Section 4.2.3 – Absorption & Emission Spectra

Metal Vapour Lamps

Metal vapour lamps use a particular heated (gas) vapour element to emit light. Two of the most common such lamps are the sodium vapour lamp and the mercury vapour lamp.

The mercury vapour lamp is a gas discharge lamp that uses an electric arc through vaporized mercury to produce light. It is typically produces blue/white light as can be seen in Figure 1 below.

Figure 1 *– Mercury vapour lamps & Mercury emission spectra*

A sodium vapour lamp is also a gas-discharge lamp that uses sodium in an excited state to produce light at a characteristic wavelength near 589 nm (yellow in colour), as can be seen in Figure 2 below.

Figure 2 *– Sodium vapour lamps & Sodium emission spectra*

Emission Spectra

You would recall that white light can be dispersed through a prism to produce a continuous spectrum ranging from violet to red and all the colours in between. Refer to Figure 3.

***Figure 3** – Continuous Spectrum*

When a sample of an element is energised, either by electricity or heat, each atom gains energy. In order to return to a stable state, they emit energy in the form of electromagnetic radiation.

An emission spectrum is the specific set of wavelengths produced by an element. The emission spectrum can be observed by dispersing the light produced by a lamp. Refer to Figure 4.

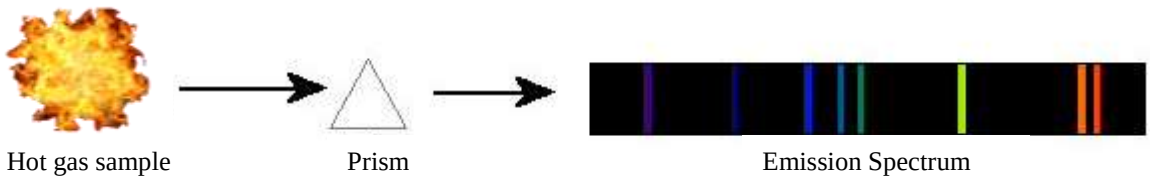


***Figure 4** – Emission Spectrum*

Each emission spectra line represents a colour (if in the visible spectrum) of particular wavelength and frequency of light that has been emitted.

NB: An emission spectra consists of coloured lines upon a black background.

When white light is passed through a cooled elemental gas some wavelengths of light are absorbed. An absorption spectrum is the specific set of wavelengths absorbed by an element. Refer to Figure 5.

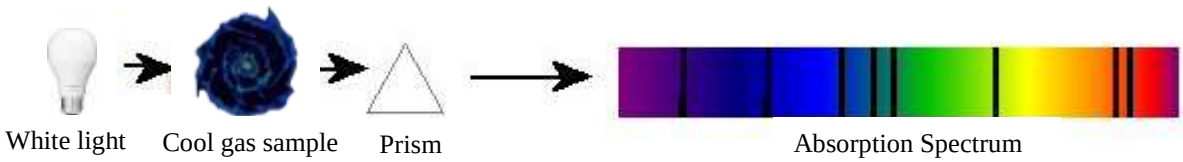


***Figure 5** – Absorption Spectrum*

Each absorption spectra line represents a colour (if in the visible spectrum) of particular wavelength and frequency of light that has be absorbed.

NB: An absorption spectra consists of black lines upon a coloured background.

The field of science involved in the analysis of light spectrum is spectroscopy. It involves separating the emitted light from an excited sample into its various spectral lines.

The combination of radiation wavelengths both emitted and absorbed is unique to each element, much like an atomic fingerprint.

Example 1

The emission and absorption spectrum for Hydrogen is shown below in Figure 6.

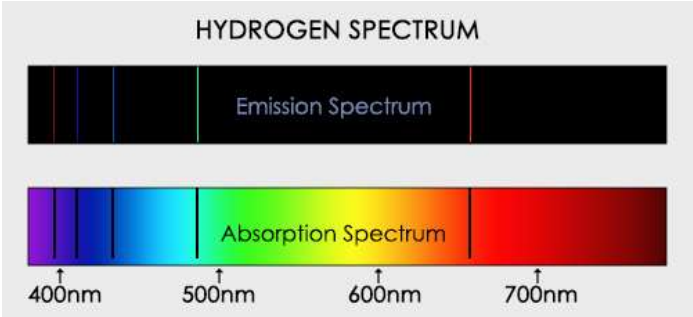


Figure 6 *– Absorption & Emission of Hydrogen*

NB: The wavelengths absorbed by an element are the same as those emitted when the same element is used in a vapour lamp.
When added together, the emission and absorption spectrum combine to produce the continuous spectrum.

Electron Energy Levels

Emission and absorption spectra can be explained with a model of the atom in which electrons occupy discrete energy levels.

In this model electrons can have only certain values of energy, according to the energy levels. Absorption and emission of light is caused by the transition of electrons between energy levels:

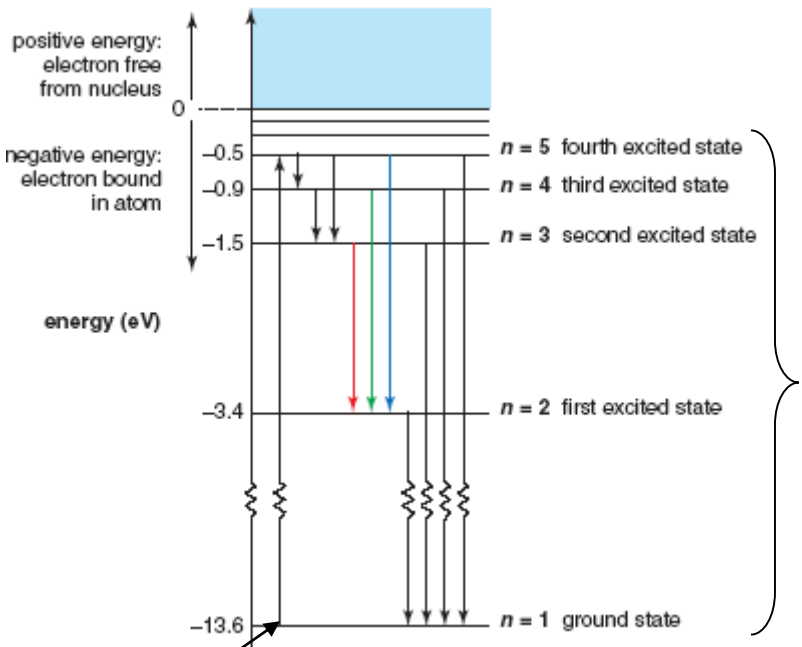
• If an electron absorbs light then it will move up to a higher energy level
• If an electron emits light then it will move down to a lower energy level

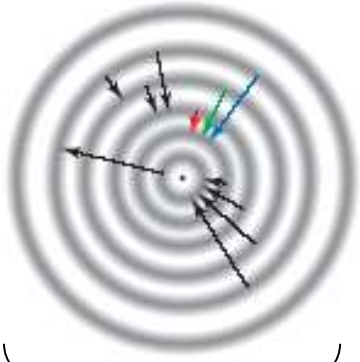
This model was proposed by Niels Bohr

Niels Bohr Atomic Model

Niels Bohr proposed an atomic model with the following features:

1. Each atom consists of a number of possible states.
   - These states have fixed or quantised energy levels
   - Electrons rotate within these fixed states following Newtonian mechanics, but do not lose electromagnetic energy





The various states or energy levels can be represented as a spectra series (left) or an orbital view (above).

Figure 7 - Atomic energy level view of the spectral series of hydrogen

The first line (directed upwards) on the energy diagram indicates the energy **received** by an electron which has been **excited** from its **ground state** to a higher **excited state** (ie. fourth excited state)

All other lines (directed downwards) represent possible energy transitions from the fourth excited state back to the ground state.

NB: Upwards arrows represent energy being absorbed by an electron.
Downwards arrow represents energy being emitted by an electron.

Task

List 5 possible energy transitions that an electron situated at the fourth excited state could undergo in order to return to the ground state (use n = 5, 4, 3, 2, 1 to label the various transitions)

Eg. 1 possible transition would be:

n(5) → n(4) → n(3) → n(2) → n(1)

2. Electrons can transfer from one state/level to another.
   - Energy must be absorbed in order to rise to a higher or excited state
   - Energy must be released, in the form of an emitted photon, in order to return to a lower state.

     $E_{photon} = hf = \Delta E$(state transition)

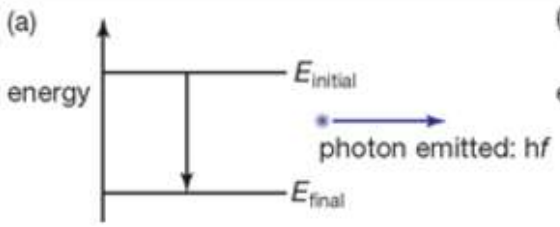


Figure 8a
Emission of light (photon)

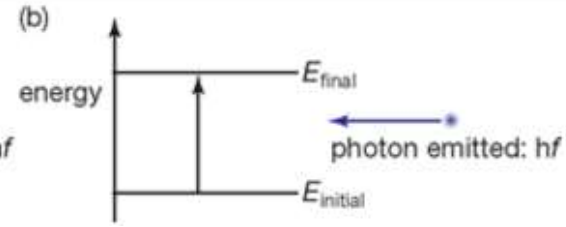


Figure 8b
Absorption of light (photon)

Emission of light: $E_{photon} = hf = \Delta E = E_{initial} - E_{final}$
Absorption of light: $E_{photon} = hf = \Delta E = E_{final} - E_{initial}$

Example 2

From the spectra series shown in Figure 7, calculate the energy and resultant wavelength of the photon:

a. Emitted when an electron "drops" from the second excited state to the ground state.
b. Required to "jump" an electron from the ground state to the first excited state.

Solutions:

a. $E_{photon} = ?$
$\lambda_{photon} = ?$
$E_3 = -1.5$ eV
$E_1 = -13.6$ eV
$h = 6.6.3 \times 10^{-34}$ Js

$E_{3\rightarrow1} = E_3 - E_1$
$= (-1.5\text{eV}) - (-13.6\text{eV})$
$= 12.1\text{eV}$
$= 12.1 \times 1.6 \times 10^{-19} = 1.94 \times 10^{-18}$ J

$E_{photon} = hf$
$= hc/\lambda_{photon}$
$\therefore \lambda_{photon} = hc/E_{photon}$
$= (6.63 \times 10^{-34} \times 3.0 \times 10^{8})/(1.94 \times 10^{-18})$
$= 1.02 \times 10^{-7}$ m (or 102 nm)

b. $E_{photon} = ?$
$\lambda_{photon} = ?$
$E_1 = -13.6$ eV
$E_2 = -3.4$ eV
$h = 6.6.3 \times 10^{-34}$ Js

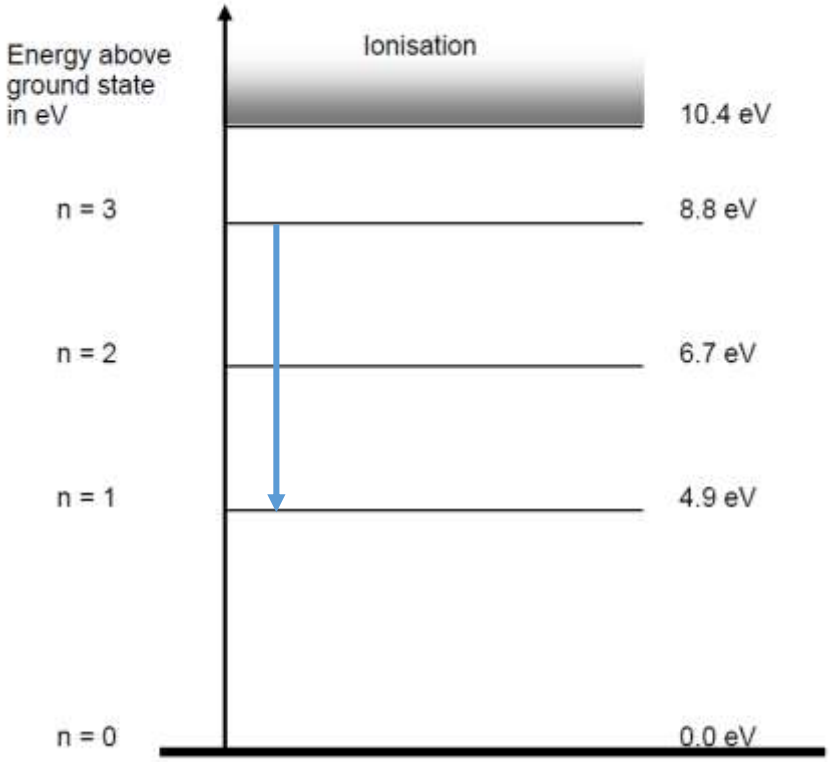
$E_{1\rightarrow2} = E_2 - E_1$
$= (-3.4\text{eV}) - (-13.6\text{eV})$
$= 10.2\text{eV}$
$= 10.2 \times 1.6 \times 10^{-19} = 1.63 \times 10^{-18}$ J

$E_{photon} = hf$
$= hc/\lambda_{photon}$
$\therefore \lambda_{photon} = hc/E_{photon}$
$= (6.63 \times 10^{-34} \times 3.0 \times 10^{8})/(1.63 \times 10^{-18})$
$= 1.22 \times 10^{-7}$ m (or 122 nm)

Exam Style Questions

Questions 1 and 2 relate to the following information.

The following figure shows the quantised energy levels of a mercury atom, relative to the ground state.



Question 1.
On the diagram draw an arrow corresponding to the emission of a 3.9 eV photon.

ΔE between two levels must equal 3.9 eV, so the transition must be from n=3 to n=1. The arrow must point down since when an electron drops down to a lower energy level it <u>emits</u> a photon.

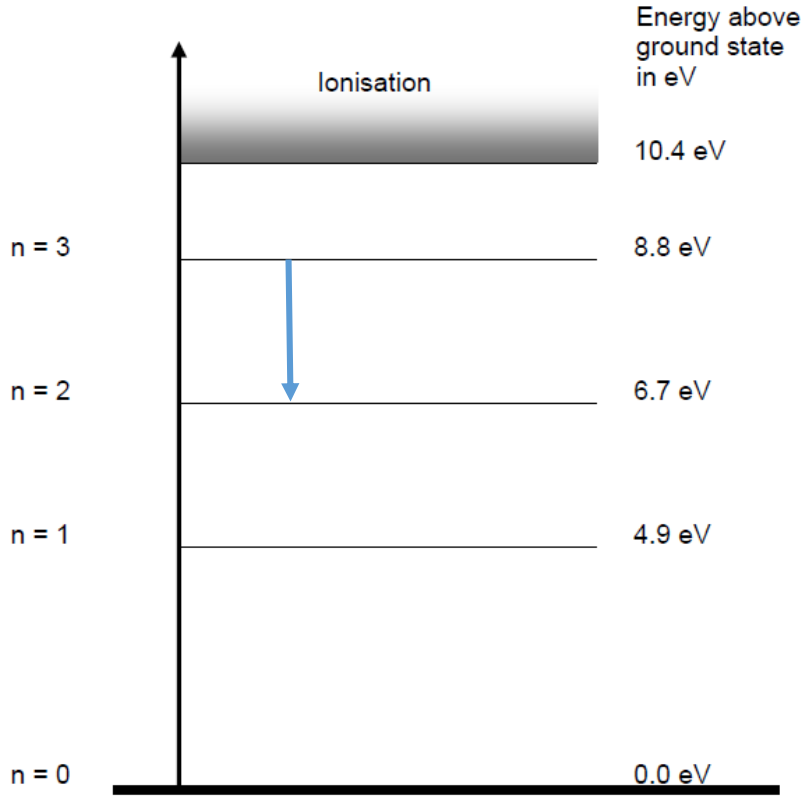
Question 2.
What is the highest frequency photon that can be emitted when an electron decays from the n = 3 level?

ΔE = hf ∴highest f corresponds to the highest ΔE ie. from n=3 to n=0
$F = \Delta E/h = 8.8/(4.14 \times 10^{-15}) = 2.13 \times 10^{15}$ *m*
NB: *Both E and h are using eV*

2.1×10^{15} Hz

Questions 3 and 4 relate to the following information.

The following figure shows the quantised energy levels of a mercury atom, relative to the ground state.



Question 3.

On the diagram draw an arrow corresponding to the emission of a 2.1 eV photon.

ΔE between two levels must equal 2.1 eV, so the transition must be from n=3 to n=2. The arrow must point down since when an electron drops down to a lower energy level it <u>*emits*</u> *a photon.*

Question 4.

What is the longest wavelength photon that can be emitted when an electron decays from the n = 2 level?

$\Delta E = hf = hc/\lambda$

$\therefore$ *longest* λ *means lowest* ΔE *ie. from n=2 to n=1 (1.8 eV)*

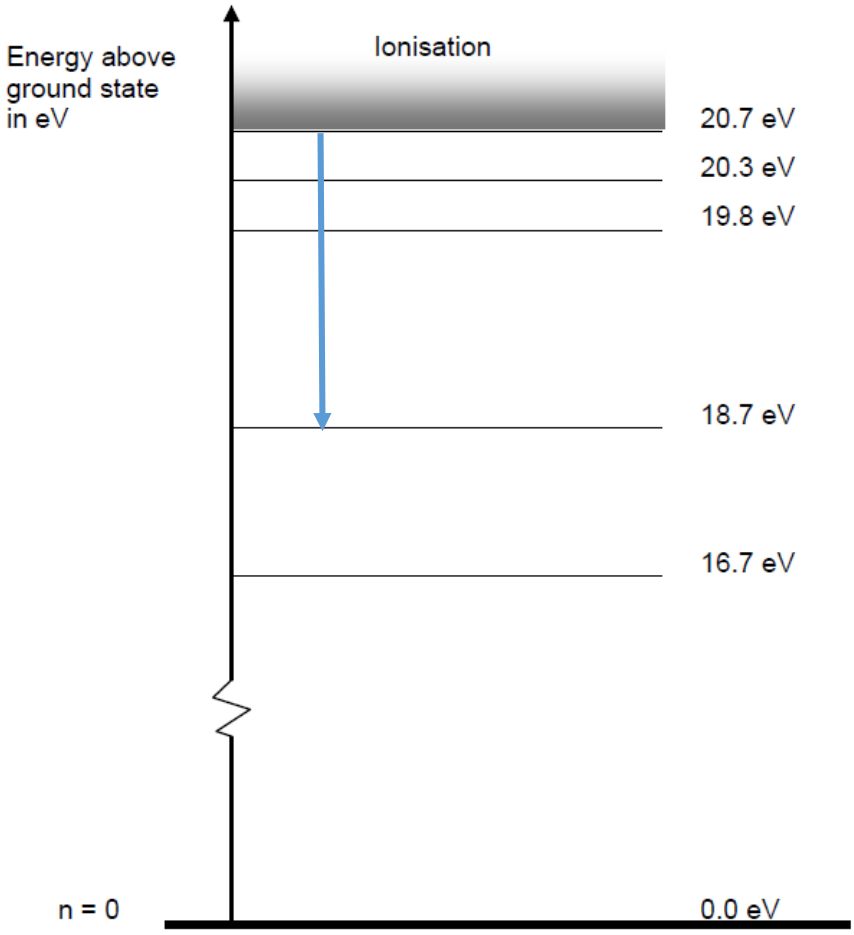
$\lambda = hc/E = (4.14 \times 10^{-15} \times 3.0 \times 10^{8})/1.8 = 6.9 \times 10^{-7} m$

NB: *Both E and h are using eV*

6.9×10^{-7} m

Questions 5 and 6 relate to the following information.

The following figure shows the quantised energy levels of a neon atom, relative to the ground state, as used in a He-Ne laser. The lasing energy produces a 632.8 nm photon beam.



Question 5.

What is the energy, in eV, of a 632.8 nm photon?

$E = hc/\lambda$
$= (4.14 \times 10^{-15} \times 3.0 \times 10^{8})/632.8 \times 10^{-9}$
$= 1.96$ eV

NB: *Both E and h are using eV*

1.96 eV

Question 6.

On the energy level diagram draw an arrow which **best** corresponds to the emission of the 632.8 nm lasing photon.

On the diagram ΔE between levels must equal 1.96 or approximately 2 eV, so the transition must be between 20.7 and 18.7 eV. The arrow must point down since when an electron drops down to a lower energy level it emits a photon.