Task 1: Ion Identification

Point A

Features	Intensity	Wavelength	Spectrum
Features A	250	832.7587	O II
Features A	400	833.3302	O II
Features A	500	834.4655	O II
Features B	50	903.6235	C II
Features B	300	903.9616	C II
Features B	750	904.1416	C II
Features B	150	904.4801	C II
Features C	15	926.2256	H I
Features C	20	930.7482	H I
Features C	30	937.8034	H I
Features C	50	949.7430	H I
Features C	100	972.5367	H I
Features C	50	915.613	N II
Features C	50	915.963	N II
Features C	60	916.019	N II
Features C	200	916.708	N II
Features C	80	1025.762	O I
Features D	50	1039.230	O I
Features D	500	1037.0182	C II
Features D	250	1036.3367	C II
Features E	25	1083.994	N II
Features E	60	1084.580	N II
Features E	15	1085.550	N II
Features E	100	1085.710	N II
Features F	1000	1215.66824	H I
Features F	500	1215.67364	H I
Features G	1000	1302.168	O I
Features G	700	1304.858	O I
Features G	300	1306.029	O I
Features H	200	1329.5775	C I
Features H	100	1329.6005	C I
Features H	80	1334.5323	C II
Features H	150	1335.7077	C II

Point B

The Hydrogen spectrum is the simplest of the whole line spectra, since Hydrogen consists only of the nucleus containing a proton (and a neutron) around which an electron moves. Electrons can move in quantized energy orbitals.
The 1s orbital is the innermost one (ground state) to which the minimum of energy corresponds.
In order for this electron to pass from the 1s orbital to an orbital with a higher energy (2s) it is necessary to administer energy which is equal to 10.19 eV.
The energy that must be administered to the electron to allow it to move to orbitals with increasing energy is gradually greater until an energy (ionization energy) is reached, by which the electron moves away from the atom.
Lines belonging to various series ascribed to it, appear simultaneously in the absorption spectrum of Hydrogen (Lyman, Balmer, Paschen, Brackett, Pfund and Humphreys series), which can be calculated using a series of empirical equations that originate from e formula elaborated by Balmer. In general, by supplying energy to gaseous Hydrogen, the number of electrons that are excited is very high. When the atoms return to unexcited conditions, in some cases the electron returns directly to the first orbit (n = 1) in other cases the electron stops momentarily on orbits with n> 1, and then falls back on lower orbits, and this accounts for the plunge near 900 Angstroms.
The Lyman series, in this case, is a sequence of lines that describes the spectral lines of the spectrum of the Hydrogen atom in the ultraviolet region caused by the transition from n ≥ 2 to n → 1 (where “n” is the principal quantum number) that is the lowest energy level for electrons. The transitions are identified with the sequence of the letters of the Greek alphabet, so that the transition from n = 2 to n = 1 is called the Lyman Alpha line (1216 Angstrom), from 3 to 1 we have Lyman Beta, from 4 to 1 the Lyman Gamma and so on.
The greater is the difference between the principal quantum numbers, the greater is the energy of the electromagnetic emission.
The version of the Rydberg equation (generalization of Balmer equation) that generated the Lyman series was:

$$\frac{1}{\lambda} = R_{\mathrm{H}}\left(1 - \frac{1}{n^2}\right) \qquad \left(R_{\mathrm{H}} \approx 1{,}0968 \times 10^7\,\mathrm{m}^{-1} \approx \frac{13.6\,\mathrm{eV}}{hc}\right)$$

Where “n” is a natural number greater than or equal to 2 (that is, n = 2, 3, 4…).
Therefore the multiple lines shown in the image of the exercise are the corresponding wavelengths n = 2 on the right, up to a n = ∞ on the left. The spectral lines are infinitely numerous, but become very dense as n = ∞ approaches (Lyman limit), so that only the first and last are clearly shown.
With Bohr's presentation of his atomic model, an explanation of why the spectral lines of Hydrogen corresponded to the Rydberg formula was obtained. Bohr had found that the electrons of the Hydrogen atom must have quantum energy levels described by the following formula:

$$E_n = -\frac{me^4}{2(4\pi\varepsilon_0\hbar)^2}\,\frac{1}{n^2} = -\frac{13.6\,\mathrm{eV}}{n^2}$$

According to Bohr's assumption, when an electron decays from an initial energy level Ei to a final Ef, the atom must emit radiation of wavelength:

$$\lambda = \frac{hc}{E_{\mathrm{i}} - E_{\mathrm{f}}}$$

In the previous formula, setting the energy of the initial state of Hydrogen corresponding to level “n” and that of the final state corresponding to level “m”, we obtain:

$$1/\lambda = R_{\mathrm{H}}\left(\frac{1}{m^2} - \frac{1}{n^2}\right)$$

Where “R_H“ is the Rydberg constant for Hydrogen in the Rydberg equation.
To relate the expressions of Bohr, Rydberg and Lyman, just replace “m” with 1 to get:

$$\frac{1}{\lambda} = R_{\mathrm{H}}\left(1 - \frac{1}{n^2}\right)$$

which represents the Rydberg formula for the Lyman serie. Therefore, each wavelength of the emission lines corresponds to an electron that decays from an energy level (greater than 1) to the first level which is precisely responsible for the plunge near 900 Angstroms.