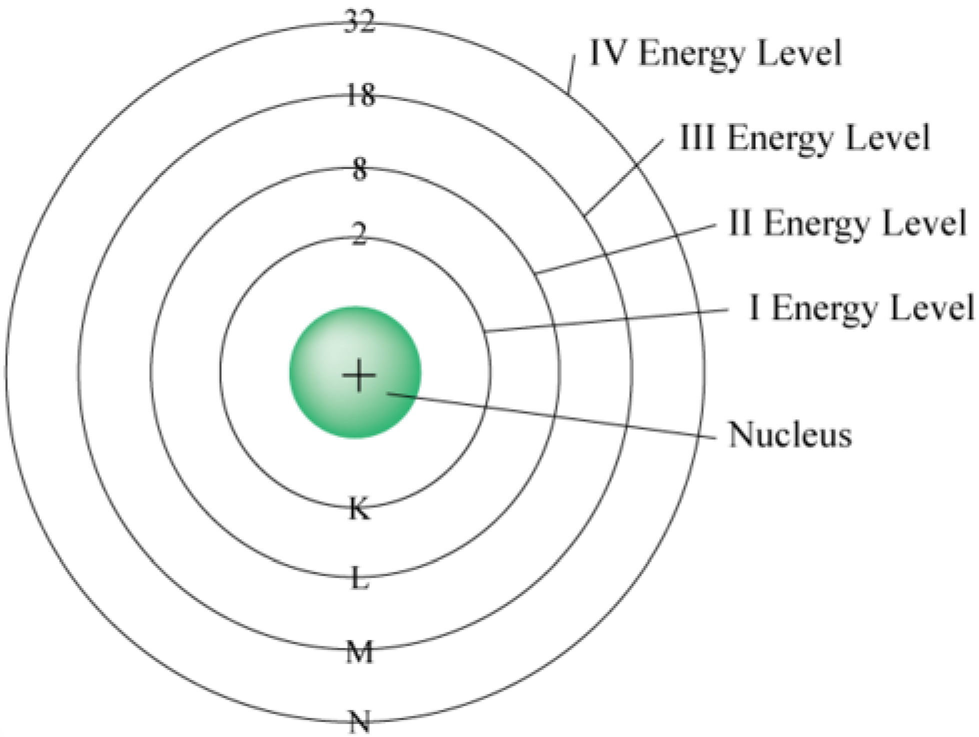
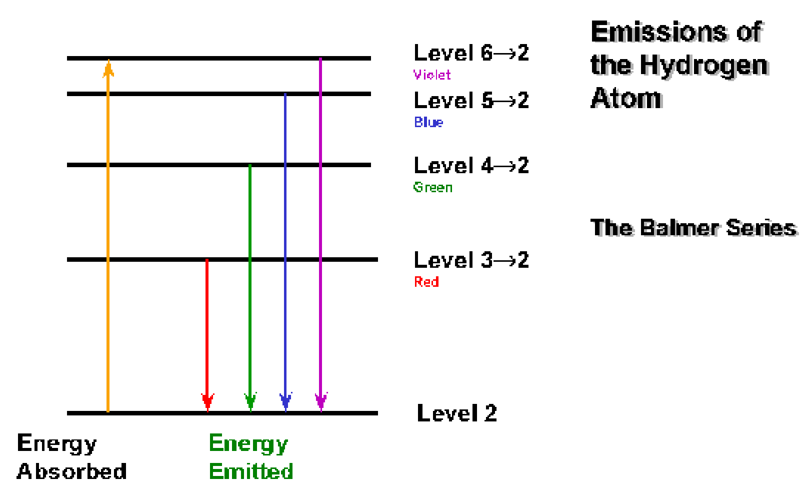
7.1 Explained

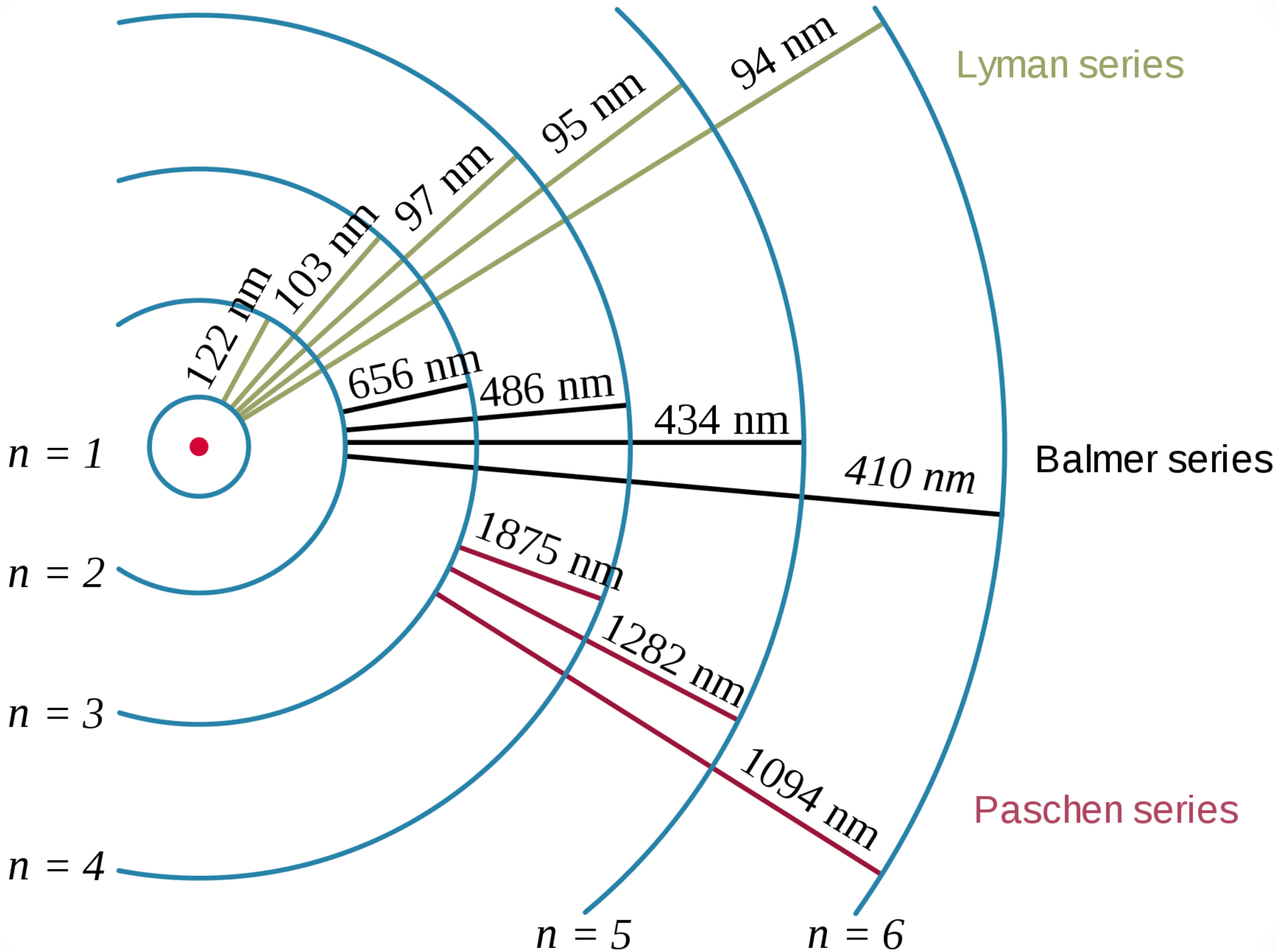
- Atoms contain electrons orbiting around the nucleus, and those electrons sit inside **energy levels** (also known as shells). Energy levels are particular orbits, or particular amounts of energy, that electrons are allowed to have. Some examples are shown below:
- Examples of Atomic Energy Levels: We can draw them as circular orbits (right), or represent them as straight lines, like floors of a hotel (left).
- The electrons in an atom can absorb energy and jump from a lower energy level to a higher one. But only if they get exactly the right amount of energy to make the jump. If they jump too far, or not far enough, they will miss! And if they're going to miss, then they stay where they are.

Main Energy Level	Sublevels	Orbitals	Maximum Electrons
4	4f	seven 4f orbitals	× 2 = 14
	4d	five 4d orbitals	× 2 = 10
	4p	three 4p orbitals	× 2 = 6
	4s	one 4s orbital	× 2 = 2
3	3d	five 3d orbitals	× 2 = 10
	3p	three 3p orbitals	× 2 = 6
	3s	one 3s orbital	× 2 = 2
2	2p	three 2p orbitals	× 2 = 6
	2s	one 2s orbital	× 2 = 2
1	1s	one 1s orbital	× 2 = 2



We can draw atomic energy level as circular orbits or represent them as straight lines, like floors of a hotel.




Lyman series
Balmer series
Paschen series
122 nm
103 nm
97 nm
95 nm
94 nm
656 nm
486 nm
434 nm
410 nm
1875 nm
1282 nm
1094 nm
n = 1
n = 2
n = 3
n = 4
n = 5
n = 6

Suppose an electron exists in a higher energy level (larger value of *n*). Systems generally like to be in as low an energy state as possible, and so the electron would fall into a lower energy state. However, this process by itself would involve a loss of energy, and so energy would not be conserved. To conserve energy, it was postulated that a particle called a **photon** carry away the excess energy in this transition, as pictured to the right.

Note that the energy of this photon is fixed, as it must be equal to the known difference in the energy levels of the electron transition. It would be natural to identify this photon with the light found when atoms emit light, and for this we need to know the frequency of the emitted light. Another hypothesis was put forward relating the energy of the photon to the frequency of the light wave:

Energy = h x frequency where ``h'' is again Planck's constant.

Note that, for a given atom, there can be many types of transitions from higher to lower energies, and thus many different energies of emitted photons and, subsequently, many different frequencies of light waves. Some of these are illustrated to the right.

Figure 12.4: Photon emission

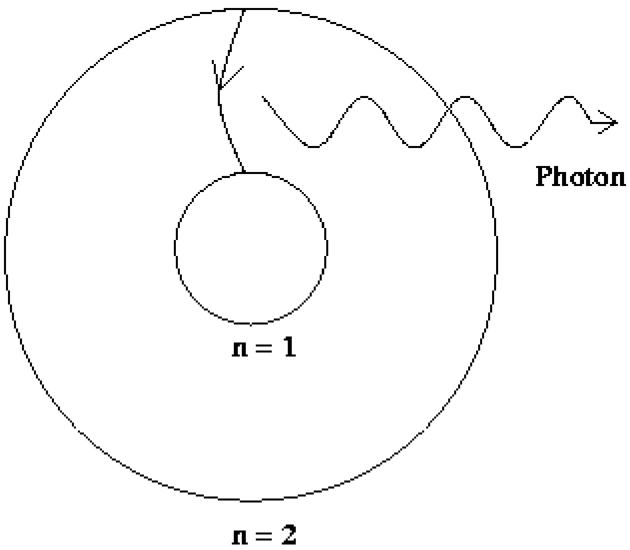
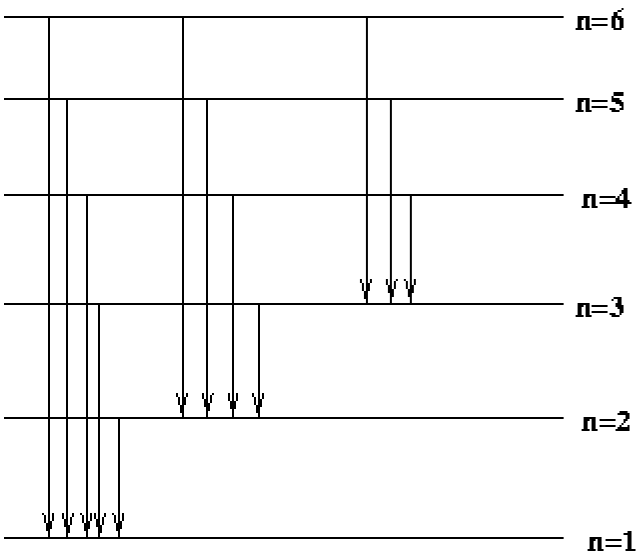


Figure 12.5: Possible transitions in the Bohr atom



Transitions ending up at the n = 2 level give rise to photons in the *Balmer series*, which is light in the visible part of the electromagnetic spectrum. Transitions to the n = 1 level give rise to the *Lyman series*, and to the n = 3 level give the *Paschen series*. Such series of distinct spectral lines are observed experimentally when a gas such as hydrogen is heated up - the added heat energy excites the electrons, which subsequently fall into lower energy levels, emitting photons of various energies and, thus, light of various frequencies. The observed frequencies correspond very well to what is predicted in the Bohr model. This success of the Bohr model in explaining the **emission spectra** of simple atoms gave birth to an intense, and still on-going, investigation into what we now call **quantum theory**.

Absorption spectrum

A related phenomenon to the emission spectra of elements is the **absorption spectra**. Imagine that we have an electron in a lower energy state, and a photon comes along. If this photon has just the right amount of energy, it can be absorbed, causing the electron to make a transition to a higher energy state.

This will only occur though if the photon has an energy equal to the energy difference between the two levels involved in the transition - if not, the photon will just pass through. Such an effect can be seen by shining light of all different frequencies through a gas of a particular element. Some of the photons will be absorbed in the gas, and so will be missing in the light that emerges from the gas. What one will therefore see from the emerging light is an almost continuous band of frequencies, but with some frequencies missing corresponding to the absorbed photons.

As with the emission spectrum, each element has its own unique absorption spectrum, as the energy levels of different elements differ. Measuring the absorption spectrum can therefore be used to identify elements in an unknown substance by comparing it to the spectrum of known elements. Among other areas, this is a major technique used to identify elements in gaseous clouds in galaxies.

Figure 12.6: Photon absorption