5. Experimental Examples

This instrument can be used to measure emission spectra. Emission spectra are produced by atoms that are in high temperature states and experience collisions by charged particles. The atoms in different states emit different spectrum. In the atomic state, the spectra are bright lines, such as Sodium lamp, Mercury lamp, and Hydrogen-Deuterium lamp; in the molecular state, the spectra are band spectrum, such as Nitrogen discharge lamp. In the red-hot state, the spectra are continuum spectra, such as Tungsten lamp.

Related to the specific structure of atomic energy levels of different elements, the spectrum of each element has its own characteristic, just like the fingerprint of a human, which can be used to identify the element.

Note: Lamps for these experiments are not included.

5.1 Observe the Spectra of a Sodium Lamp and a Mercury Lamp

Objective

Learn to measure the wavelength of an emission source and use a spectrometer.

Learn to recognize the characteristic spectra of Sodium lamp and Mercury lamp.

Principle

Sodium lamp and Mercury lamp are widely used in college physics experiments. Sodium lamp is a vapour discharge glass tube filled with Sodium and Neon gas. First, the Neon discharges and emits red light after electrification, then the Sodium liquid droplet will be vaporized and emits light instead of the Neon. The radiation is nearly monochromatically yellow in colour and is stable after a few minutes. The *Na D* lines are 589.59 nm and 588.99 nm, whose mean value is used as a monochromatic light source.

The working principle of a low-pressure Mercury lamp is almost the same as that of a low-pressure Sodium lamp. The pressure of Mercury vapour is very low (below 1 kpa). The luminous efficiency of the lamp is not high and the lamp is also an arc vapour discharge lamp. In the visible wave band, the following spectra are strong enough and very easy to be observed. 579.07 nm (yellow), 576.96 nm (yellow), 546.07 nm (green), 491.60 nm (cyan), 435.83 nm (blue), 407.78 nm (purple), 404.66 nm (purple).

Procedure

1. Observe the spectra of a Sodium lamp:

a) Turn the power supply of the lamp and warm it up for about 5~10 minutes.

b) Acquire the spectrum of the Sodium lamp with the LEOI-100 spectrometer. Learn to select the output slit (observe window or CCD) by diversion mirror.

c) Scan from 300 nm to 800 nm, observe the spectra of the Sodium lamp and note the values.

2. Observe the spectra of a Mercury lamp.

a) Turn the power supply of the lamp and warm it up for about 5~10 minutes.

b) Scan from 300 nm to 800 nm, observe the spectra of the Mercury lamp and note the values.

5.2 Measure the Rydberg Constant

Objective

1. Measure the Hydrogen spectrum.
2. Verify the Balmer law and the Bohr model.
3. Measure the Balmer lines of Hydrogen atom and calculate the Rydberg constant, then compare

with the theoretical value.

Principle

Bohr Model

In 1913, Bohr proposed a model for the Hydrogen atom to explain the spectrum of a Hydrogen atom. Bohr's model was based on the following assumptions:

1) The electron in a Hydrogen atom travels around the nucleus in a circular orbit. Although the electron performs an accelerated motion, it does not radiate energy. The electron is in a stable state called a steady state.

2) Light is absorbed when an electron jumps to a higher energy orbit; light is emitted when an electron falls to a lower energy orbit. The energy of the light emitted or absorbed is equal to the difference between the energies of the orbits. If an electron leaps from one steady state E_m to another state E_n, it will absorb or emit a photon with a frequency of:

$$\nu = \frac{E_n - E_m}{h}$$

3) Only a limited number of orbits with certain energies are allowed. In other words, the orbits are quantized and the angular momentum of the electron on the orbit is integral times of $\hbar (\hbar = h/2\pi)$:

$$L = m\upsilon r = n\hbar \text{, where } n = 1, 2, 3\ldots$$

According to Bohr assumptions, the energy levels in the Hydrogen atom can be calculated. When the electron is on the orbit, the static attraction force is the centripetal force. For circular motion:

$$m\frac{\upsilon^2}{r} = \frac{e^2}{4\pi\varepsilon_0 r^2} \Rightarrow r = \frac{e^2}{4\pi\varepsilon_0 m\upsilon^2}$$

By Bohr's quantized assumption $r = \frac{n\hbar}{m\upsilon}$, so:

$$\frac{n\hbar}{m\upsilon} = \frac{e^2}{4\pi\varepsilon_0 m\upsilon^2}$$

It is easy to get: $\upsilon = \frac{e^2}{4\pi\varepsilon_0 n\hbar} = \frac{\alpha c}{n}$, $r = \frac{4\pi\varepsilon_0 n^2\hbar^2}{me^2} = a_0 n^2$

Where, $\alpha = \frac{e^2}{\hbar c}$ is the fine structure constant, $a_0 = \frac{4\pi\varepsilon_0\hbar^2}{me^2}$ is the Bohr's radius. The energy of the electron is the sum of kinetic energy and potential energy. When it circles on orbit n, the energy becomes:

$$E_n = \frac{1}{2}m\upsilon^2 - \frac{e^2}{4\pi\varepsilon_0 r} = -\frac{me^4}{8n^2\varepsilon_0^{\ 2}h^2}$$

From the above discussion, it can be seen that the orbital radius r, the motion velocity v, and the energy E are quantized. The state (n=1) is called the ground state, the states (n≥2) are called the excitation states. Substitute n into the above formula, one can get the energy of each energy level. According to Bohr's assumptions, when the electron falls from a higher energy level E_n to a lower energy level E_m, it emits a photon. The wave number of the photon is:

$$\sigma = \frac{1}{\lambda} = \frac{v}{c} = \frac{E_n - E_m}{hc} = -\frac{me^4}{8\varepsilon_0^{\ 2}h^3c}(\frac{1}{n^2} - \frac{1}{m^2}) = R(\frac{1}{m^2} - \frac{1}{n^2}),\ \ n>m$$

where $R = \frac{me^4}{8\varepsilon_0^{\ 2}h^3c}$ is the Rydberg constant. In reality, considering the mass of the Hydrogen atomic nucleus is about 1836 times of that of an electron, the reduced mass, $Mm/(M+m)$, instead of the mass of the electron, m should be used.

Balmer lines

Blamer lines are the electron leaps from the second energy level to others, the wave number of the Balmer liners is described by:

$$\sigma = \frac{1}{\lambda} = R(\frac{1}{2^2} - \frac{1}{n^2}),\ n=3,4,5\ldots \tag{1}$$

$$R = \frac{e^4}{8\varepsilon_0^{\ 2}h^3c}\frac{Mm}{M+m} = R_\infty \frac{M}{M+m} \tag{2}$$

$R_\infty = \frac{e^4 m}{8\varepsilon_0^{\ 2}h^3c}$ is the Rydberg constant when assuming the mass of the nucleus is infinity.

Procedure

1. Turn on the power supply of a Hydrogen lamp and warm it up.

2. Calibrate the spectrometer by a Mercury lamp and then measure the spectra of the Hydrogen lamp.

3. Calculate the Rydberg Constant.