The Development of Atomic Theory

Democritus (400 BC)	- suggested that matter is composed of indivisible particles called atoms.
John Dalton (1803)	- reintroduced the atomic theory based on the available evidence
J.J. Thomson (1897)	- identified electrons in cathode rays - proposed the plum pudding model: a sphere of positive electrical charge in which electrons are embedded.
Ernest Rutherford (1911)	- carried out the Gold Foil Experiment in which positive α-particles were deflected by gold foil, proving the existence of the nucleus - all positive charge and most of the mass is concentrated in the nucleus of the atom. The electrons are found in the empty space around the nucleus.
James Chadwick (1932)	- discovered the other heavy nuclear particle, the neutron.

Summary of Subatomic Particles

Particle	**Location**	**Charge**	**Mass (u)**
proton	nucleus	+ 1	1.0073
neutron	nucleus	0	1.0087
electron	outside nucleus	- 1	0.00055

Electromagnetic Radiation

All electromagnetic radiation is made up of (you guessed it) electric and magnetic fields. Each type of electromagnetic radiation has its own wavelength, frequency and energy. As well as having wave-like properties, electromagnetic radiation also has particle-like properties. The elementary particle of electromagnetic radiation is the photon and all forms of electromagnetic radiation travel at the speed of light ($3.00 \times 10^8 m/s$).

Basic Wave Terms

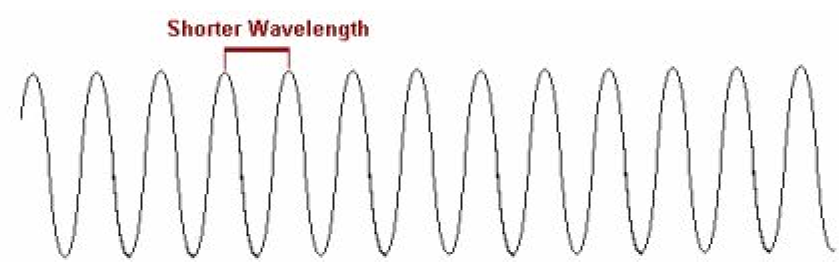
Wavelength (lambda or λ): The length of one complete wave.
Units: metres (m)

Frequency (*f*): The number of waves that will pass a fixed point per unit time.
Units: Hertz (Hz) or cycles per second (c/s)

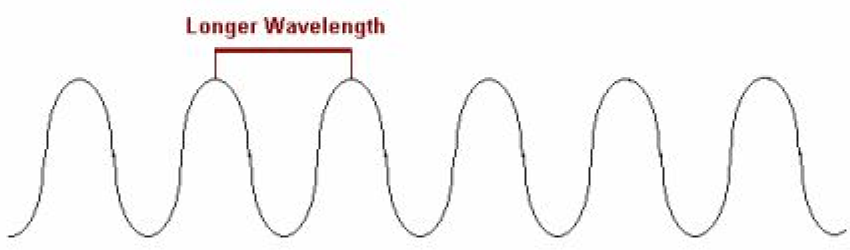
Assume two electromagnetic waves are traveling at the speed of light. Wave “A” has shorter wavelength and a higher frequency. Wave “B” has a longer wavelength and lower frequency:

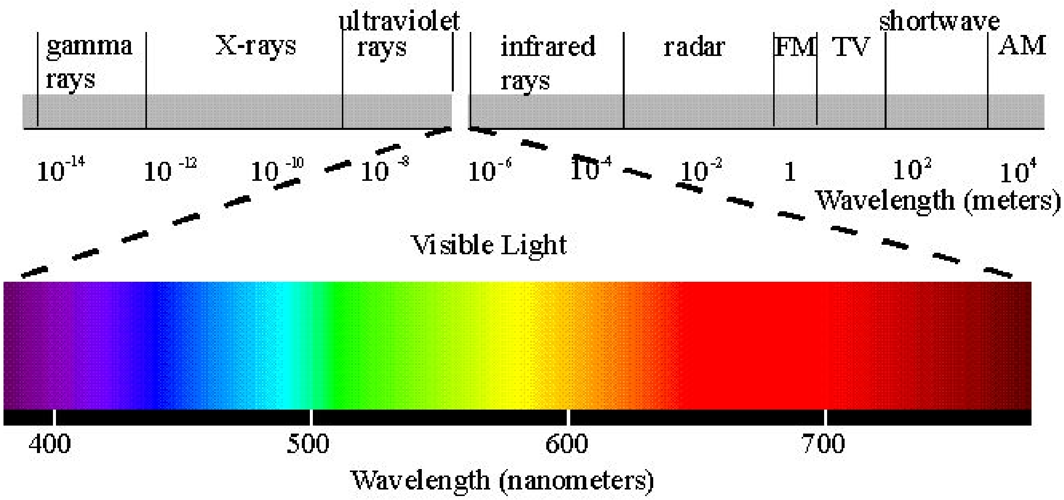
direction of propagation $\longrightarrow$

A.



B.





(1 nm = 1 x 10^{-9} m)

The Spectrum of Electromagnetic Radiation

Name of Radiation	Wavelength (m)	Frequency (c/s or Hz)
Radio Waves	**10^{2}**	**10^{6}**
Microwaves		
Radar Waves		
Infrared Light	**Increasing**	**Increasing**
Visible Light *	**10^{-7}** **Wavelength**	**10^{14}** **frequency**
UV Light		
X-Rays		
Gamma Rays	**10^{-14}**	**10^{22}**

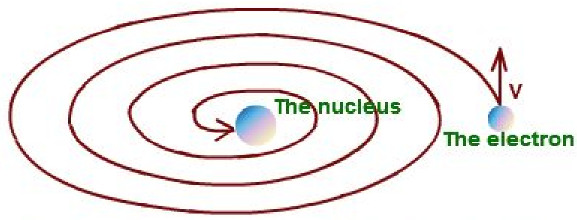
The Spectrum of Visible Light

Colour	Wavelength (m)	Frequency (c/s or Hz)
Red	6.5 x 10^{-7}	4.6 x 10^{14}
Orange		
Yellow		
Green		
Blue		
Indigo		
Violet	4.1 x 10^{-7}	7.3 x 10^{14}

THE DEVELOPMENT OF BOHR'S ATOMIC THEORY

PROBLEMS WITH RUTHERFORD'S PLANETARY MODEL OF THE ATOM:

1) According to the classical laws of physics, as a charged particle changes direction in space, it should give off energy. Therefore, as electrons orbit around a nucleus, they should lose energy by emitting light. This loss of energy should cause the electrons to spiral into the nucleus, resulting in the collapse of the atom. The Rutherford atom would be unstable.



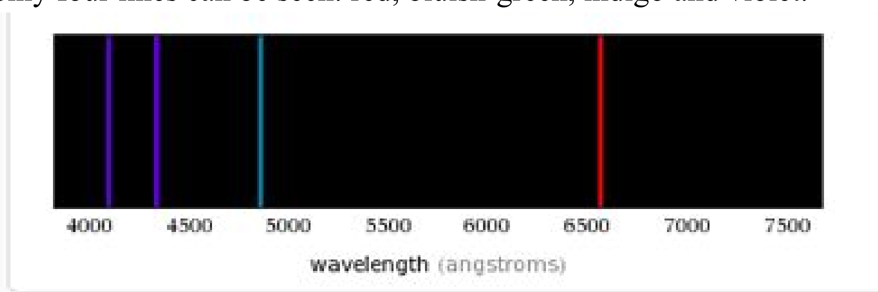
In the planetary model of atom, the electron should emit energy and spirally fall on the nucleus.

2) If white light is passed through a prism, a continuous spectrum of colours is observed. Each colour in the spectrum has a different energy; from red with the lowest energy to violet with the highest energy. However, it was well known that excited atoms emit energy as a characteristic line spectrum.

Atoms do not emit a continuous spectrum, nor do they collapse, Rutherford's model had to be modified.

NIELS BOHR (1913)

When a sample of an element is sealed in a discharge tube and subjected to a high voltage, light is given off by the element in the tube. When this light is passed through a spectrometer, a line spectrum is observed as opposed to a continuous spectrum. Furthermore, each element gives off a characteristic line spectrum. For example, if the light emitted by hydrogen is passed through a spectrometer only four lines can be seen: red, bluish-green, indigo and violet.



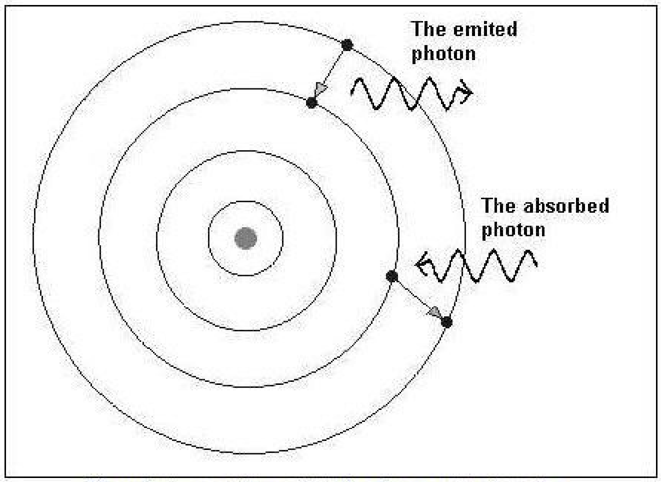
Bohr developed a model of the atom which explained the line spectrum of hydrogen and why the atom doesn't collapse. His theory is made up of two postulates:

Postulate 1:
Electrons can only move in certain fixed orbits. Each orbit corresponds to a specific energy level and an electron can move within an orbit without losing any energy.

Postulate 2:
An electron can only move from one orbit (or energy level) to another when it gains or loses energy.

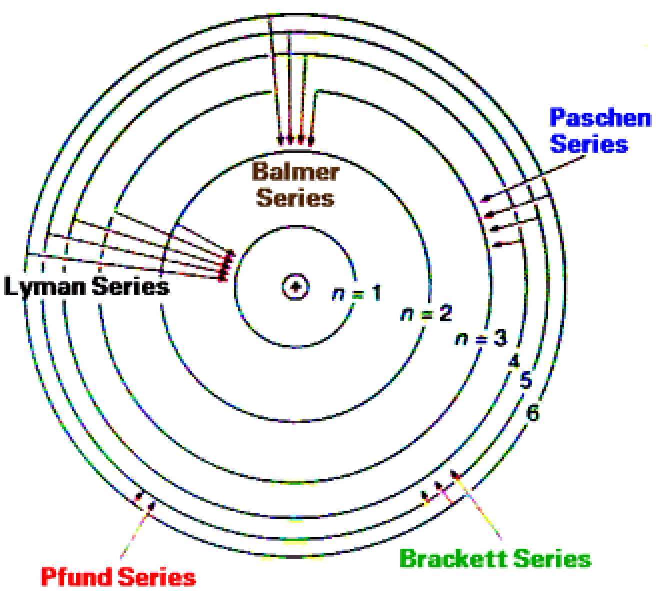
BOHR'S EXPLANATION FOR THE HYDROGEN SPECTRUM

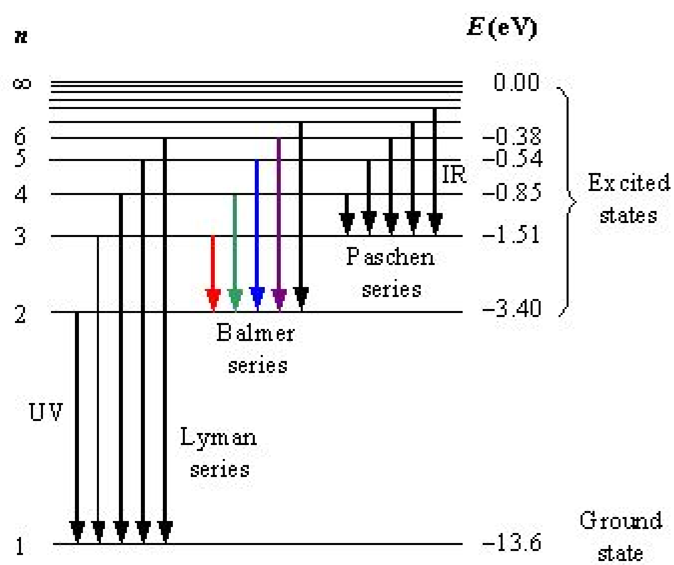
In a discharge tube, the electrons become temporarily excited and thus move to orbits that are farther from the nucleus. These transitions are only temporary, though, so the electrons return to their normal (or ground) states. When they do so, they give off energy. Since electrons can only undergo transitions between certain specific energy levels, only certain quantities (quantums) of energy are given off. Since the colour of light corresponds to the energy of possessed by a quanta (or photon), only certain coloured lines are observed.



The electron emits or absorbs the energy changing the orbits.

Bohr was able to create the first explanation for the periodic behaviour of the elements. The similarity in the chemical properties of the elements results from the filling of successive energy levels (2,8,8,18…). The end of Mendeleev's periods thus corresponds to a full energy level. Unfortunately, Bohr's theory only worked for atoms or ions containing 1 electron (H atom, He^+ ion).





Energy levels of the hydrogen atom with some of the transitions between them that give rise to the spectral lines indicated.

The Wave-Particle Nature of Light

Wave Properties of Light:

The wavelength and frequency of light are related by the speed of light.

$$\lambda = \frac{c}{f}$$

Where: c = speed of light
$= 3.00 \times 10^{8} m/s$

f = frequency
(cycles/s or hertz)

λ = wavelength (m)

Particle Properties of Light:

A particle of light is a photon. The <u>energy of a photon</u> is some multiple of an energy quantum called Planck's constant.

$$E = hf$$

Where: E = energy of a single photon (kJ)

f = frequency (cycles/s, Hz, s^{-1})

h = Planck's Constant
$= 6.63 \times 10^{-37} kJ \cdot s$

The <u>energy per mole</u> of photons can also be calculated:

$$E = hfN_A$$

Where: E = energy per mole of photons (kJ/mol)

N_A = Avogadro's Number ($6.022 \times 10^{23} \frac{particles}{mol}$)

CALCULATING TRANSITION ENERGY:

In the Bohr model, atoms can absorb or emit specific quantities (quanta) of energy which correspond to the wavelengths visible in the line spectra of elements.

Although it represented a great step, the Bohr model of the atom was only accurate for the one electron systems of the hydrogen atom. The energy of any orbit (n) of the hydrogen atom can be calculated by:

$$E_n = \frac{-1312}{n^2}$$ where E_n is the energy of any orbit (kJ/mol)

The transition energy (energy gained or lost as electrons change energy levels) can therefore be calculated. This is equal to the energy emitted by 1 mole of excited hydrogen atoms.

$$E_{transition} = E_{n_f} - E_{n_i} = hfN_A$$

$$\therefore E_{transition} = \left(\frac{-1312}{n_f^2}\right) - \left(\frac{-1312}{n_i^2}\right) = hfN_A$$

Sample Problem:
Calculate the frequency and wavelength of electromagnetic radiation emitted from a hydrogen atom if an electron moves from the third energy level to the first energy level.