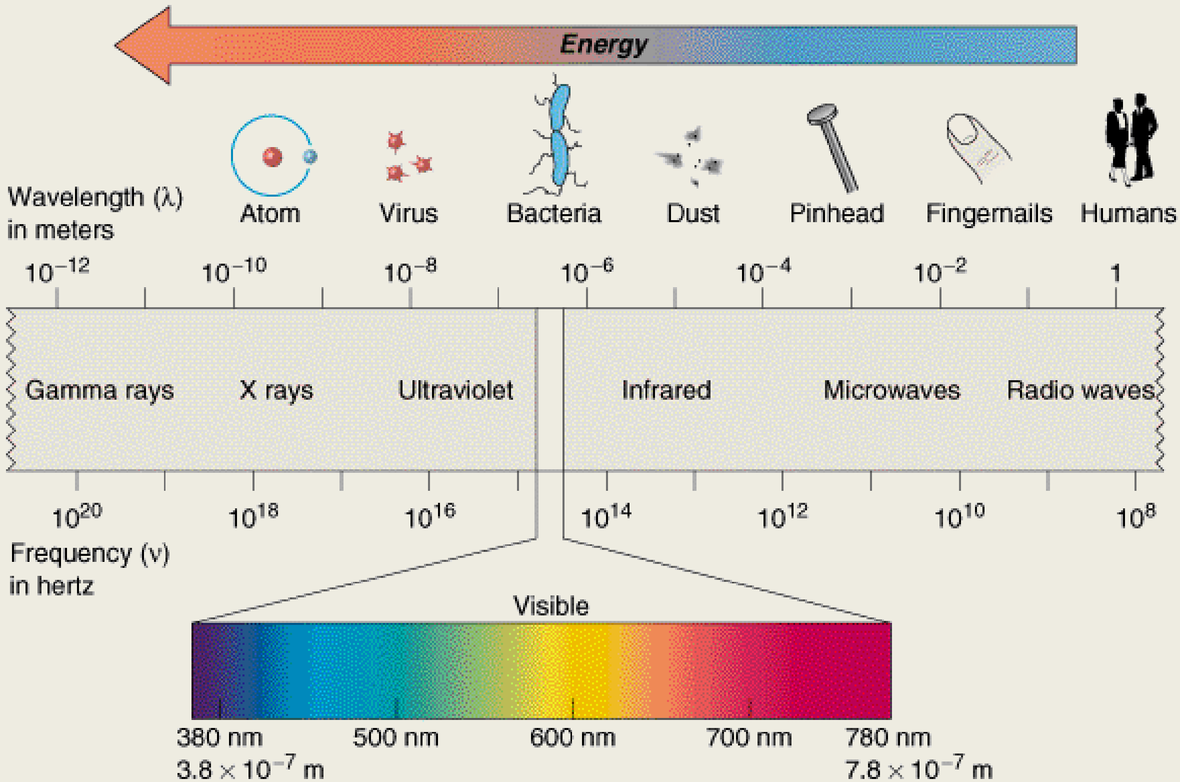
AH Chemistry – Unit 1

Electromagnetic Radiation and Atomic Spectra

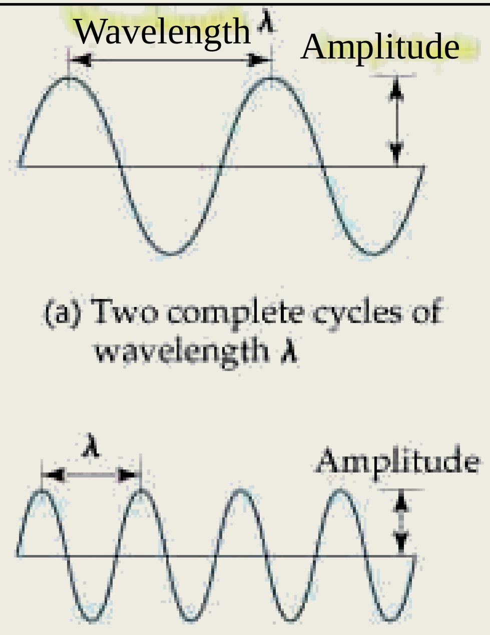
Introduction



This topic describes the ***Electromagnetic Spectrum*** and how it can interact with atoms, ***Spectroscopy***.

Much information about ***Electronic Structure*** comes from spectroscopic evidence.

The Wave Nature of Light



(a) Two complete cycles of wavelength λ

(b) Wavelength half of that in (a); frequency twice as great as in (a)

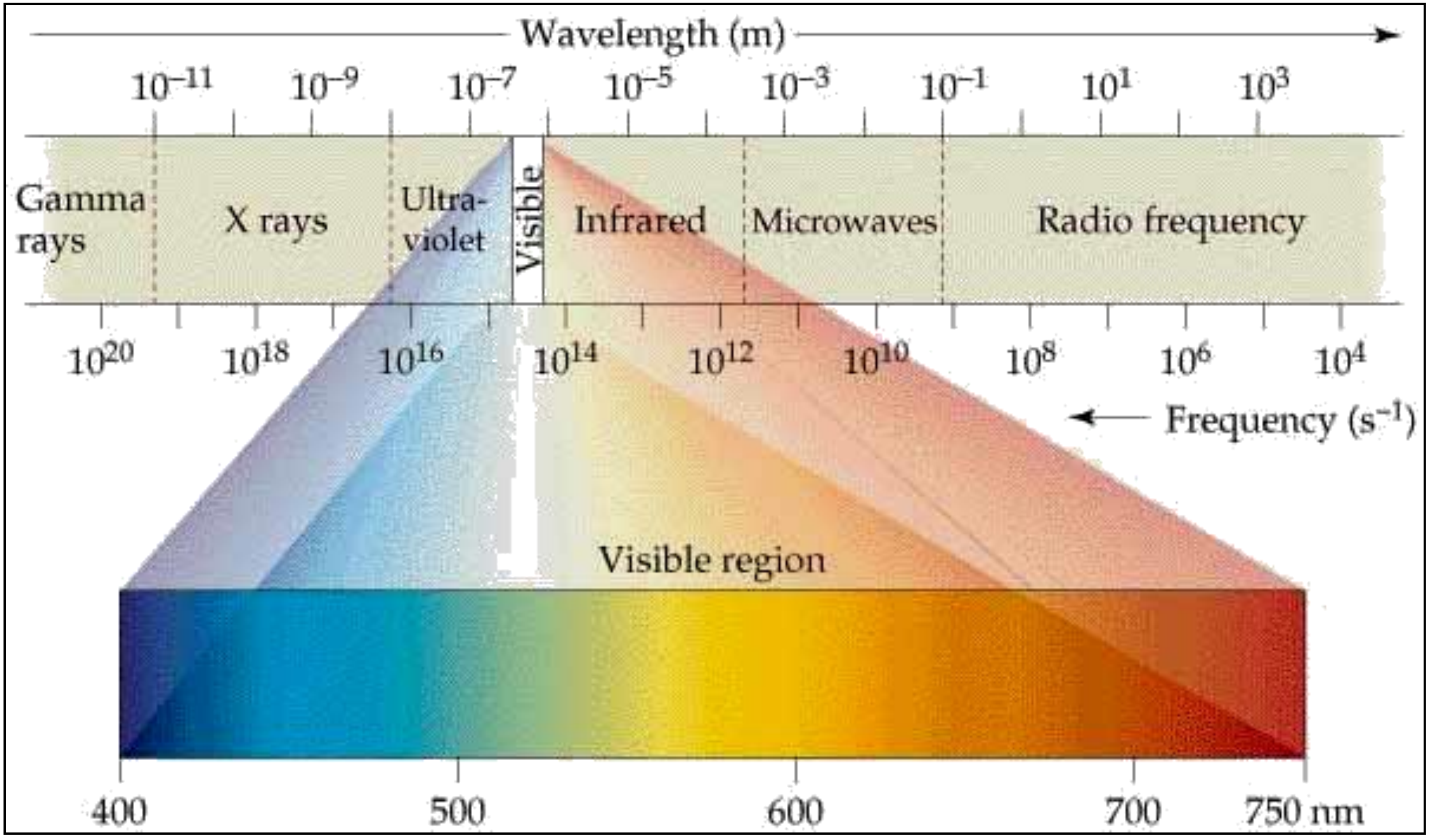
(c) Same frequency as (b), smaller amplitude

- All waves have a characteristic ***wavelength***, λ, measured in metres (m) to nanometres (nm)
- The ***frequency***, ν, of a wave is the number of waves which pass a point in one second measured in Hertz (Hz) or per seconds (s^{-1})
- The ***speed*** of a wave, *c*, is given by its frequency multiplied by its wavelength:

$$c = \nu\lambda$$

- For light, $c = 3.00 \times 10^8\ ms^{-1}$
- Another unit of 'frequency' used in spectroscopy is the ***wavenumber*** ($1/\lambda$), ***nu*** $\bar{\nu}$, measured in m^{-1} or cm^{-1}

Electromagnetic Radiation



The Particle Nature of Light 1

Planck*:** energy can only be absorbed or released from atoms in certain amounts called ***quanta

To understand ***quantization***, consider the notes produced by a violin **(*continuous*)** and a piano **(*quantized*):**

a violin can produce ***any note*** *by placing the fingers at an appropriate spot on the fingerboard.*

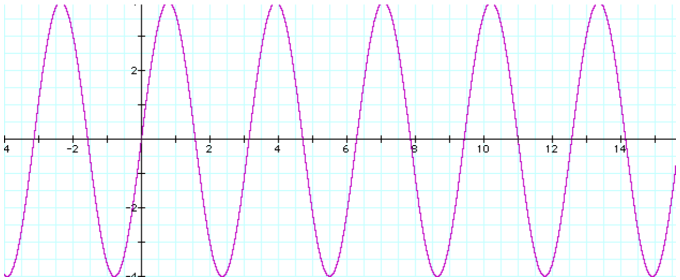
A piano can only produce ***certain notes*** *corresponding to the keys on the keyboard.*

The Particle Nature of Light 2

Electromagnetic Radiation can also be thought of as a stream of very small particles known as ***photons***

Electromagnetic Radiation exhibits wave-particle ***dual properties*.**

The ***energy*** **(E)** of a ***photon*** (particle) is related to the ***frequency*** (wave) of the radiation as follows:



$$E = h\nu$$

where h is ***Planck's constant*** (6.63×10^{-34} J s).

Energy Calculations 1

$$E = h\nu \qquad \text{or} \qquad E = h\,c\,/\,\lambda$$

The ***energy*** calculated would be in Joules (J) and would be a very small quantity.

Normally, we would calculate the energy transferred by the ***emission*** or ***absorption*** of ***one mole of photons*** as follows:

$$E = Lh\nu \quad \text{or} \qquad E = L\,h\,c\,/\,\lambda$$

Where L is the ***Avogadro Constant***, 6 .02 x 10^{23} and E would now be in J mol^{-1} or kJ mol^{-1}.

Energy Calculations 2

For example, a neon strip light emitted light with a wavelength of 640 nm.

640 nm = 640 x 10^{-9} m = 6.40 x 10^{-7} m.

For each ***photon*****:**

$E \quad = \quad h \quad c \quad / \quad \lambda$

$= 6.63 \times 10^{-34} \times 3.00 \times 10^{8} / 6.40 \times 10^{-7}$

$= 3.11 \times 10^{-19}$ J

For ***1 mole of photons*****:**

$E = 3.11 \times 10^{-19} \times 6.02 \times 10^{23}$ J

$= 1.87 \times 10^{5}$ J mol^{-1}

$= 187$ kJ mol^{-1}

Atomic Spectra

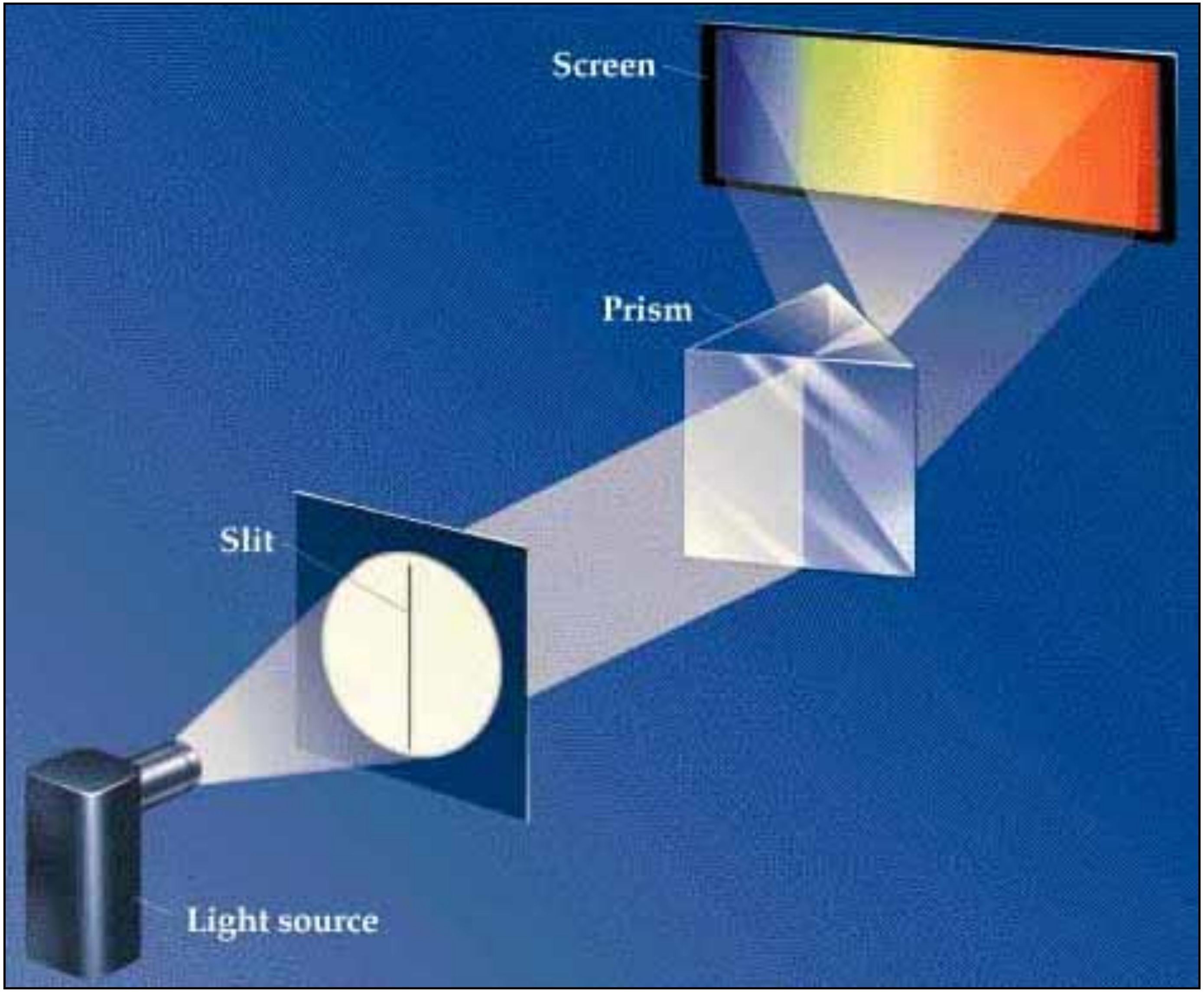
Atomic emission spectra provided significant contributions to the modern picture of atomic structure

Radiation composed of only one wavelength is called ***monochromatic***

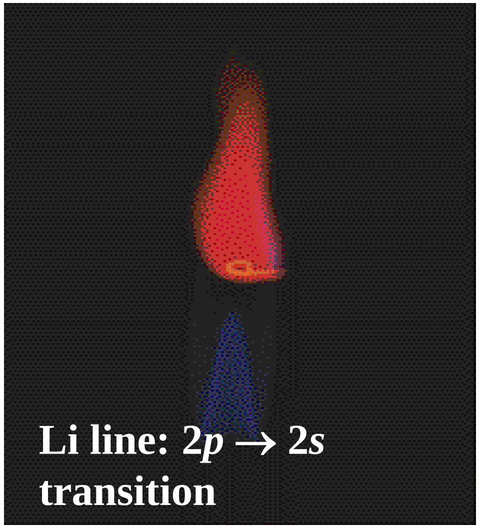
Radiation that spans a whole array of different wavelengths is called ***continuous***

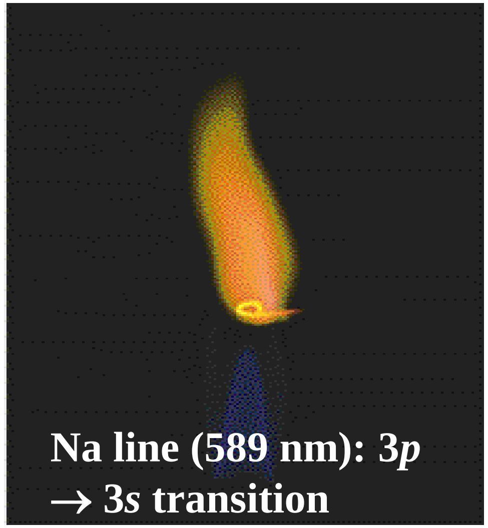
White light can be separated into a continuous ***spectrum*** of colors.

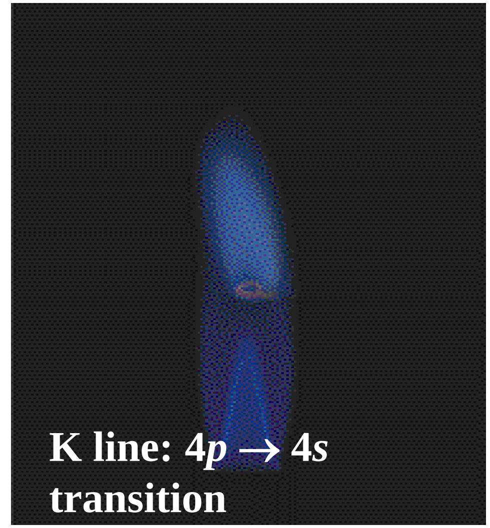
Atomic Spectra 2



Emission Spectroscopy



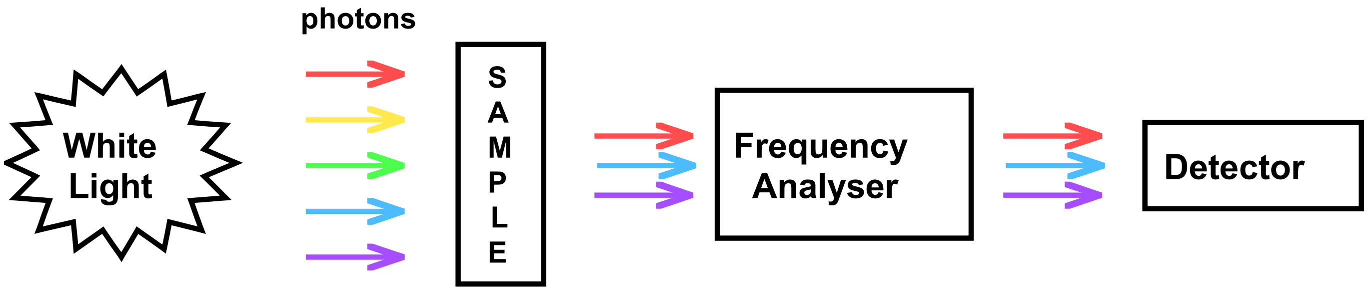




In ***emission spectroscopy***, light of certain wavelengths is emitted as 'excited' electrons drop down from higher energy levels.

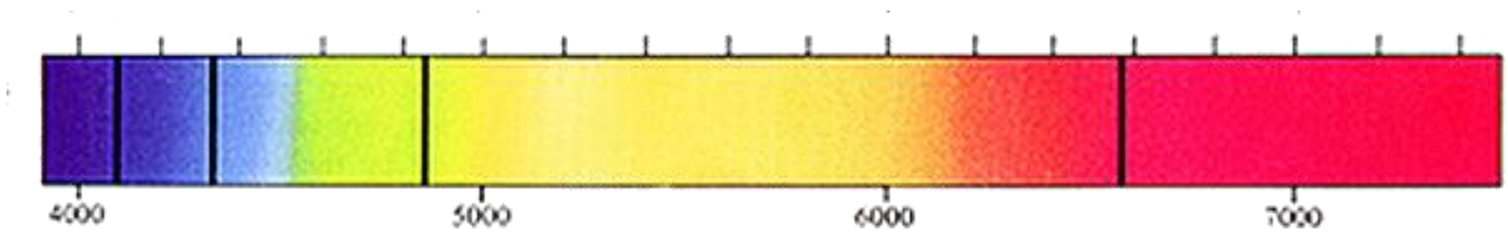
The spectrum is examined to see the wavelengths emitted.

Absorption Spectroscopy



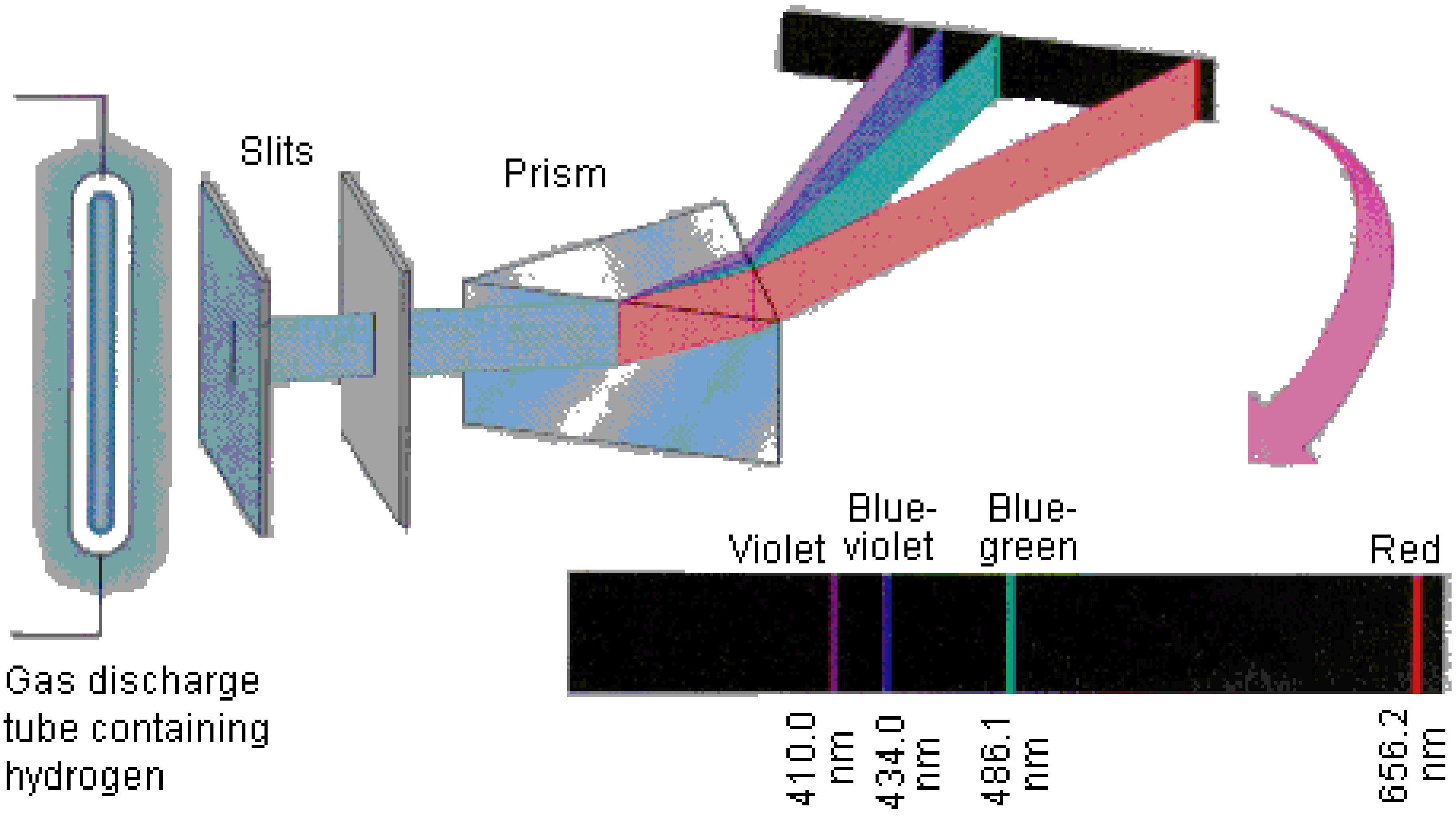
In ***absorption spectroscopy***, light of certain wavelengths is absorbed and electrons are promoted to higher energy levels.

The spectrum is examined to see which wavelengths have been absorbed.



The ***intensity*** of the light absorbed is proportional to the ***quantity*** of the atoms/ions in the sample. Calibration samples can be prepared, intensities measured and unknown concentrations determined (PPA)

Hydrogen Emission Spectrum



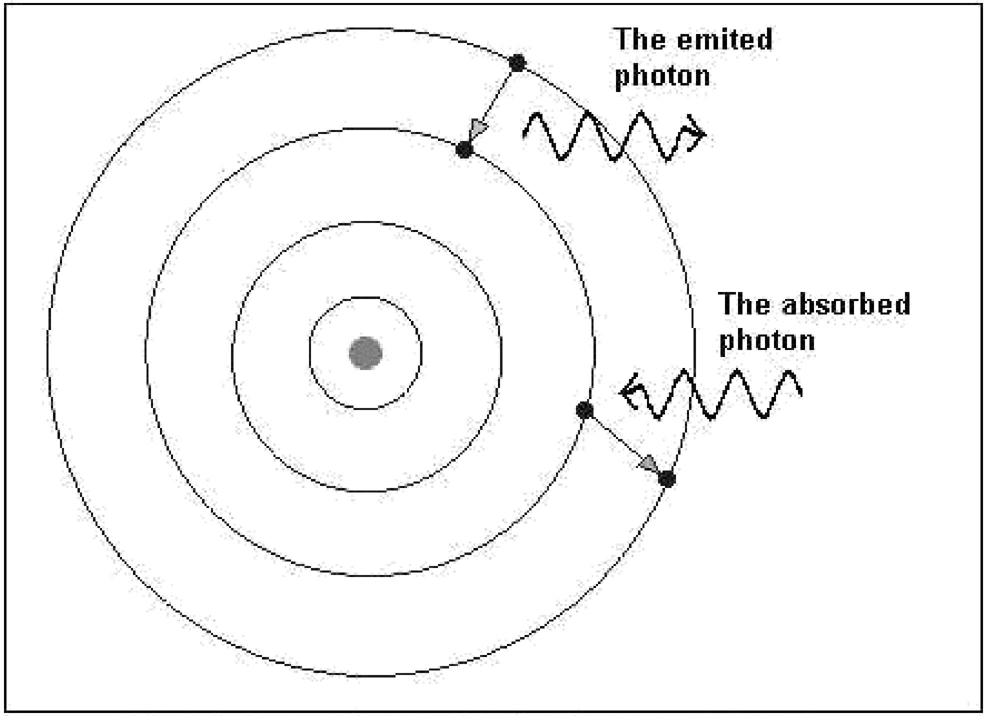
Line Spectra

The spectrum obtained when hydrogen atoms are ***excited*** shows four lines: red, blue-green, blue and indigo

Bohr deduced that the colours were due to the movement of ***electrons*** from a higher energy level back to the 'ground state'.

The significance of a Line Spectrum is that it suggests that electrons can only occupy certain ***fixed energy levels***

Bohr's Model



The electron emits or absorbs the energy changing the orbits.

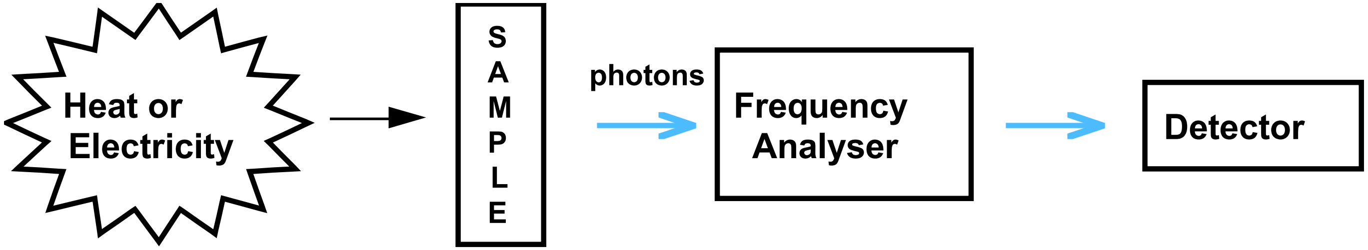
These fixed energy levels are what we have always called our ***electron shells***

A photon of light is emitted or absorbed when an electron changes from one energy level (shell) to another.

Emission Spectroscopy 2

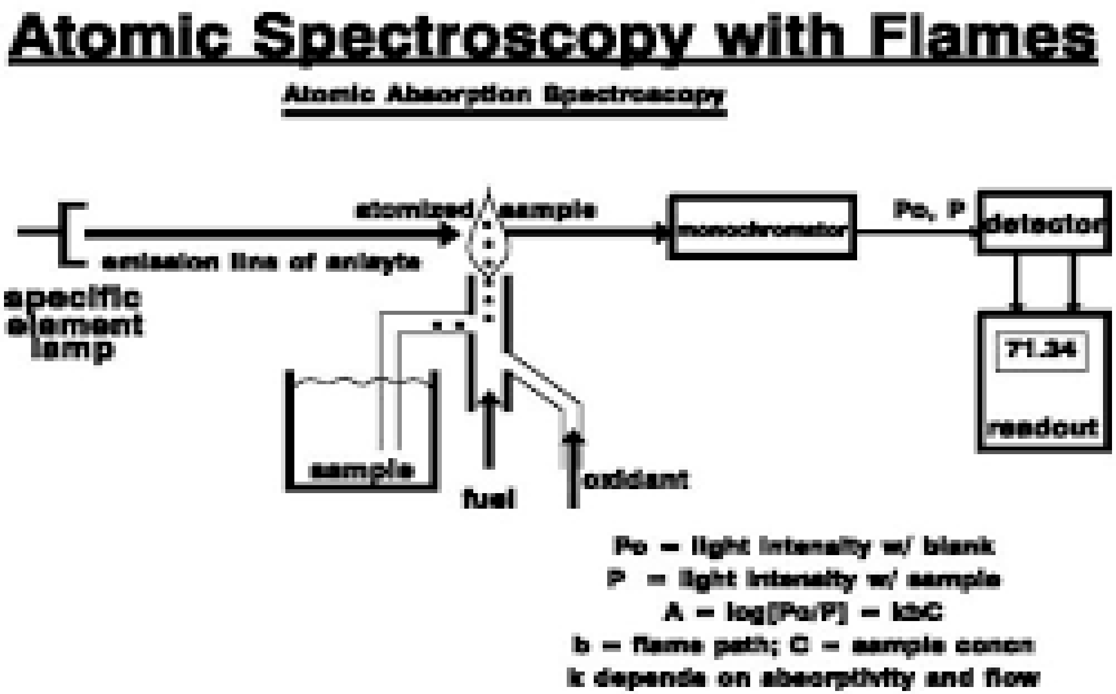
Each element provides a ***characteristic spectrum*** which can be used to ***identify*** the element. Analysing light from stars etc, tell us a lot about the elements present.

The ***intensity*** of a particularly strong line in an element's spectrum can be measured. The ***intensity*** of the light emitted is proportional to the ***quantity*** of the atoms/ions in the sample. ***Calibration samples*** can be prepared, intensities measured and unknown concentrations determined. ***Analytical tool***.



How much of an element?

If a blank or reference sample is introduced, it is possible to quantify how much of an element is present in a sample, as the intensity of light absorbed will be proportional to concentration.



$$A \, \alpha \, C$$