The energy of light quanta of red color = 2.0 eV which is lower than the work function.

(3) Spectral lines in photo-emission →

→ Observations

↳ one see only certain lines (wavelengths) that are present in any photo emission spectrum.

(*) classically, it should have all ~~the~~ lines.

12th January 2024

Recall :

(*) Spectral lines in photo-emission :

Observations : → Only ~~as~~ certain wavelengths λ are present in any photo-emission spectrum.

(*) Johannes Rydberg (1888) :

Visible (to human eyes) spectral lines from hydrogen gas can be expressed as,

$$\boxed{\frac{1}{\lambda} = R_H \left(\frac{1}{2^2} - \frac{1}{n^2} \right)}$$

$n = 3, 4, 5$

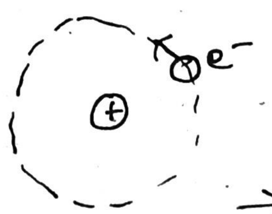
R_H → Rydberg constant

(*) Rutherford (1911) :

From ~~sca~~ scattering experiments

↳ An atom consists of concentrated +ve charge at the centre, surrounded by -vely charged electrons.

Eg : Hydrogen atom →



This is what was hypothesized.

→ A particle in a circular orbit is an accelerating particle

→ Maxwell : An accelerating charged particle will continuous radiate EM waves

So it will radiate energy and fall into proton. (matter of minutes)

Classical physics →

Force balance :

$$\boxed{\frac{m_e v^2}{r} = \frac{1}{4\pi\varepsilon_0} \cdot \frac{e^2}{r^2}}$$

So we can calculate energy of the electron →

$$\boxed{E = \frac{1}{2} m_e v^2 - \frac{1}{4\pi\varepsilon_0} \cdot \frac{e^2}{r}}$$

Eliminating using the force balance equation,

$$E = \frac{1}{2}\left(\frac{1}{4\pi\varepsilon_0} \frac{e^2}{r}\right) - \frac{1}{4\pi\varepsilon_0} \frac{e^2}{r}$$

$$\Rightarrow \boxed{E = -\frac{1}{2} \frac{e^2}{4\pi\varepsilon_0} \left(\frac{1}{r}\right)}$$

Classical bound state formula

(Total energy is half of potential)

As r is continuous, the electron could be anywhere.

Classically, the electron's orbit can have any r.

⇒ Any energy.

Then how do we explain only certain spectral lines?
This suggests that emitted light can have any wavelength → NOT consistent with expt.

Maxwell : An atom is not stable!

* Niels Bohr (1913):

Somehow electrons stay only on those orbits where angular momentum is quantized.

$$\boxed{m_e v r = n\hbar}\ , \hbar = \frac{h}{2\pi}\ , n = 1, 2, 3, 4, \dots$$

☐ You can backcalculate from this to get Rydberg formula.

from previous formula,

$$m_e^2 v^2 r^2 = \frac{m_e}{4\pi\varepsilon_0} e^2 r = n^2 \hbar^2$$

☐ How did he ~~arr~~ arrive at angular momenta as the thing that is quantized? ① Planck ($E = h\nu$) ② Total Energy

You can start with Rydberg and arrive at Bohr hypothesis ☒ Try

$$\Rightarrow \boxed{\frac{1}{r} = \frac{m_e}{n^2\hbar^2} \frac{e^2}{4\pi\varepsilon_0}}$$

$$\Rightarrow E_n = \frac{-m_e}{2\hbar^2}\left(\frac{e^2}{4\pi\varepsilon_0}\right)^2 \frac{1}{n^2}$$

* Now if an electron jumps from $n = 3, 4, 5, \dots$ to $n = 2$ energy levels..

Energy of light quanta,

$$\boxed{h\nu = E_n - E_2} \qquad \boxed{\nu = \frac{c}{\lambda}}$$

$$\Rightarrow \frac{1}{\lambda} = R_H\left[\frac{1}{2^2} - \frac{1}{n^2}\right]$$

Where, $\boxed{R_H = \frac{1}{hc}\left(\frac{m_e}{2\hbar^2}\right)\left(\frac{e^2}{4\pi\varepsilon_0}\right)^2}$

(*) $\boxed{E_7}$ (A3/Q1)

~~B~~ An experimental physicist ~~p~~ finds the wavelength of an EM wave from Hydrogen gas to be ~~100~~ 1010 nm. With a spectrometer having accuracy of 1%. Using, Bohr model, determine the initial and final energy levels of the ~~electron~~ corresponding electron. (given $E_1 = -13.6\,eV$)

(*) Key lessons from Planck, Einstein and Bohr →

1. Energy of the electromagnetic waves are quantized
2. An electron absorbs energy in quanta
3. An electron emits energy in quanta.

The problem lies in the determinism in classical physics

Under const force: $x = x_0 + u_0 t + \frac{1}{2}\left(\frac{F}{m}\right)t^2$
provided we supply x_0 and u_0.

(*) Is it possible to determine x_0 and u_0 ~~if~~ for any particle, even in principle?

→ All classical laws are 2nd order diff eqns, so two constants of integration. – are they possible to supply.