BOHR MODEL OF HYDROGEN ATOM

A. LINE SPECTRA

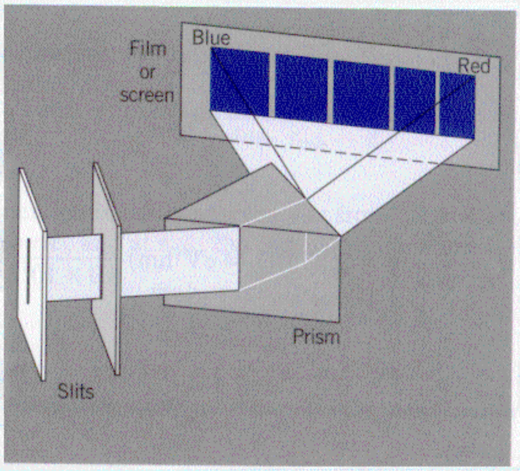
It has been known for a long time that salts of certain metals ~~make~~ change the color of a flame

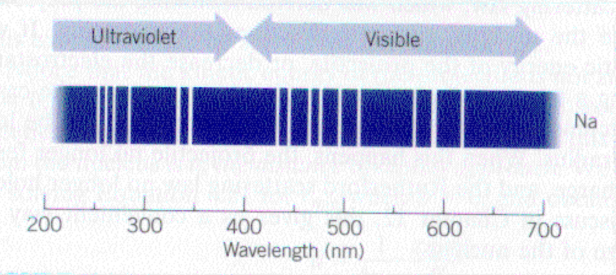
Eg.: Sodium salts make the flame bright yellow
Barium (Ba) " " " " green
Strontium (Sr) " " " " red
Lithium (Li) " " " " also red

~~[illegible]~~ This effect has been used in fireworks since very old days, and is being used today (July 4th fireworks!)

In the mid-XIX century scientist started to study these colors in greater details — by applying the so-called "spectral analysis", in which light is split into elementary color components (using a prism, or grating).

The experimental set-up in those early experiments is shown schematically below:



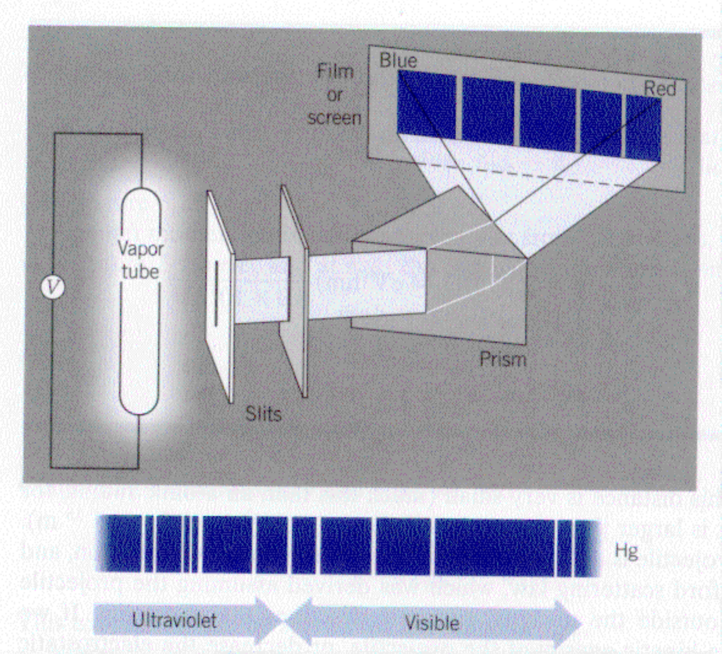


The results of those experiments showed that the salts produced characteristic spectra, consisting of narrow "lines" of pure colors

On a black-and white photograph, those "line spectra" looked very much like the <u>barcodes</u> used currently to mark commercial products.

Soon, it was discovered that each element investigated produced a unique, distinct "barcode". Chemists started using spectral analysis for detecting trace elements (also, crime investigators – e.g., a favorite poison used by murderers in XIX century was arsenic oxide As_2O_3, and spectral analysis made it possible to prove that the victim was poissoned by that substance).

Later, it was discovered that Light is emitted when an electric discharge occurs in a tube containing a vapor of an element, or a dilute gas like H, O, He, N, etc.



This new technique led to further progress in spectral analysis. However, it was still a total mystery WHY different elements produced different patterns of lines in the spectra.

The first step toward solving this puzzle was made by Johannes Balmer, a Swiss schoolteacher. He studied the wavelengths emitted by hydrogen gas in a discharge tube, and he discovered that all wavelengths he observed could be very accurately calculated from the formula:

$$\lambda = (364.5\,\text{nm}) \times \left(\frac{n^2}{n^2 - 4}\right)$$

where n takes an integer value $(n > 2)$.

Balmer studied only the visible wavelength range. However, his discovery stimulated other scientists to examine the infrared and ultraviolet ranges.

Soon it was found that there are more such "series" as Balmer discovered:

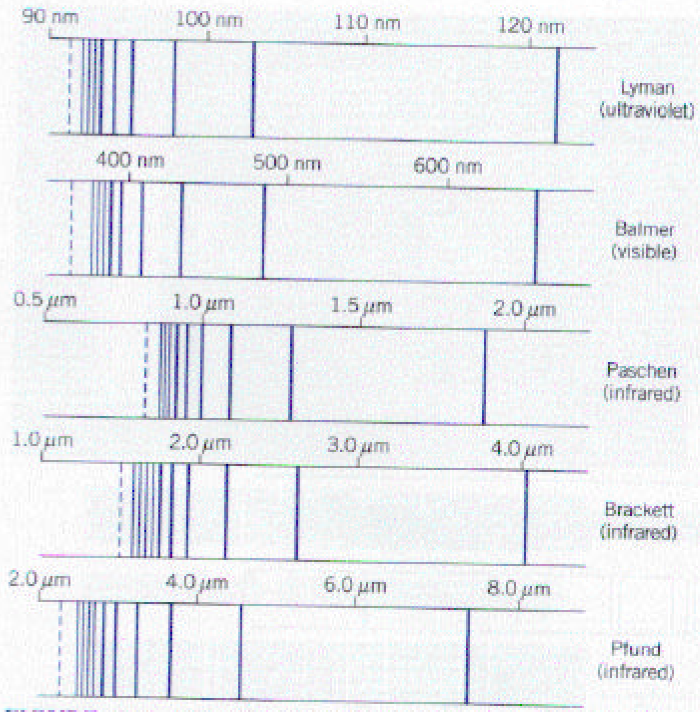


FIGURE 6.18 Emission and absorption spectral series of hydrogen. Note the regularities in the spacing of the spectral lines. The lines get closer together as the limit of each series (dashed line) is approached. Only the Lyman series appears in the absorption spectrum; all series are present in the emission spectrum.

The wavelengths in all those series were found all to satisfy the following equation:

$$\lambda = \lambda_{min} \frac{n^2}{n^2 - n_0^2} \qquad (n > n_0)$$

where λ_{min} and n_0 were different for each series:

λ_{min} and n_0 for different spectral series of Hydrogen atom:

	λ_{min}[nm]	n_0	range
Lyman series	91.13	1	ultraviolet
Balmer "	364.5	2	visible
Pashen "	820.1	3	infrared
Brackett "	1459	4	"
Pfund "	2280	5	"
Humphreys "	3282	6	"

The work of those gentlemen showed that there is a clear regularity in the emitted wavelengths. Still, the reason why it was so remained a puzzle.

Now, NIELS BOHR arrives at the stage...

Inspired by the Rutherford's experiment, Bohr considered various ideas of how the atom is built. And one idea that struck him turned out to be a great success!