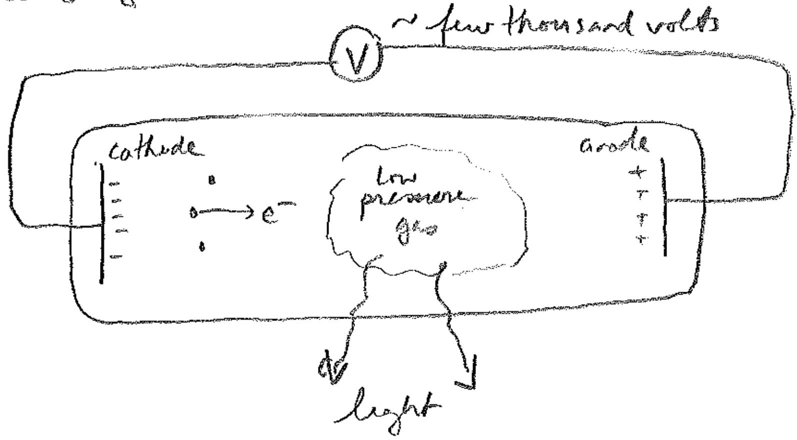
Discrete spectra emitted by excited low pressure gases

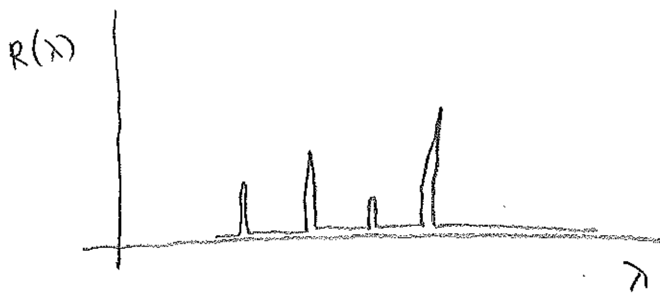
[How excite gases? Bombard them w/ electrons]

[DEMO: show them one.]

Gas discharge tube

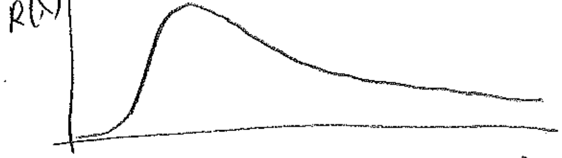


[apply large voltage between electrodes. strong $\vec{E}$ field strips e^- from cathode which accel. to anode collide w/ gas atoms & xfer kinetic energy; excited gas atoms then release energy to radiation]



Each type of gas emits light w/ certain characteristic wavelengths (called its spectrum)

as opposed to continuous thermal spectrum emitted by hot solids, liquids, & high pressure gases



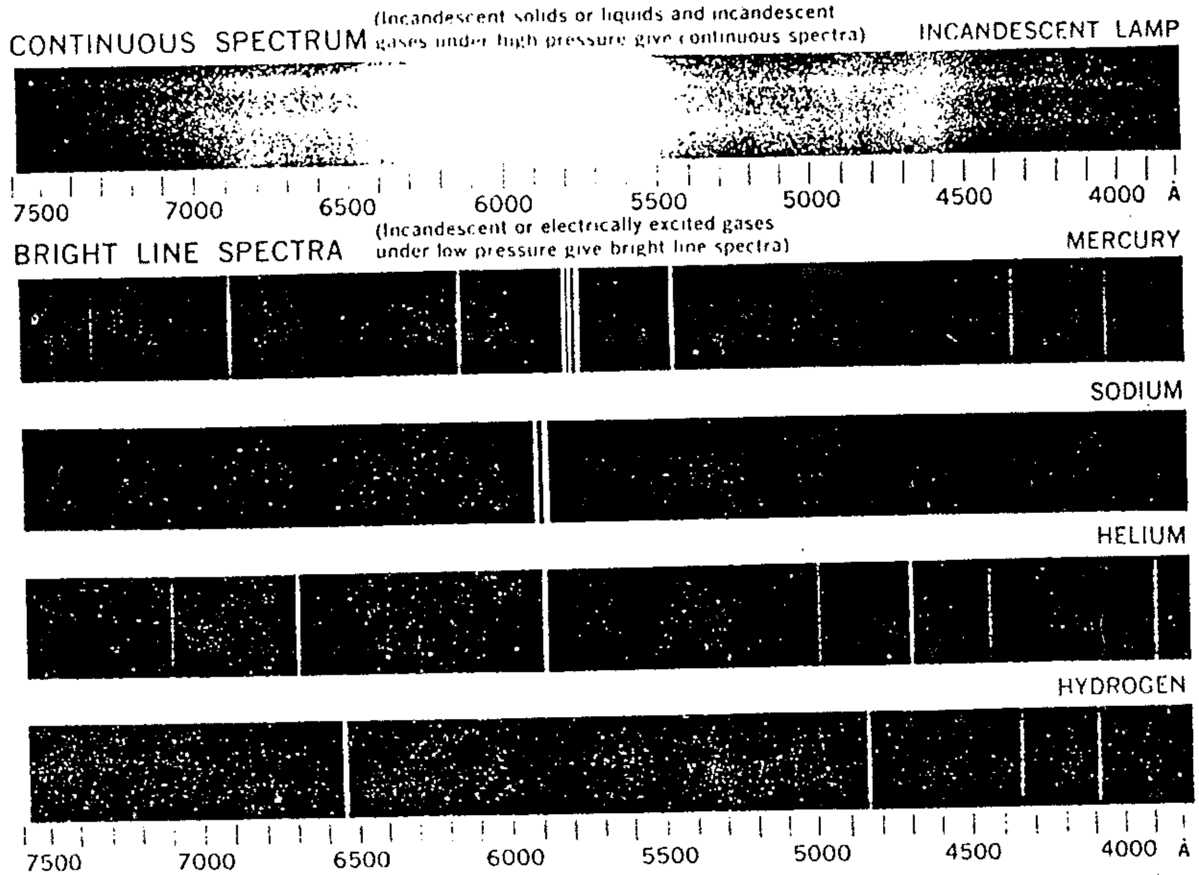
[show transparency]

[characteristic spectrum acts as a fingerprint; used to identify]

V1.5

Better to do actual demo using ~~tools~~ spectra tubes

EMISSION SPECTRA



Adapted from the SPECTRUM CHART, Welch Scientific Company

[DEMO: hand out diffraction gratings
close shutters, turn off lights

show:

- thermal spectrum from incandescent bulb (again)
- mercury [used in fluorescent lamps]
- [talk about sodium yellow light used in parking lots]
- helium
- neon [used in neon signs]
- hydrogen]

Hydrogen Balmer series

(red)	656 nm	H_α
(aqua)	486 nm	H_β
(blue)	434 nm	H_γ
(violet)	410 nm	H_δ

[story of Balmer, swiss math teacher]

Balmer guessed a formula for these wavelengths

$$\lambda = (\text{const}) \left(\frac{4n^2}{n^2-4} \right)$$

for integers n

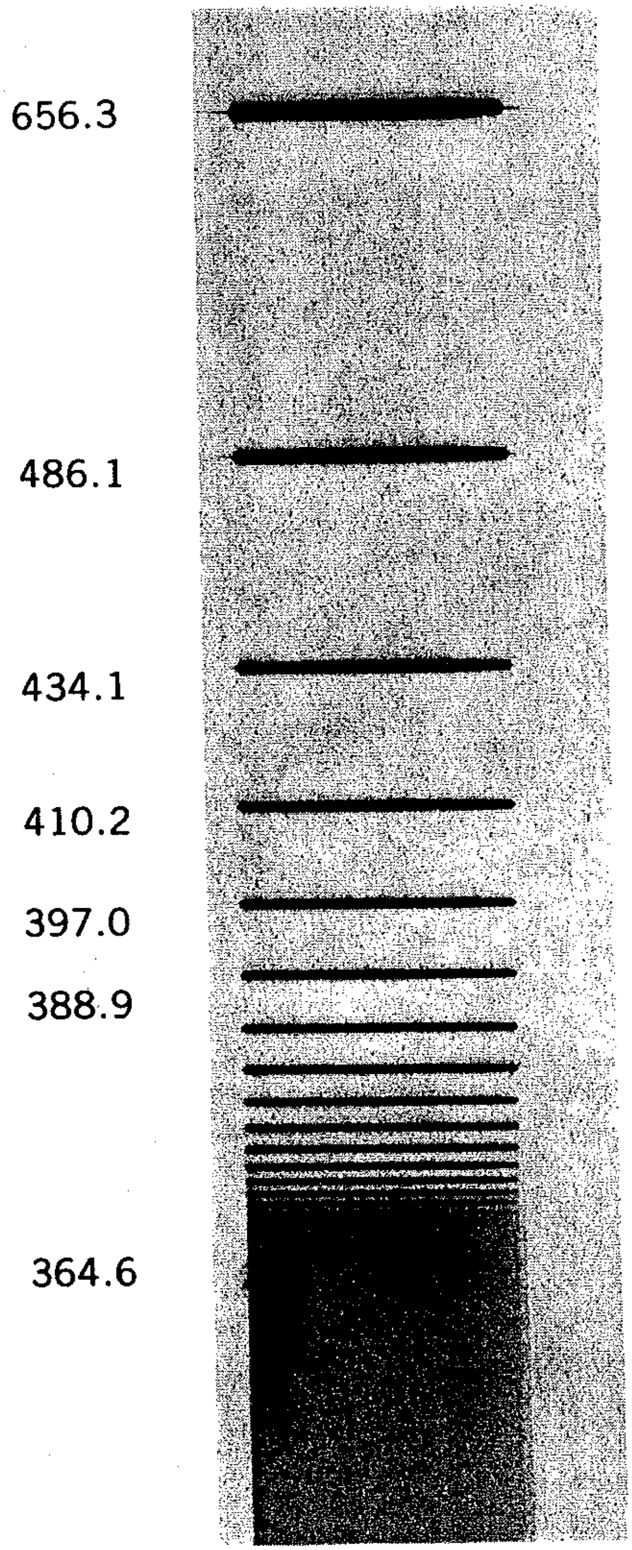
$$\begin{cases} H_\alpha : n=3 \\ H_\beta \quad n=4 \\ H_\gamma \quad n=5 \\ H_\delta \quad n=6 \end{cases}$$

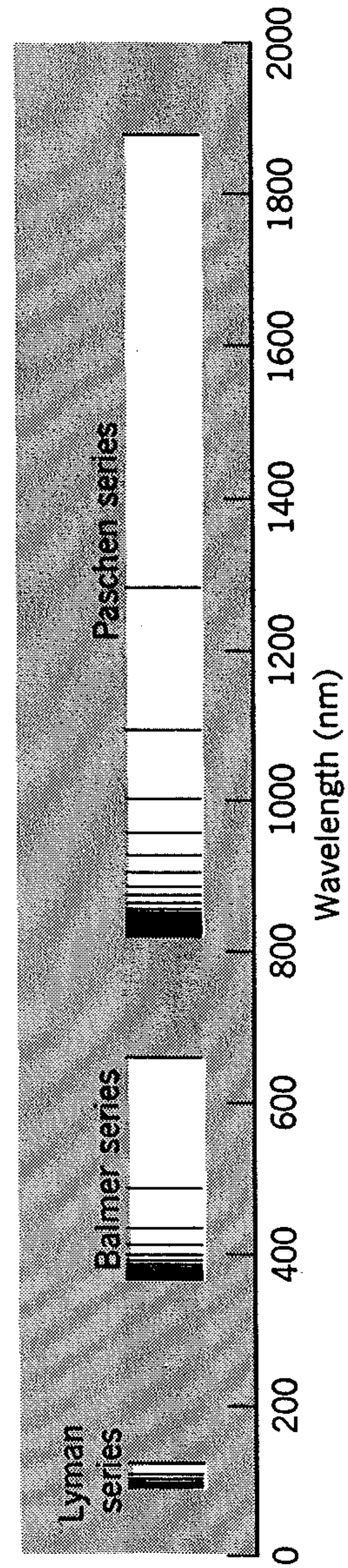
$\Rightarrow$ (const) $= 91.15\,nm$

[What happens for $n=7$ & above?
Predictions for new lines w/ λ in the ultraviolet
Can't see them w/ eye but can photograph them
$\longrightarrow$ show transparency.

They are there! Process of science:
make predictions, then test them]

V3.5





HRK, ch 51, fig 1 and fig 2

$$\frac{1}{\lambda} = \frac{1}{(\text{const})}\left(\frac{n^2-4}{4n^2}\right) = \frac{1}{(\text{const})}\left(\frac{1}{4} - \frac{1}{n^2}\right)$$

Rydberg proposed a more general formula

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

R = Rydberg const $= 1.097 \times 10^7\ \text{m}^{-1}$

m, n different integers w/ $m < n$

$m = 1 \Rightarrow \frac{1}{\lambda} = R\left(1 - \frac{1}{n^2}\right)$, $n = 2, 3, \ldots$ Lyman series (UV)

[see transparency]

$m = 2 \Rightarrow$ Balmer series

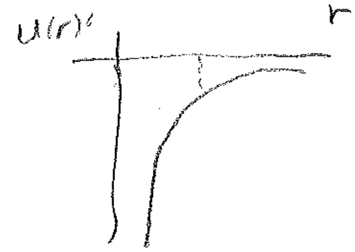
$m = 3 \Rightarrow \frac{1}{\lambda} = R\left(\frac{1}{9} - \frac{1}{n^2}\right)$, $n = 4, 5, \ldots$ Paschen series (IR)

Skip this. It is too confusing & does not help at this stage, because too many unanswered questions

Hydrogen atom = proton + electron bound by Coulomb force

Potential energy

$$U(r) = \frac{Kq_1q_2}{r} = -\frac{Ke^2}{r}$$



just explain in words

[electron in orbit ⇒ acceleration ⇒ radiation

frequency of radiation ⇒ frequency of orbital motion

faster for small r.

At Bohr radius $r \sim 0.05\text{nm}$, $\lambda \sim 40\text{nm}$ (UV).]

$$\left[\begin{array}{l} \frac{Ke^2}{r^2} = m\omega^2 r \\ \omega = \sqrt{\frac{Ke^2}{mr^3}} \\ \lambda = \frac{c}{f} = 2\pi c\sqrt{\frac{mr^3}{Ke^2}} \end{array}\right]$$

Two problems of the classical model. As radiates energy, U ↓

so r ↓, & f gradually increases.

① expect continuous spectrum

② unstable in abt 15 ps

[but we know Bohr model gnd st stable. How?]

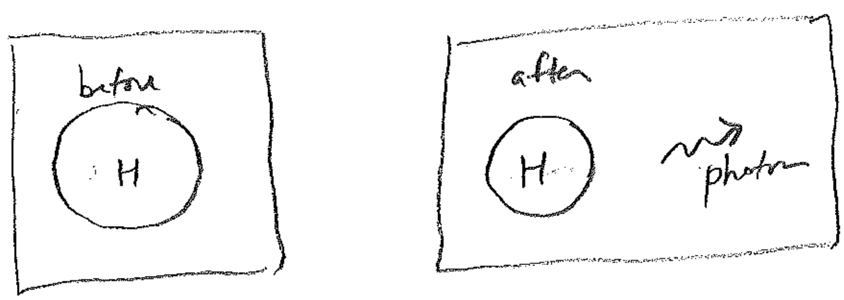
Quantum hypothesis: spectral lines w/ wavelength λ correspond to photons of energy

$$E_{photon} = hf = \frac{hc}{\lambda} = hcR\left(\frac{1}{m^2} - \frac{1}{n^2}\right)$$

$$hcR = 13.6\ eV$$

If these photons are emitted by excited hydrogen atoms energy conservation says

$$E_{photon} = E_H^{before} - E_H^{after}$$

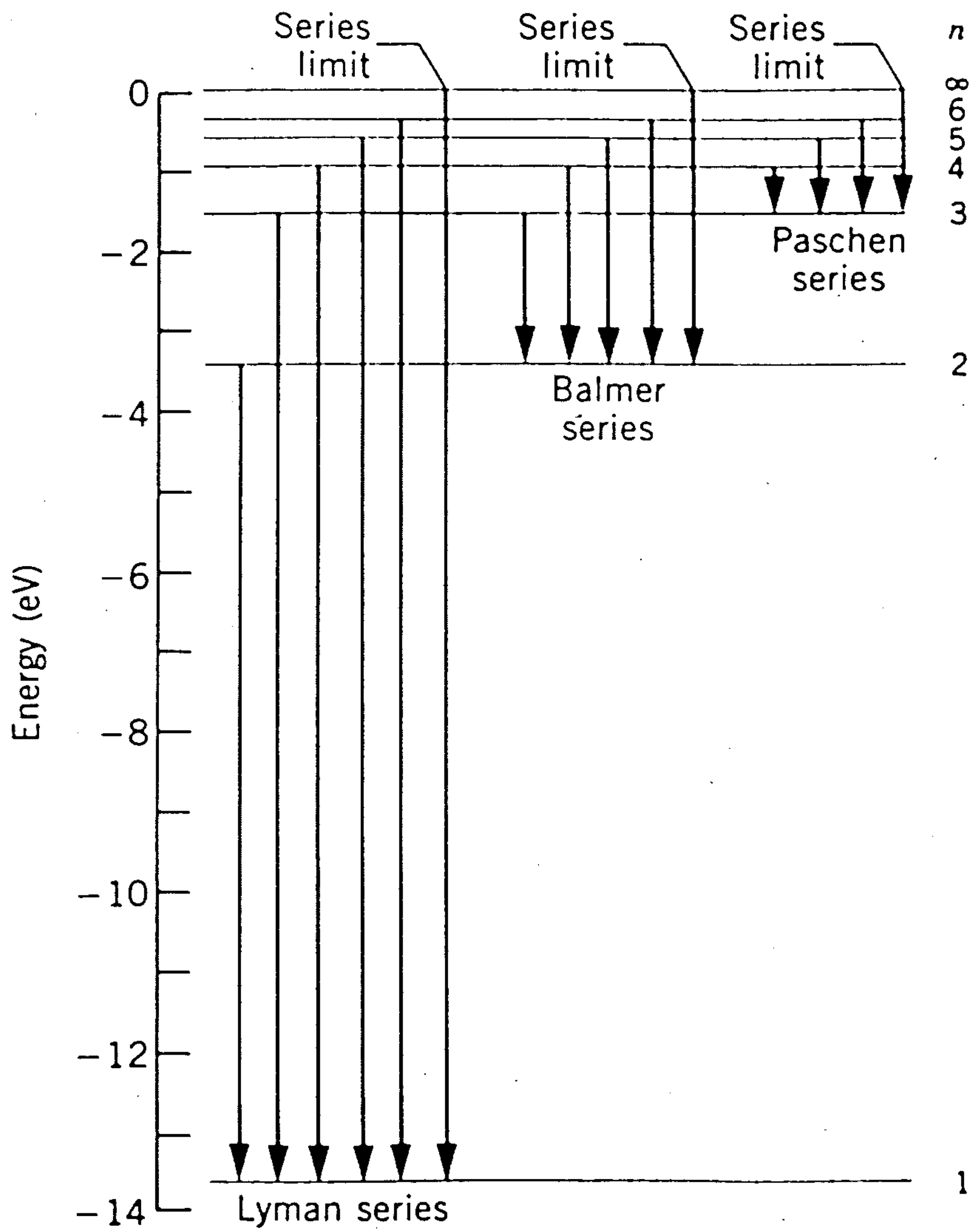


This suggests that hydrogen atoms have discrete energy levels

$$E_n = -\frac{hcR}{n^2}, \quad n = 1, 2, 3$$

$$E_{photon} = E_n - E_m$$

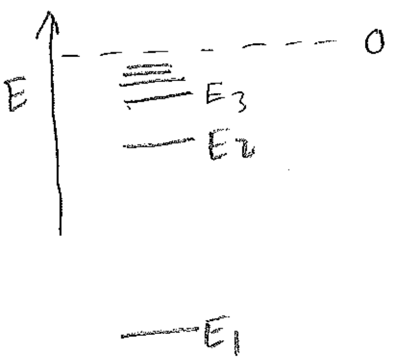
n = initial state
m = final state



Series limit
Series limit
Series limit
n
∞
6
5
4
3
2
1
0
−2
−4
−6
−8
−10
−12
−14
Energy (eV)
Paschen series
Balmer series
Lyman series

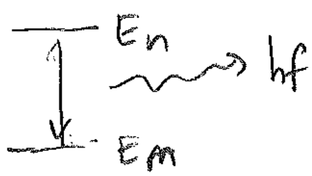
The lowest energy level of hydrogen ($n=1$) is called the ground state: $E_1 = -hcR = -13.6\,\text{eV}$.

Higher energy levels ($n \geq 2$) are called excited states and range from $E_2 = -\frac{13.6\,\text{eV}}{4} \approx -3.4\,\text{eV}$ to just below zero.



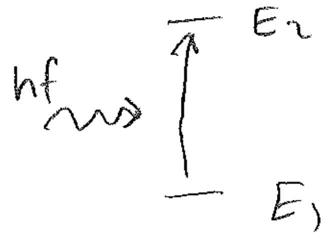
If $E \geq 0$, atom is ionized and the electron is free.

In a transition from level n to level m, an atom emits a photon whose energy is the <u>difference</u> of hydrogen energy levels



$$hf = E_n - E_m$$

An atom in the ground state cannot emit light
but it can absorb a photon
whose energy is sufficient to reach an excited state



Since $E_2 - E_1 \approx 10\,eV$

$$hf \gtrsim 10\,eV$$

$$\Rightarrow \lambda \lesssim 100\,nm \quad (UV)$$

Hydrogen gas is transpared to visible light
but can absorb ultraviolet light

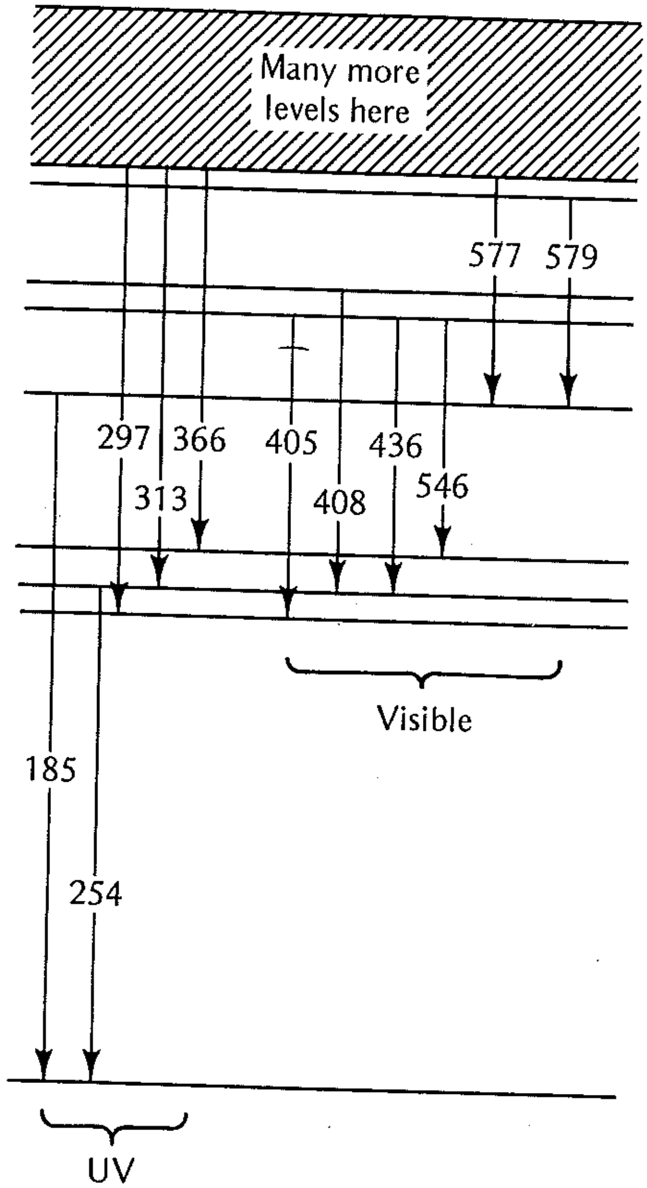
Similarly of most other gases.

Other gases also have discrete energy levels.
There is no simple formula for them.
but in all case

$$E_{photon} = E^{before}_{atom} - E^{after}_{atom}$$

[see Hg]

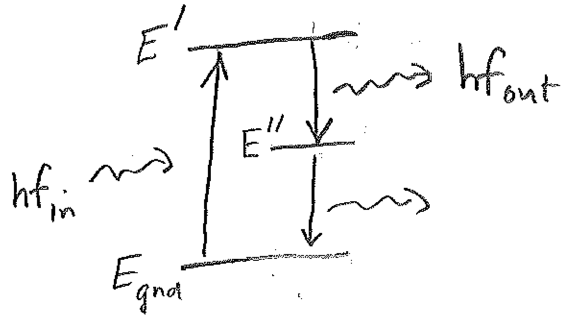
V7-5



Mercury vapor energy levels

Fluorescence

process by which an atom in the ground state absorbs a photon of wavelength λ_{in} and subsequently emits a photon of wavelength λ_{out}



[which is larger: λ_{in} or λ_{out}?

$E' - E_{gnd} > E' - E''$

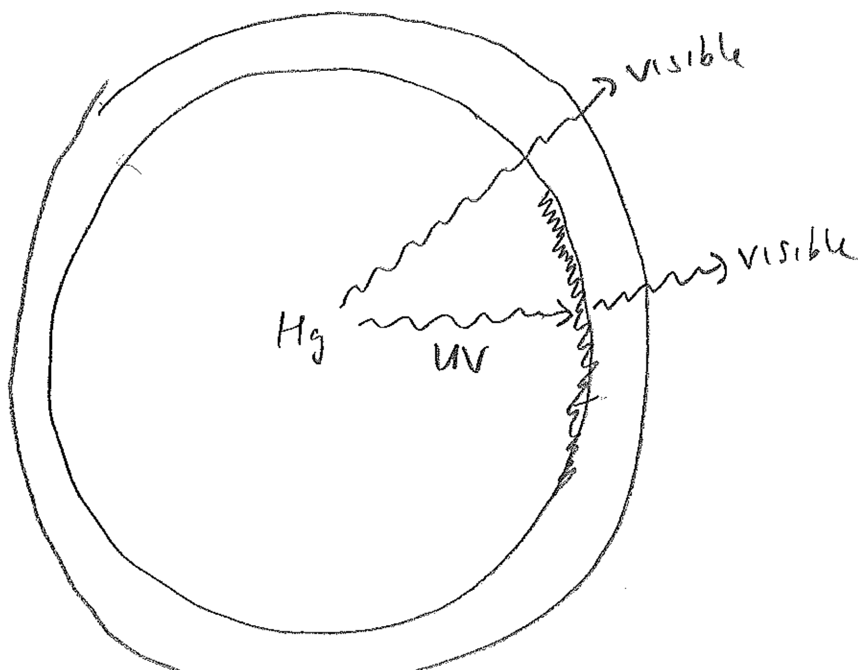
$hf_{in} > hf_{out}$

$\lambda_{in} < \lambda_{out}$]

In this process, UV light can be converted to visible light

Fluorescent lamp

excited Hg atoms emit both visible (bluish)
and UV light.
The UV is absorbed by the glass and
by a fluorescent coating on inside



The coating re-emits light in the visible (white)

[DEMO: half-coated fluorescent;
UV lamp on broken tube;
quinine water]