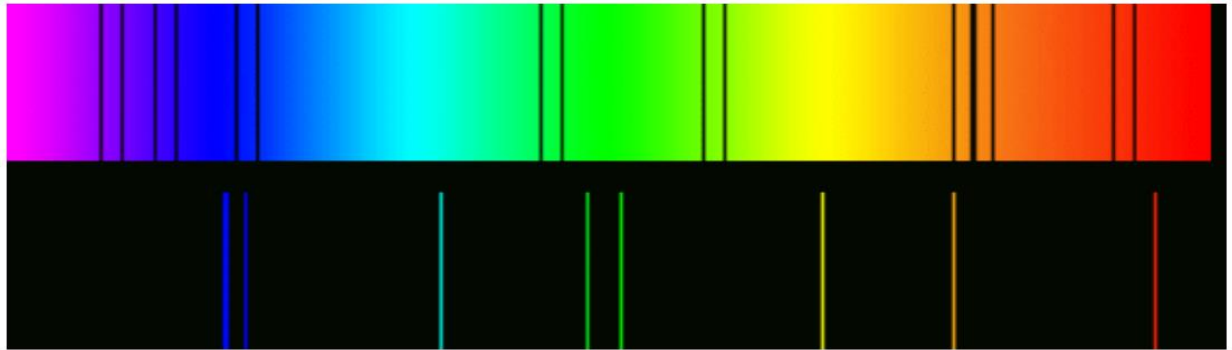


Simple physics of spectral emission of atoms



A bit of theory. An atom can only be in special stationary or quantum states, each of which corresponds to a certain energy. In a stationary state, an atom does not emit electromagnetic waves. Emission and absorption of energy by an atom occurs during a jump-like transition from one stationary state to another. The emission of light by an atom occurs when an electron passes from the upper orbit to the lower one under the action of a quantum (a certain portion) of energy.

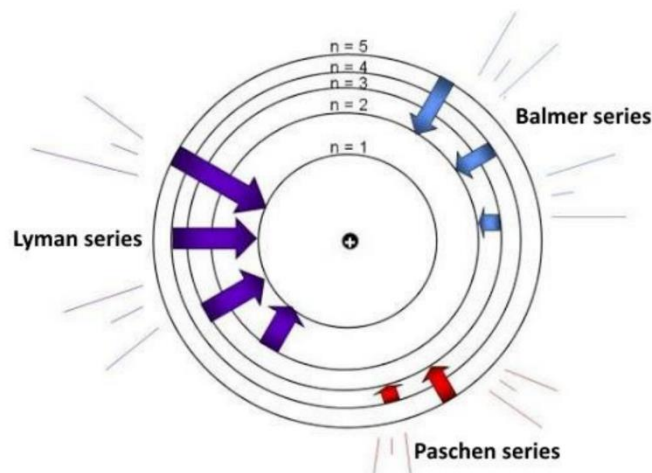


Fig.1

But after all, all the lower orbits, according to the theory, are already occupied and cannot accept electrons above the limit. This means that you first need to "knock out" the electron from the lower orbit. And where will he go? For simplicity, we will assume that the outermost orbit is also fully occupied. Will it start to form a new orbit?! Will he start free floating through the expanses of matter in search of a berth? Okay, let's leave him alone.

This raises questions. Why is the lower electron knocked out, which is more strongly bound to the nucleus, and besides, it is also pressed by the outer electrons to the nucleus? It is more logical to assume that it is the outer electrons that are most likely to be knocked out, which are the first to meet with the perturbing particle and, moreover, are weaker associated with the nucleus! But then there would be no emission spectra. The place of this outer electron knocked out into the interatomic space can only be occupied by another similar "interatomic" electron. In this case, only one frequency would be emitted. No spectrum.

Let's go further. It is also argued that an electron from the upper orbit on its way to the lower one can make a couple of stops in the lower orbits, thereby emitting a couple of photons. But, after all, these orbits are occupied, according to your own rules!

And finally the most important counterargument. How can all electrons in one atom simultaneously make all potentially possible transitions! Well, even if not in one and not all, and not at the same time. Let's say the first atom emits the first frequency, the second atom emits the second frequency, and the third atom emits the third. After that,

the first atom emits the second frequency, the second - the fifth, the third - the twentieth, etc. But, as a result of all this chaos, a stable spectral picture is obtained!? Why, this is just a miracle of probability theory!

And how can one and the same action cause different electronic transitions in identical atoms? In any case, it turns out that the spectrum is not formed immediately, not simultaneously, but sequentially.

But it seems to me that if a single atom could be excited with a single action, it would emit the entire spectrum at once. There is only one small thing left: to manage such an experiment!

Generally speaking, it is difficult to find a more confusing, contradictory and contrived theory than the quantum theory of the structure of the atom.

Let us propose a simpler hypothesis explaining even the emission spectrum of a single atom with one action. It is believed that the frequency f of photons emitted by an atom is equal to the difference between two frequencies, the largest of which represents the boundary of the radiation series.

As is known, the wavelengths of the spectral lines of hydrogen are determined by the Rydberg formula.

$$\frac{1}{\lambda} = R \left(\frac{1}{n_1^2} - \frac{1}{n^2} \right), \quad (1)$$

where λ is the wavelength of electromagnetic radiation, R - Rydberg constant, n_1 - integer defining the series, n - integers, defining spectral lines of this series.

Since the frequency is the reciprocal of wavelength, the Rydberg formula can be represented as the difference between the boundary frequency of a series f_b and the frequencies of the spectral lines of a given series f_n .

$$f = f_b - f_n \quad (2)$$

It was this frequency difference formula that led to the creation of the quantum theory of atomic structure, which assumes that the emission of a spectral line with frequency f by an atom is caused by an electron jumping from a higher orbit where the electron has more energy to a lower orbit. That is, the energy of the emitted photon is equal to the difference in the energy of the electron in the highest and lowest orbits. Recall that photon energy is proportional to frequency $E = hf$.

Let's make permutation in the formula (2):

$$f_b = f + f_n \quad (3)$$

That is, the more energetic (and, quite logically, larger) photon f_b can be split into the sum of two (3, 5, 10, ... n) smaller photons. But what if the emission of a series of photons by an atom is caused by the splitting of a large photon into many small ones!

Fig.2 shows the frequencies of the spectral lines of hydrogen. For clarity, the numerical values of frequencies are presented in vertical columns from top to bottom.

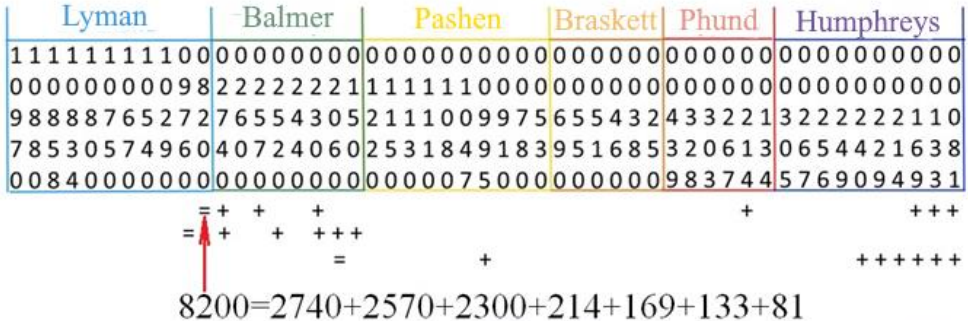


Fig.2

The frequency (as well as the wavelength) of any of the spectral lines can be represented by the sum of the frequencies of other lines. For example, a photon with a frequency of 8200 (the last column in the Lyman series) can split into photons (indicated by pluses) with frequencies: $8200 = 2740 + 2570 + 2300 + 214 + 169 + 133 + 81$. And so on for all frequencies with many combinations of terms.

And about how it happens physically. In spectral analysis, a sample of a chemical element is heated. It is heated by energetic photons, that is, photons with a high frequency, which bombard the nuclei of atoms. In this case, these energetic photons are split, falling on the uneven surface of the nucleus.

The structure of this surface is unique for each chemical element, therefore the composition and sizes (frequencies) of these fragmentation photons are also unique for each element.

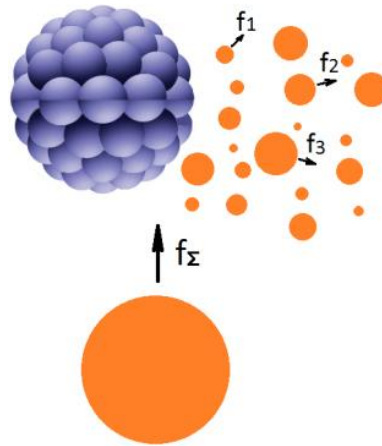


Fig.3

And so, instead of such simplicity, we are offered dozens of layers of theoretical heaps, stretching one after another in an attempt to adapt to reality!