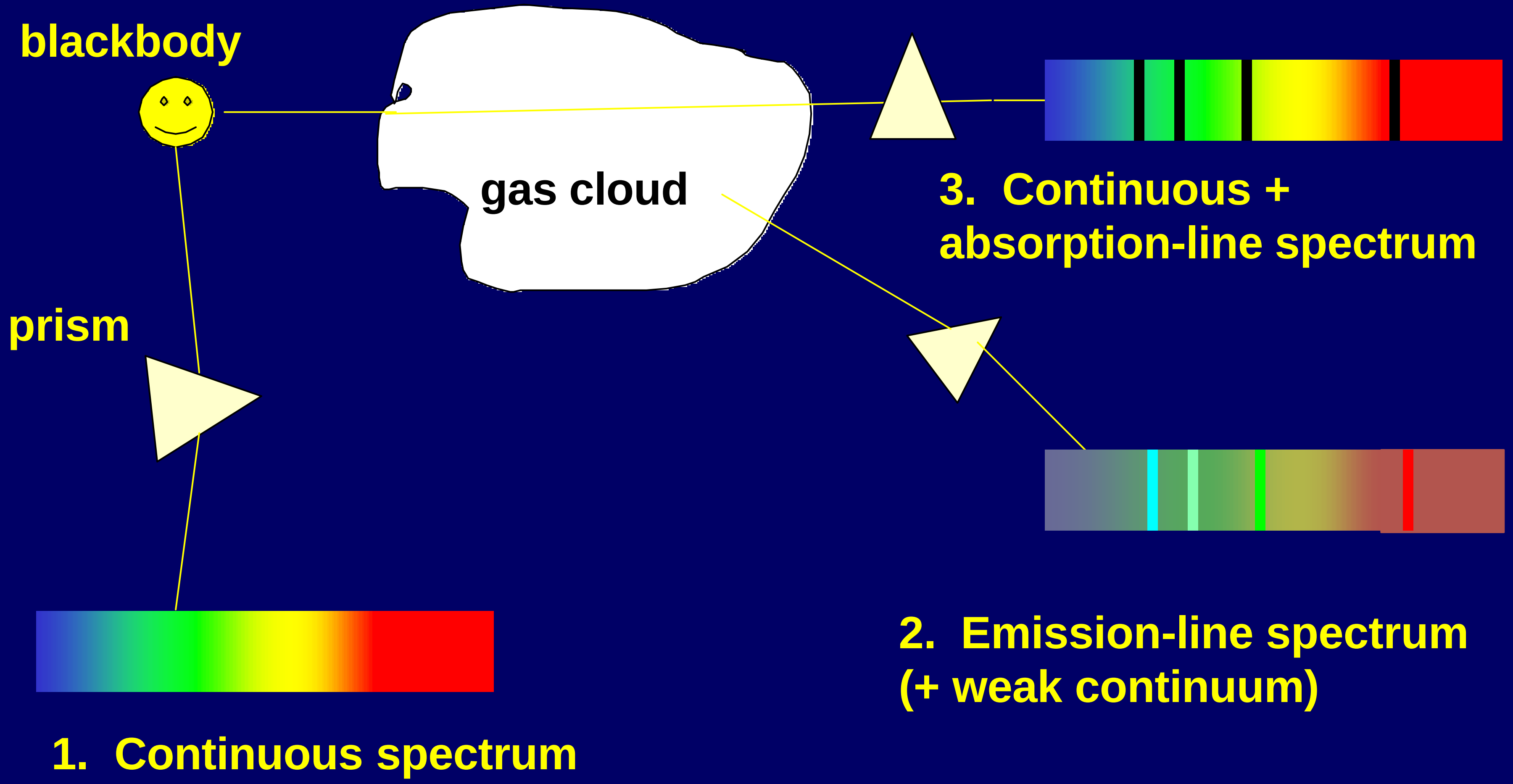
Spectral Analysis

– light is dispersed into a spectrum using a diffraction grating (or prism) in a spectrograph

– Fraunhofer (1815) first extensive study of the Sun - identified about 600 dark lines

– Fraunhofer lines - strongest named A,B,...

Kirchhoff's laws
3 types of spectra



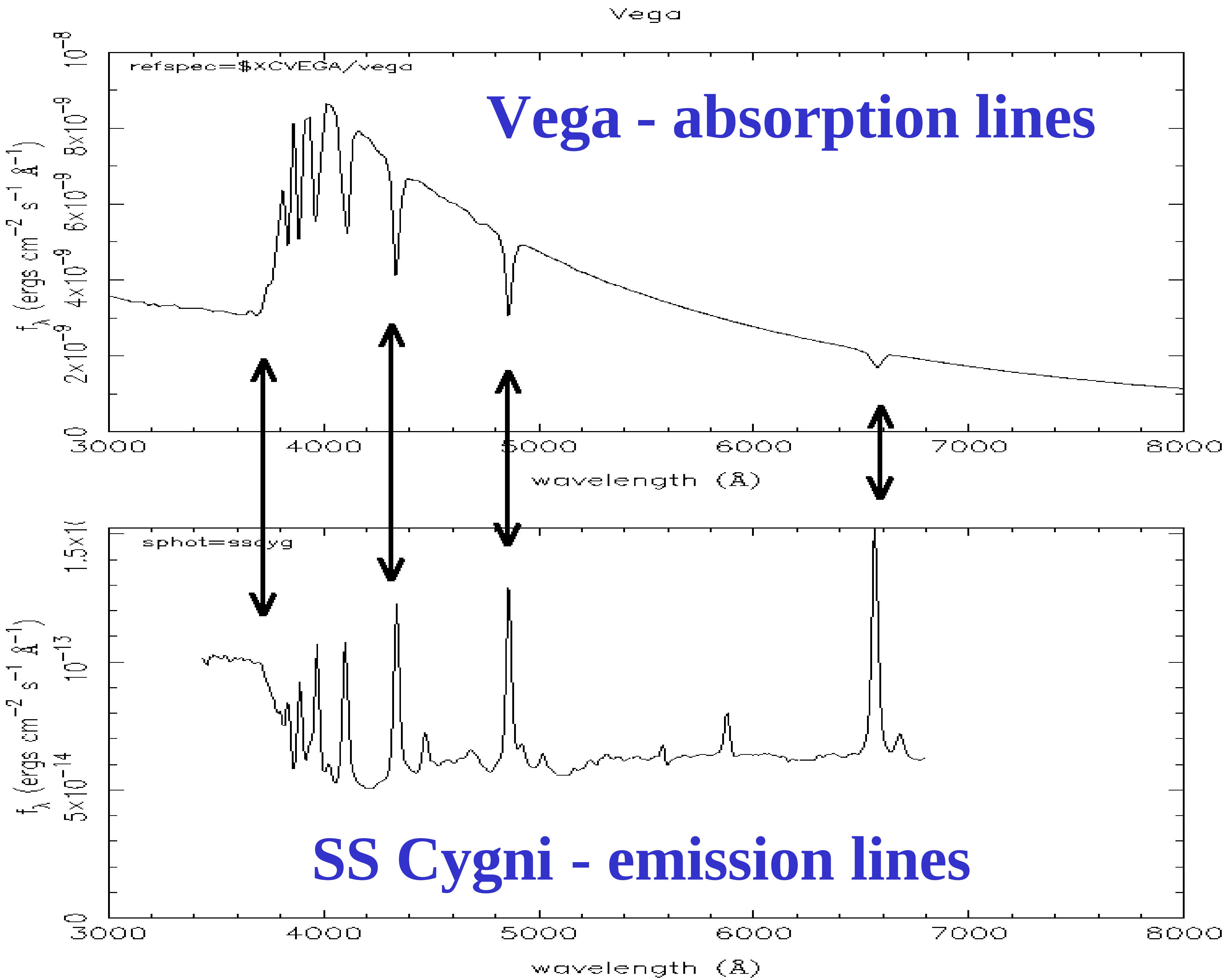
Kirchhoff's laws

A *hot opaque body*, e.g. a hot dense gas, produces a *continuous spectrum*, e.g. a "black-body" spectrum.

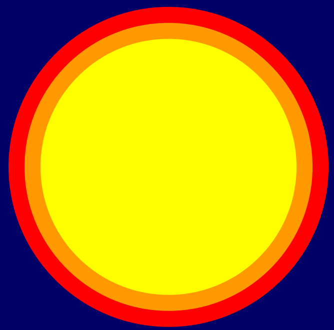
A *hot transparent gas* emits an *emission-line spectrum* , bright spectral lines, sometimes with a faint continuous spectrum.

A *cool transparent gas* in front of a continuous spectrum source produces an *absorption-line spectrum* - a series of dark spectral lines.

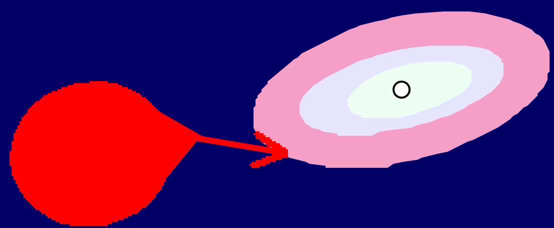
(Kirchoff and Bunsen laboratory experiments, 1850s)


Vega
refspec=$XCVEGA/vega
Vega - absorption lines
wavelength (Å)
sphot=sscyg
SS Cygni - emission lines
wavelength (Å)

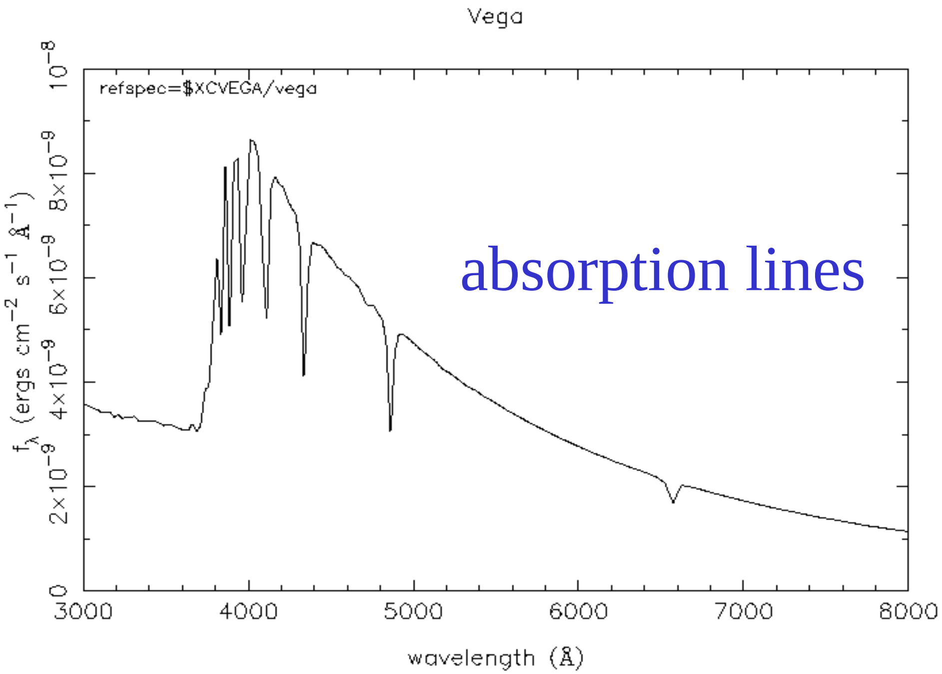
Kirchoff's Laws

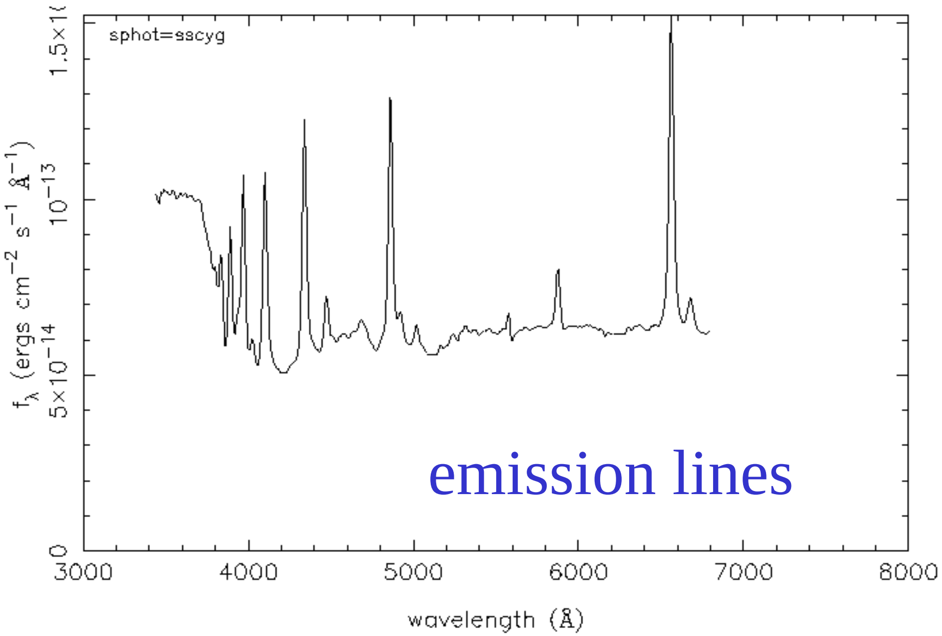


hot opaque interior,
cool transparent atmosphere



hot transparent gas
(accretion disc)





Spectral Fingerprints

– *E*ach element / molecule absorbs and emits only certain specific wavelengths of light.

– Spectral lines are diagnostic of the chemical composition and

– physical conditions (temperature, pressure)

Atoms and Ions

- atoms: **nucleus (protons + neutrons) + electrons**

	mass	charge
proton	**1**	**+1**
neutron	**1**	**0**
electron	**1/1836**	**-1**

atom - neutral : equal numbers of protons (+ve) and electrons (-ve)

ion - ionised : electrons removed, positive net charge

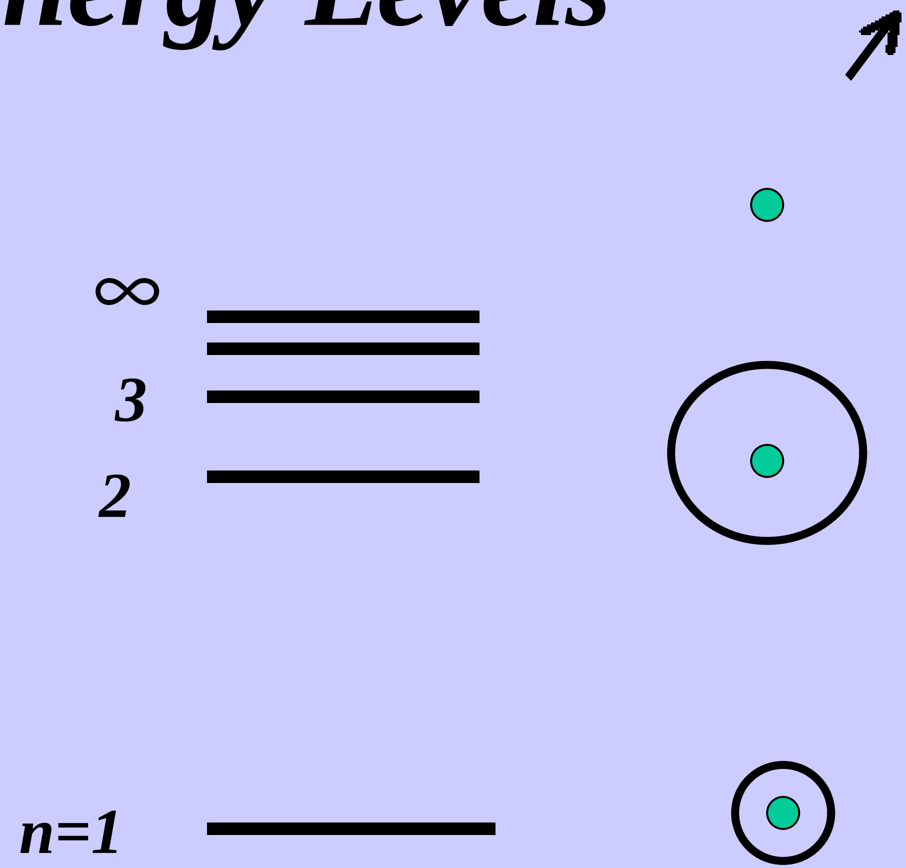
Atoms and Ions

- Examples:
  - Hydrogen (H)   (1 proton) + 1 electron
  - H II   (1p)   charge +1
  - Helium (He)   (2p + 2n)   + 2 e-   0
  - He II   (2p + 2n)   + 1 e-   +1
  - He III (2p + 2n)   +2
  - Oxygen (O)   (8p + 8n)   + 8 e-   0
  - OIII   (8p + 8n)   + 6 e-   +2
  - OVI   (8p + 8n)   + 3 e-   +5

Atomic Energy Levels

e.g. Hydrogen:

$$E_n = -\frac{e^2}{r_n} = -\frac{I}{n^2}$$



$$E_1 = -13.6\,\mathrm{eV}$$

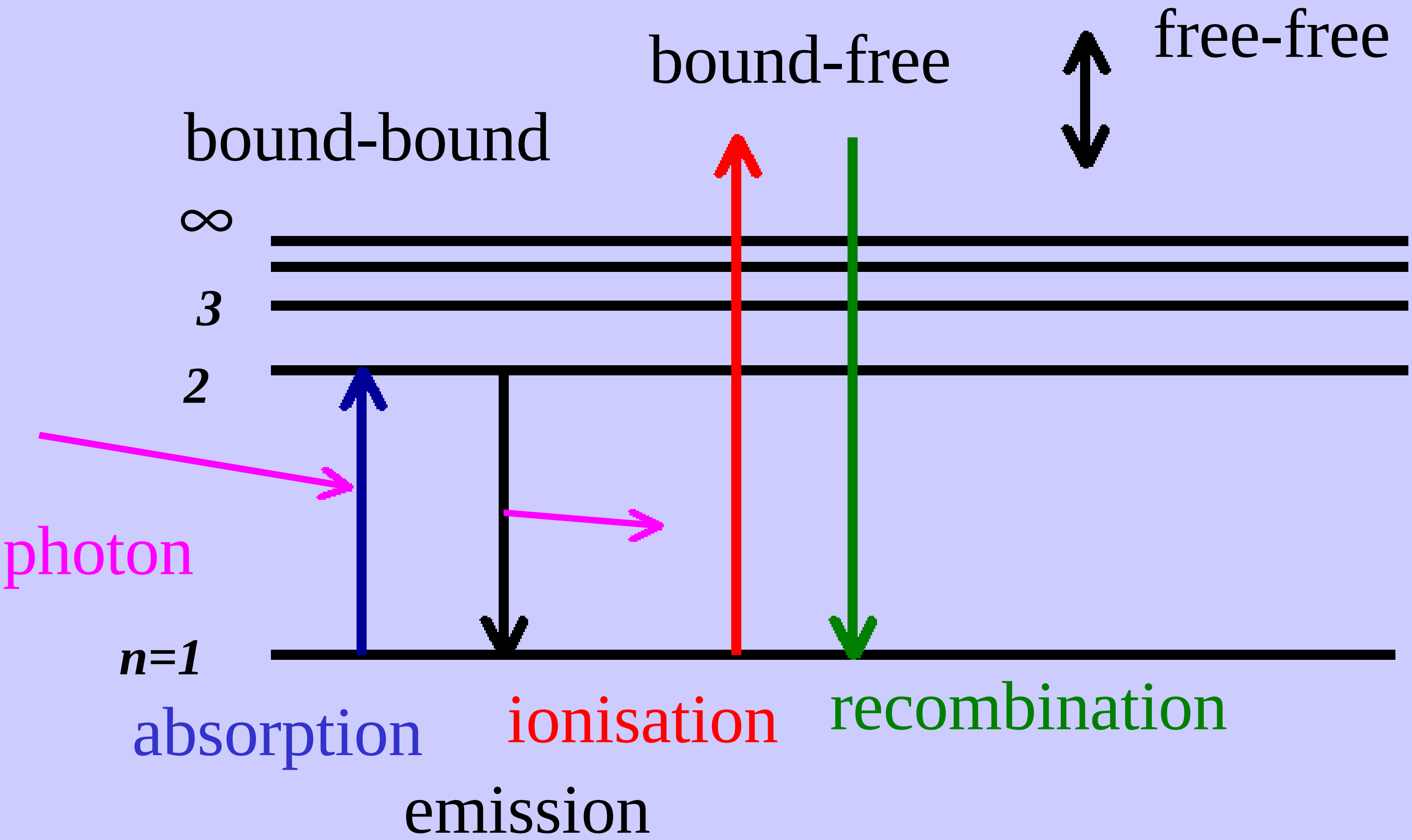
$$E_\infty = 0$$

$$I = 13.6\,\mathrm{eV} = \mathrm{Ionisation\ \ Potential}$$

$$e = \mathrm{proton\ charge}$$

$$r = \mathrm{size\ of\ \ electron\ orbit}$$

Atomic Transitions



- **Energy change associated with each transition is**

$$E = h\,f = \frac{hc}{l}$$

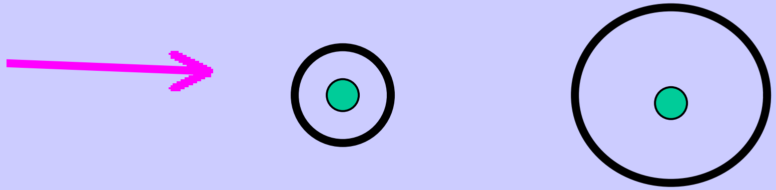
(*h* is Planck's constant)

higher *E* ® higher frequency (shorter wavelength) photon is absorbed/emitted

- **energy changes are very small**
  - **measured in ELECTRON VOLTS (eV)**
    1 eV = 1.602 ´ 10^{-19} J

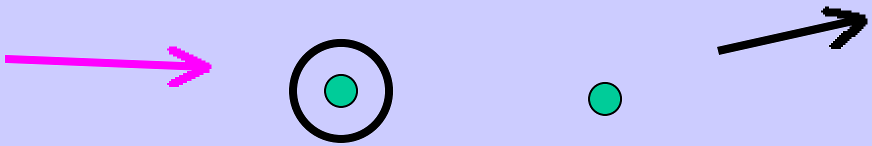
Energy Conservation

bound-bound



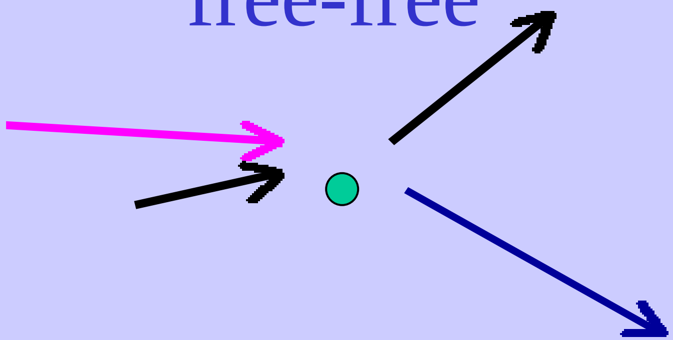
$$\frac{h\,c}{l} + E_1 = E_2$$

bound-free



$$\frac{h\,c}{l} + E_1 = E_\infty + \frac{1}{2} m v^2$$

free-free



$$\frac{h\,c}{l_1} + \frac{1}{2} m v_1^2 = \frac{h\,c}{l_2} + \frac{1}{2} m v_2^2$$

- Example:
  - H atom
    - (see handout: energy level diagram)
  - Energy difference between the ground state (n=1) and the first excited state (n=2) :

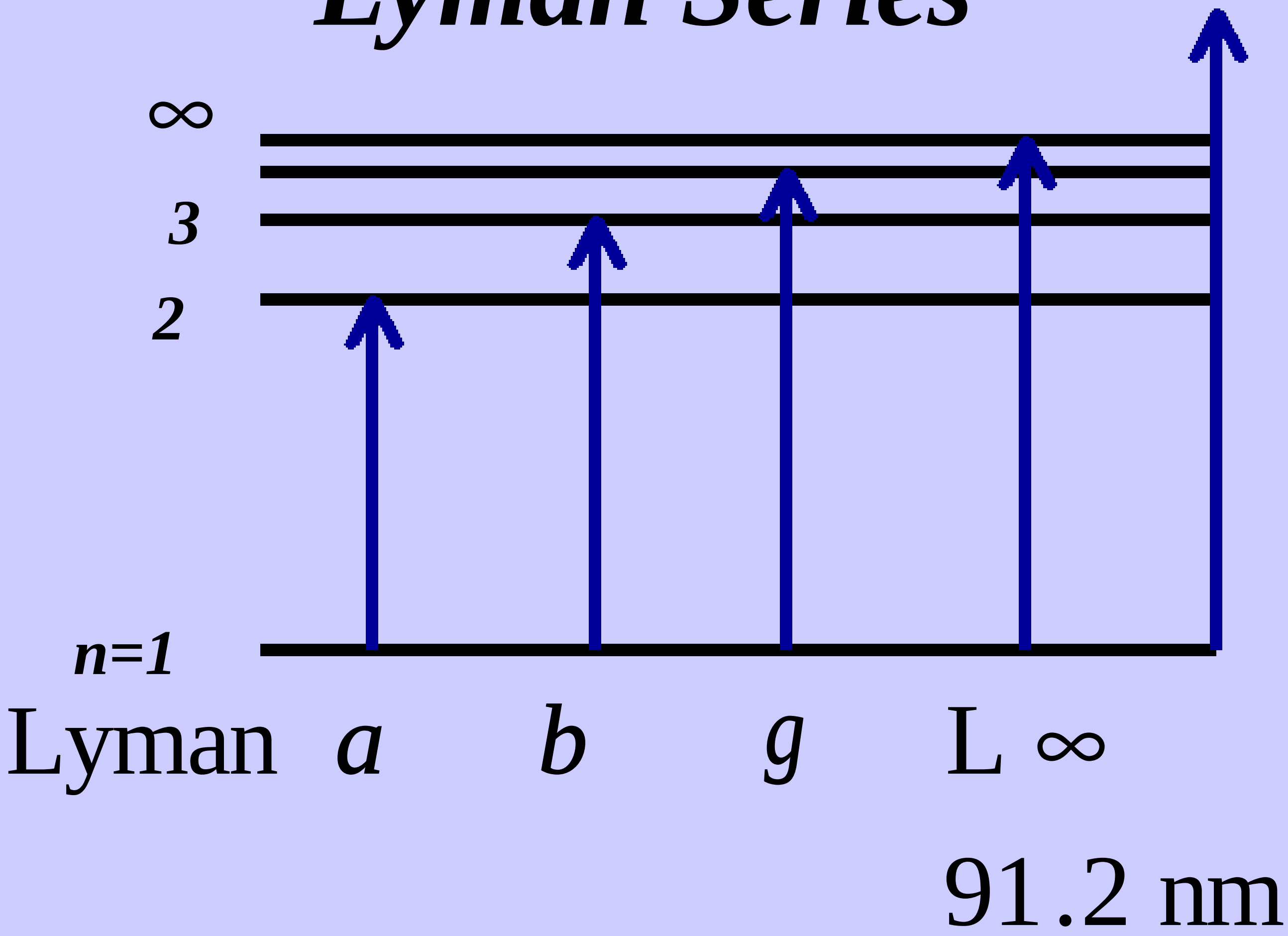
$$E = E_2 - E_1 = I \times \left(\frac{1}{1^2} - \frac{1}{2^2} \right) = (13.6 \text{ eV})(0.75) = 10.2 \text{ eV}$$

$$l = \frac{h\,c}{E} = \frac{(6.626 \times 10^{-34} \text{ J s})(3 \times 10^{8} \text{ m/s})(10^{9} \text{ nm/m})}{(10.2 \text{ eV})(1.602 \times 10^{-19} \text{ J/eV})}$$

$$= 121.6 \text{ nm}$$ in UV part of spectrum

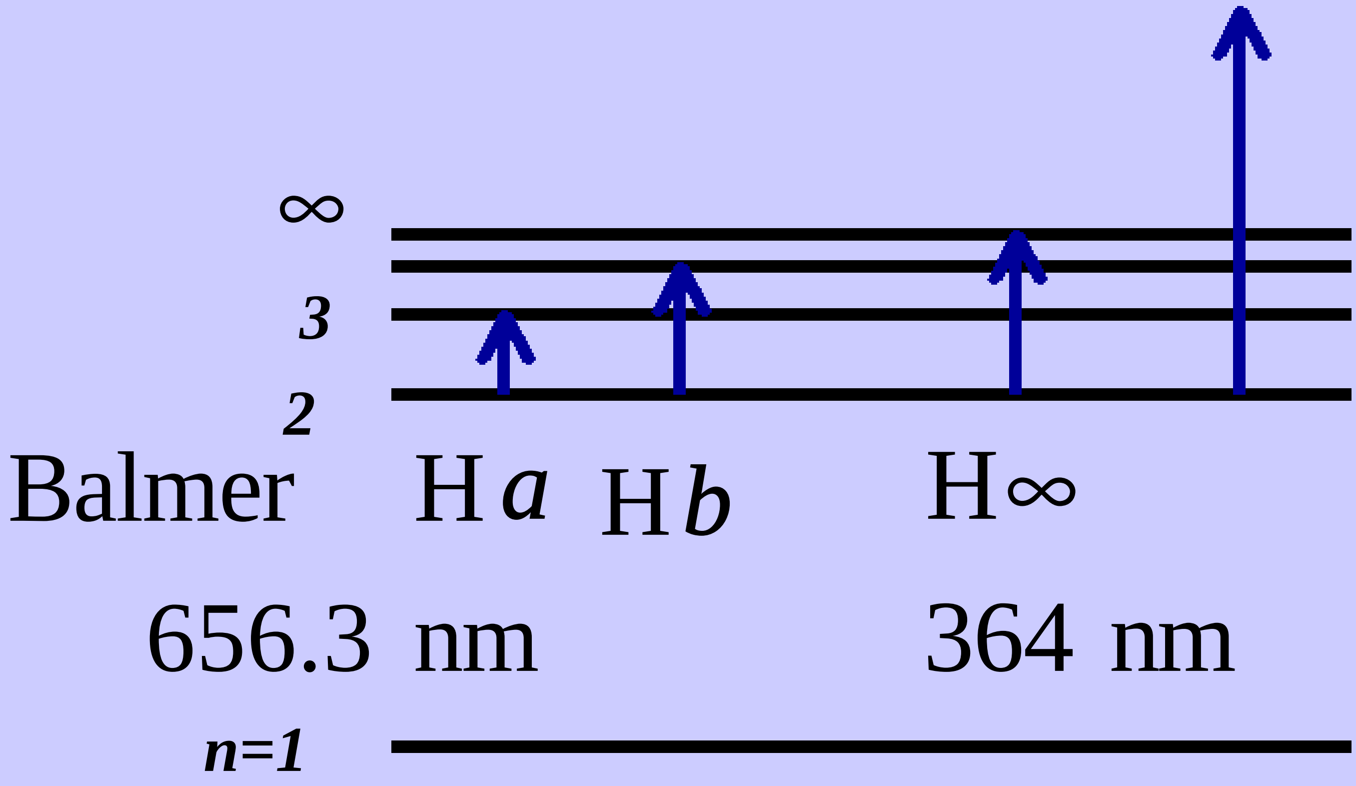
Lyman α line (Lα)

Lyman Series



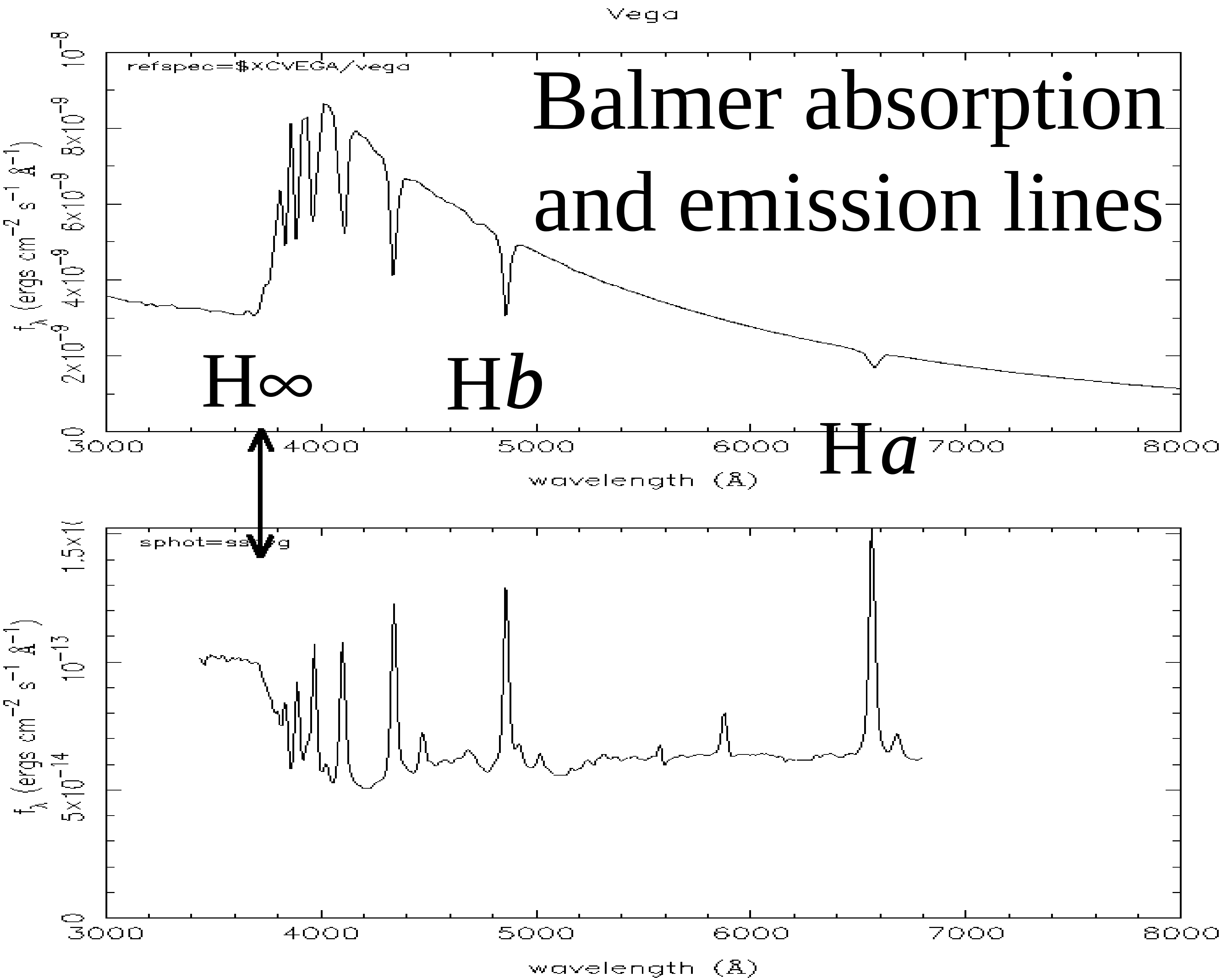
Lyman continuum photons ionise H from the ground state.

Balmer Series



Balmer continuum photons ionise from n=2

Balmer absorption and emission lines



Rydberg formula

$$\frac{1}{\lambda} = R\left(\frac{1}{n_\ell^2} - \frac{1}{n_u^2}\right)$$

- Rydberg constant $R = 1097 \times 10^7\ \text{m}^{-1}$ $\quad \frac{1}{R} = 91.2\ \text{nm}$

n_ℓ - principal quantum number of lower level
n_u - " " " of upper level

LYMAN series	$n_\ell = 1$	$n_u = 2, 3, 4 \ldots$	UV
BALMER	2	3, 4, 5 …	visible
PASCHEN	3	4, 5 …	near IR
BRACKETT	4	5, 6 …	IR
PFUND	5	6, 7 …	IR

- Hydrogen is simplest.
- Multi-electron atoms have more complicated energy-levels.
- ions with 1 electron are like Hydrogen but with larger Ionisation Potential due to higher charge of nucleus