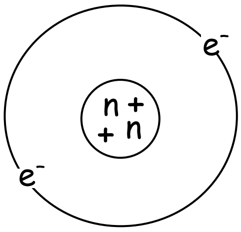
1.1 Electromagnetic Radiation & Atomic Spectra

Electronic Structure



Particle	Position	Charge	Mass (amu)
Proton	In nucleus	+1	1
Neutron	In nucleus	0	1
Electron	Outside nucleus	–1	$^1/_{1840}$

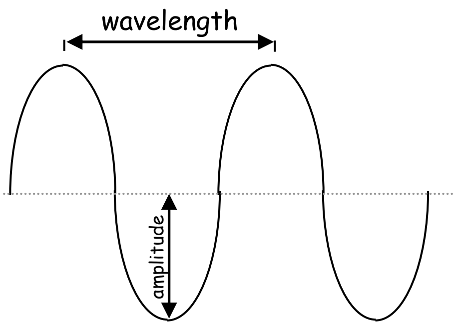
Electromagnetic Radiation

James Clerk-Maxwell developed a theory for describing electromagnetic radiation as wave phenomena in 1864.

Types of EM Radiation

EM Radiation	Gamma rays	X rays	UV radiation	Visible light	Infra-Red radiation	Microwaves	Radio & TV waves
Wavelength	10^{-14}m	10^{-10}m	10^{-8}m	10^{-7}m violet 450nm, red 700nm	10^{-6}m	10^{-3}m	10^{-1} to 10^{4}m
Frequency	high →						low
Energy	high →						low
Uses	sterilising, cancer treatment	picturing broken bones	sunbeds		heater	cooking	home

Describing EM Waves



λ wavelength – measured in m or nm
the distance between wave crest to wave crest

f frequency – measured in s^{-1} or Hertz Hz
the number of waves per second

c velocity – speed is 3×10^8 m s^{-1}
speed of the waves

$\bar{\nu}$ wavenumber – measured in m^{-1} or cm^{-1}
1/wavelength

$$c = \lambda \times f$$

speed (m s^{-1}) = wavelength (m) x frequency (Hz or s^{-1})

Question

1. Radio 4 on LW has a wavelength of 1500m. Calculate the frequency of this wave.
2. Radio Clyde 1 has a frequency of 102.5MHz.
    a) Calculate the wavelength in m.
    b) Calculate the wavenumber in cm^{-1}.

Waves and Particles

Under certain circumstances, EM radiation can be regarded as a stream of very small particles called *photons* rather than waves. These photons transfer small bundles (packets) of energy.

This is called *Wave – Particle Duality*.

The energy of photons of radiation can be calculated by:

$$E = h \times f$$

energy (J) = Plank's constant (J s) x frequency (Hz or s^{-1})

where Plank's constant h = 6.63 x 10^{-34} J s

therefore

high frequency ⟷ high energy
(low wavelength)

for 1 mol of photons

$$\mathbf{E = L \times hf}$$

where L is Avogadro's number = 6.02×10^{23} mol^{-1}

$$E = Lhf \quad \text{but} \quad \mathbf{f} = \frac{c}{\lambda}$$

$$\mathbf{E} = \frac{Lhc}{\lambda}$$

and $1/\lambda = \bar{\nu}$ $\qquad$ $\mathbf{E} = Lhc\bar{\nu}$

<u>Questions</u>

1. Calculate the energy, in kJ mol^{-1}, corresponding to

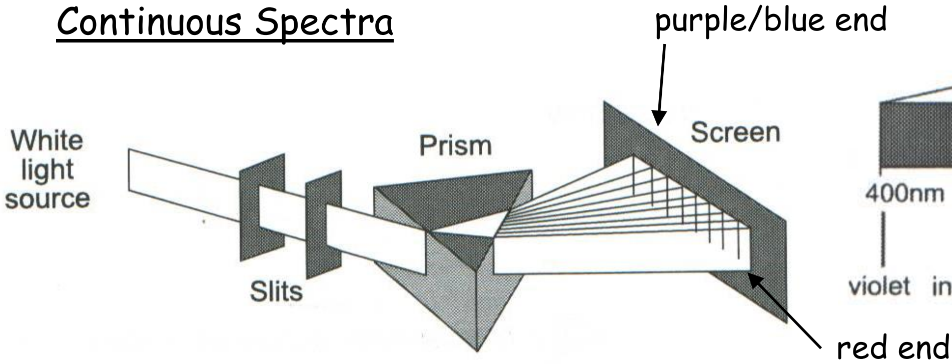
    a) a wavenumber of 1000cm^{-1}

    b) a wavelength of 620 nm.

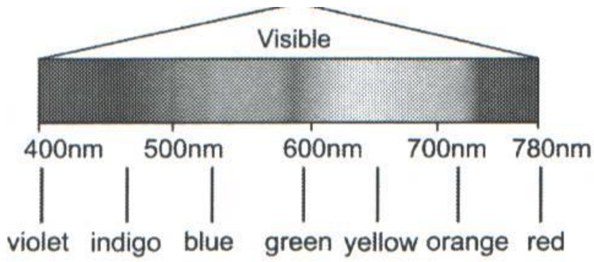
2. The bond enthalpy of a Cl – Cl bond is 243kJ mol^{-1}. Calculate the maximum wavelength of light that would break one mole of these bonds to form individual chlorine atoms.

Spectroscopy

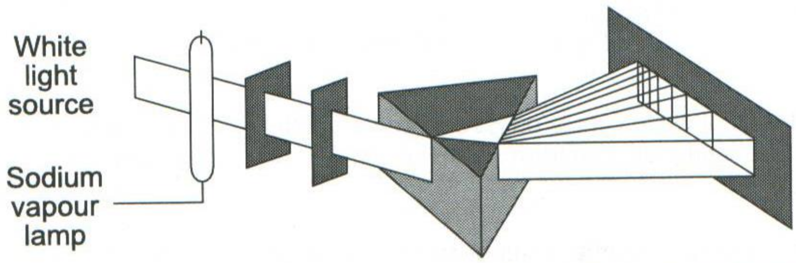
Continuous Spectra



Continuous spectrum



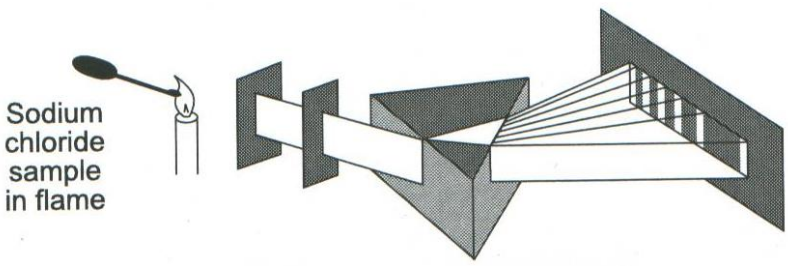
Line Absorption Spectra



Line absorption spectrum

- Continuous colour spectrum produced but with dark lines
- Dark lines represent certain wavelengths which have been removed by absorption by the source (in the diagram a sodium vapour lamp)

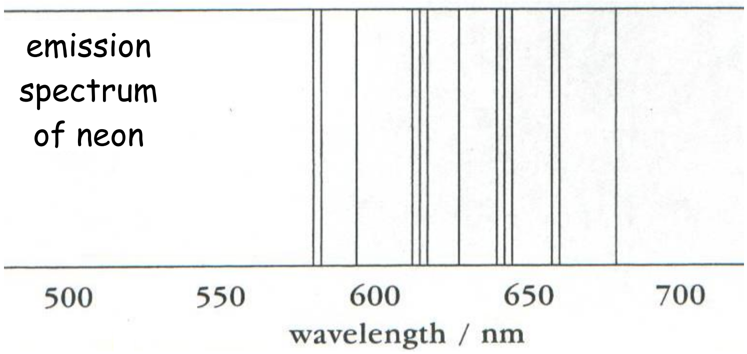
Line Emission Spectra



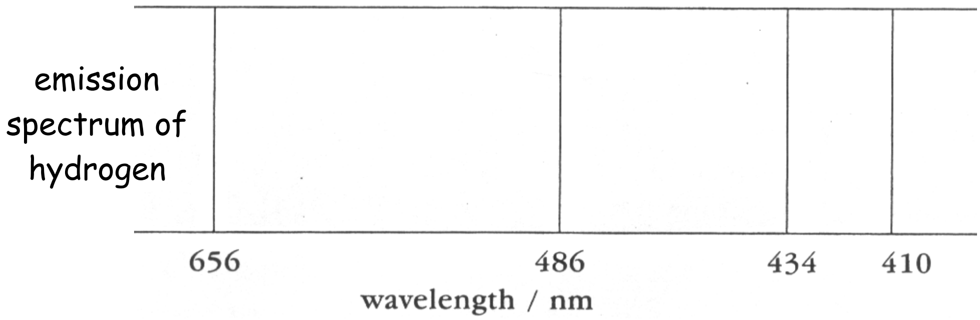
Line emission spectrum

- Series of lines produced (not a continuous spectrum)
- Each line represents a distinct wavelength of emitted light
- The position of each line leads to an emission spectrum of each element

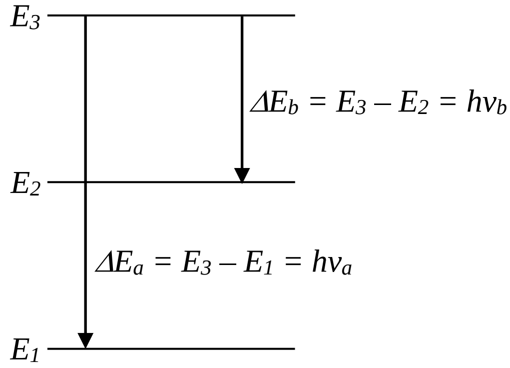
<u>Using Emission Spectra</u>



- Glowing neon light have a distinctive red colour
- Red colour observed is a combination of a series of lines (mainly from the red end of the visible spectrum)

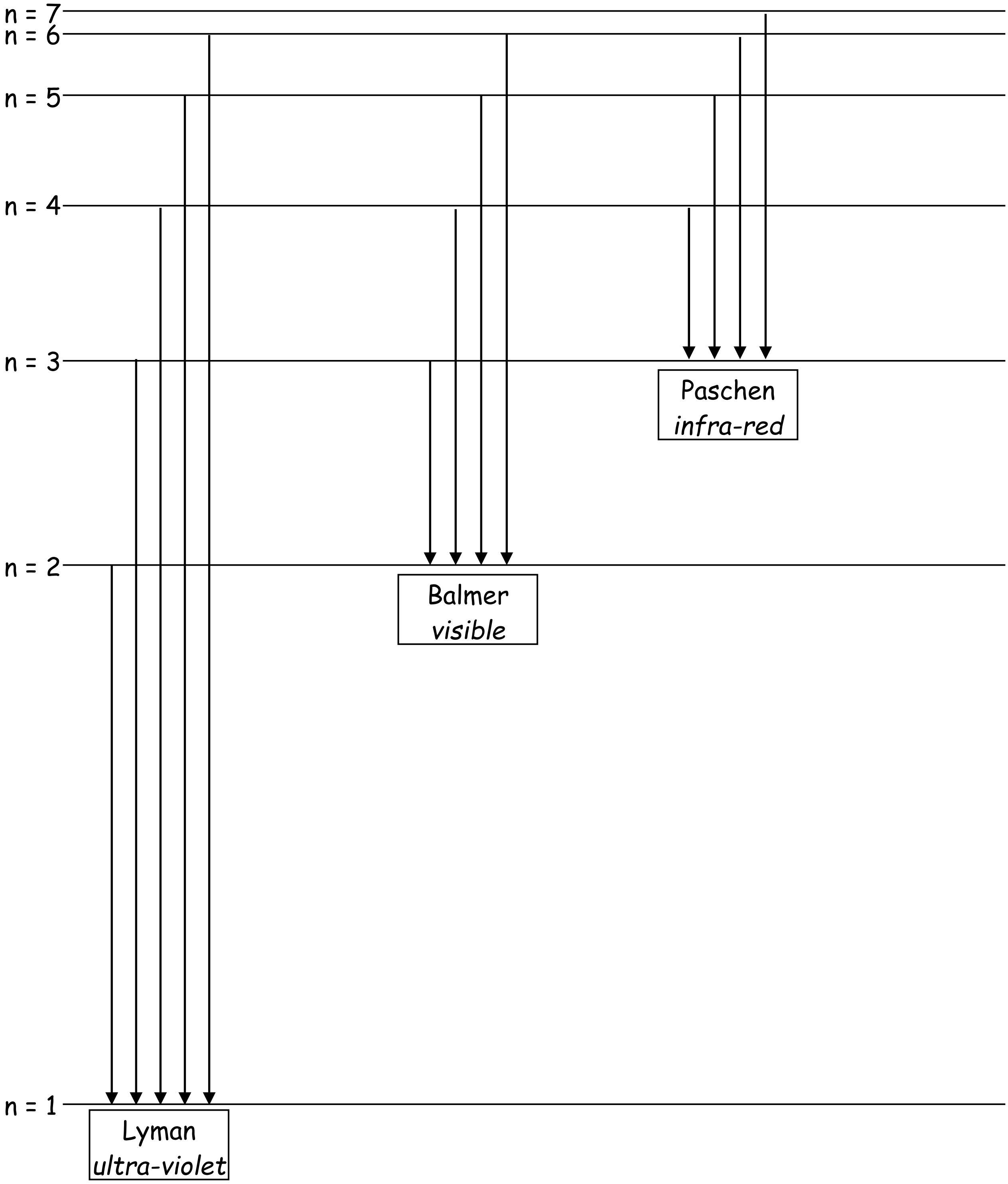


- atoms are *excited* when they have absorbed energy
- when energy is absorbed, electrons move up to an outer shell
- energy, in the form of a photon, is released when electrons move to a lower energy level (electron shell)
- the energy is released as a particular wavelength of light (ie $E=\frac{hc}{\lambda}$)
- differences between energy levels are fixed in any particular atom
  - electrons at each level have fixed energy
  - energy of electrons is <u>quantised</u>
- each wavelength/line on emission spectra is formed from a specific energy difference as electrons fall back from their excited state.
  - This allows various energy levels to be calculated from each wavelength obtained from an emission spectra.

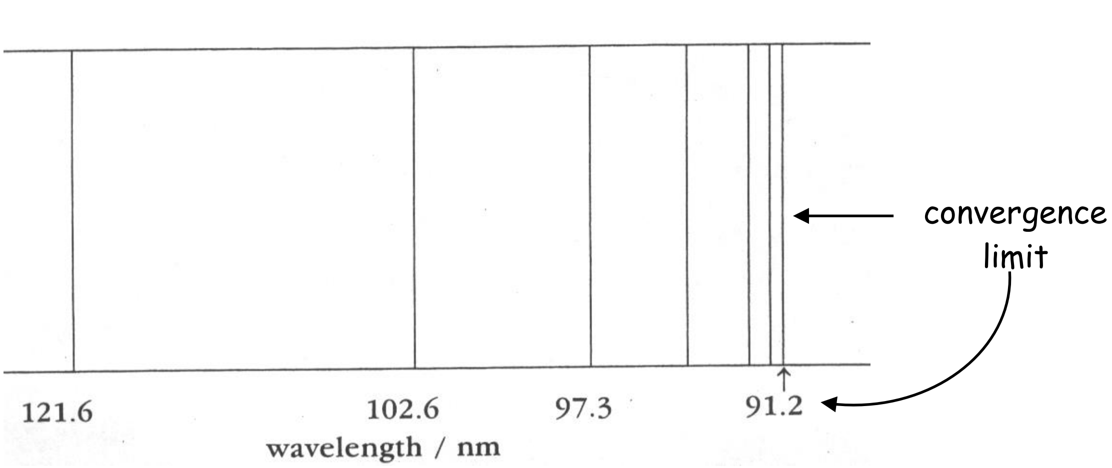


The full emission spectrum of hydrogen has one set of lines in the uv region, one in the visible region and several in the IR region of the EM spectrum.

Name of Series	Energy level to which excited electron falls	Region of the Electromagnetic Spectrum of Photon
Lyman	n = 1	ultra-violet
Balmer	n = 2	visible
Paschen	n = 3	infra-red
Brackett	n = 4	infra-red
Pfund	n = 5	infra-red



Emission spectrum of hydrogen
<u>Lyman Series</u>



- UV light emitted as electrons return to energy level n = 1 (innermost electron shell)
- Lines on emission spectrum get closer together and converge
- The energy difference between ground state and the convergence limit is called the *ionisation energy*.

<u>Question</u>

1. Calculate the ionisation energy for hydrogen if the wavelength of the line at the convergence limit is 91.2nm.

However,

Bohr's work on hydrogen emission spectra was unable to explain the behaviour of bigger elements. In deed as equipment improved, some lines on emission spectra proved to be doublets or triplets under higher resolution.

This was later explained by the fact that electron shells have electron subshells within them.