

# Atoms and electrons

Atoms are so small that, even today, direct visual inspection is a challenge.

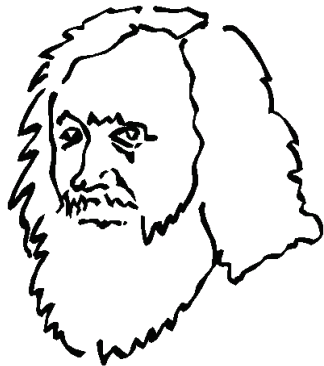
Our model of the atom changes as our experimental ability improves. Today's atomic theory tries to explain the observations made with accelerators (Large Hadron Collider). The current "quark model" of the atom is a hypothesis based on current atomic theory.



Demokrit

## *The Greek Model:*

Democritus (400 BC) concluded that matter could not be divided into smaller and smaller pieces forever. Eventually, the smallest piece of matter would be found. He used the word *atomos* to describe the smallest possible piece of matter.



Dalton

## *The Dalton Model:*

Points of Dalton's Theory (1803):

1. All elements are composed of indivisible particles.
2. Atoms of the same element are exactly alike.
3. Atoms of different elements are different.
4. Compounds are formed by joining atoms of two or more elements.

# Atoms and electrons



J. J. Thomson

## *The Thomson Model:*

J. J. Thomson - English scientist who discovered electrons in 1897.

Sometimes called the "plum pudding" model, Thomson thought of an atom as being composed of a positively charged material with the negatively charged electrons scattered through it.



Rutherford

## *The Rutherford Model:*

Ernest Rutherford - British physicist who discovered the nucleus in 1908.

Rutherford's model proposed that an atom is mostly empty space. There is a small, positive nucleus with the negative electrons scattered around the outside edge.

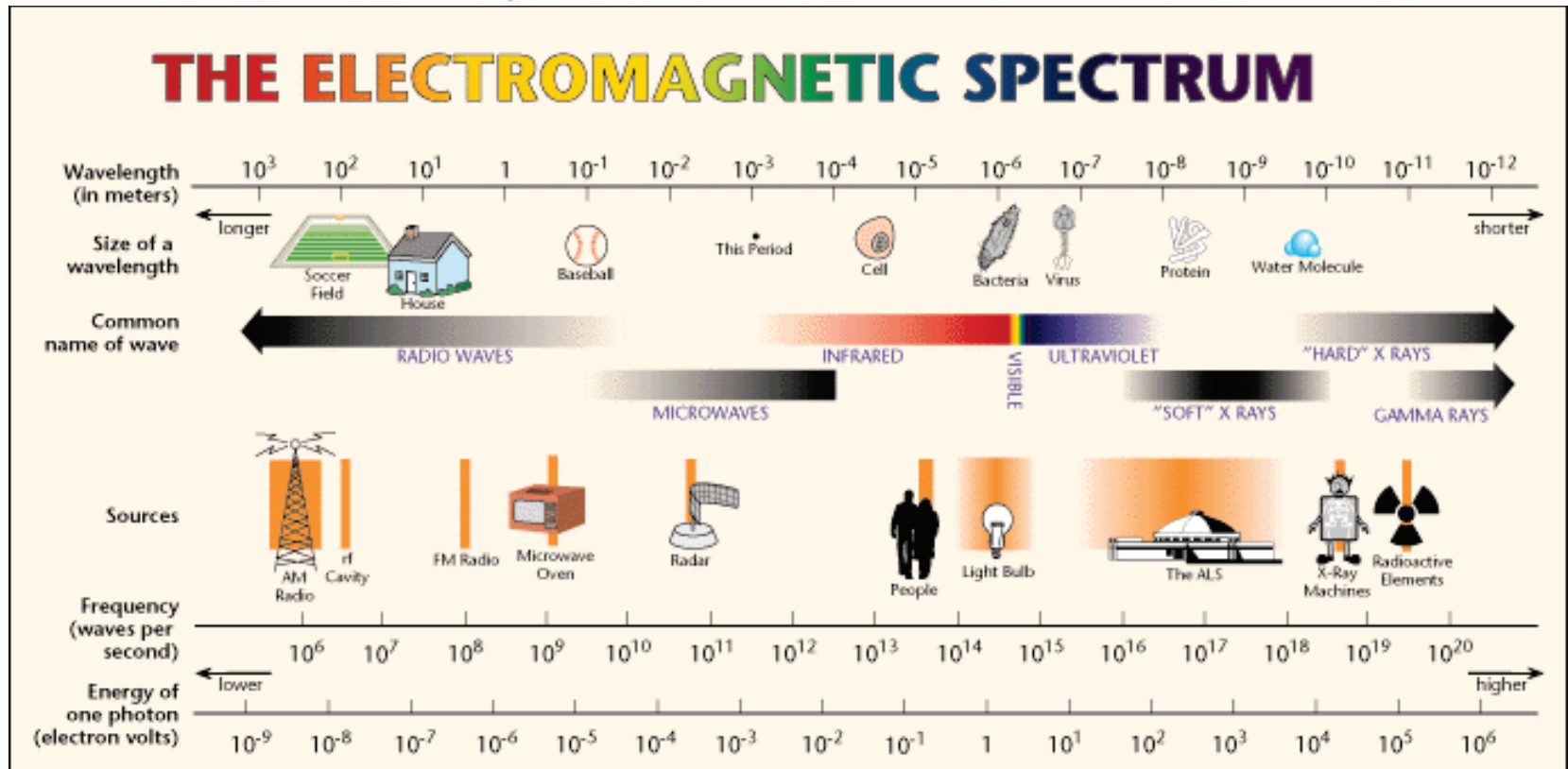
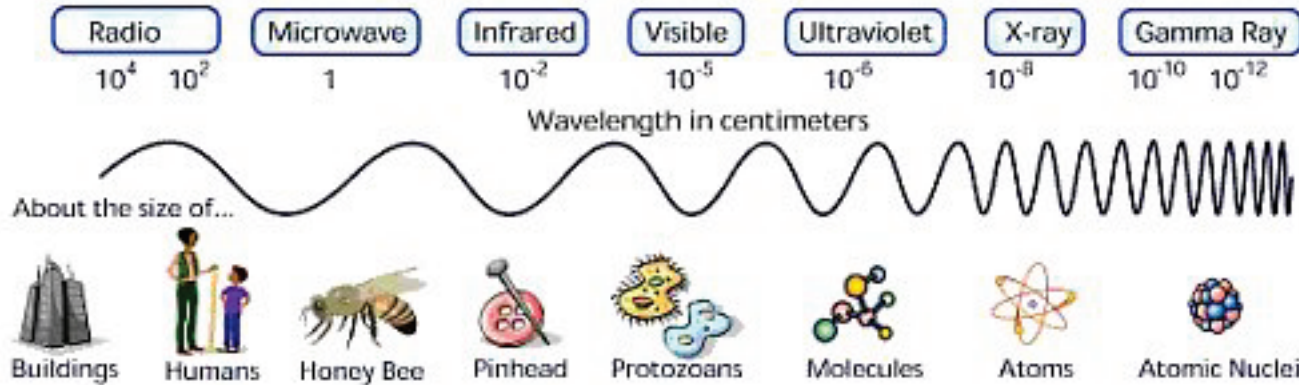


## *The Bohr Model:*

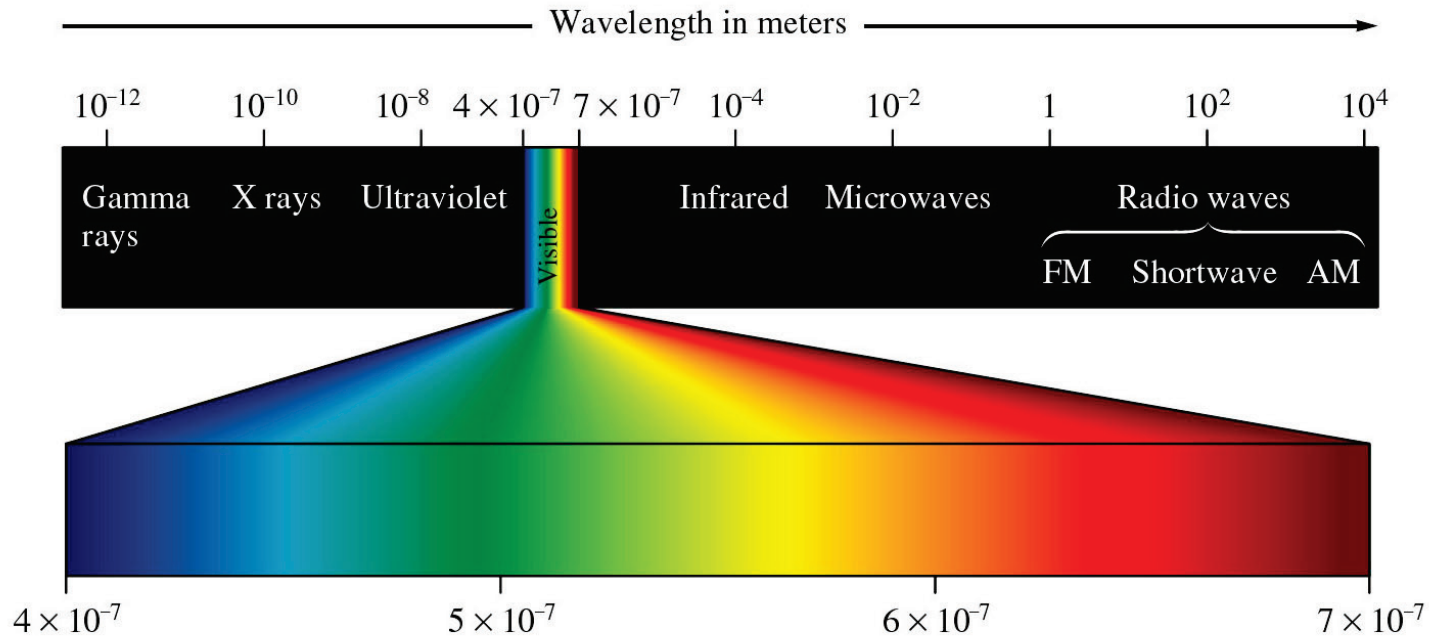
Niels Bohr - Danish scientist who proposed the Planetary Model in 1913.

Electrons move in definite orbits around the nucleus, like planets moving around the nucleus. Bohr proposed that each electron moves in a specific energy level.

# The electromagnetic spectrum



# Visible light waves



Visible light waves are the only electromagnetic waves we can see. We see these waves as the colors of the rainbow.

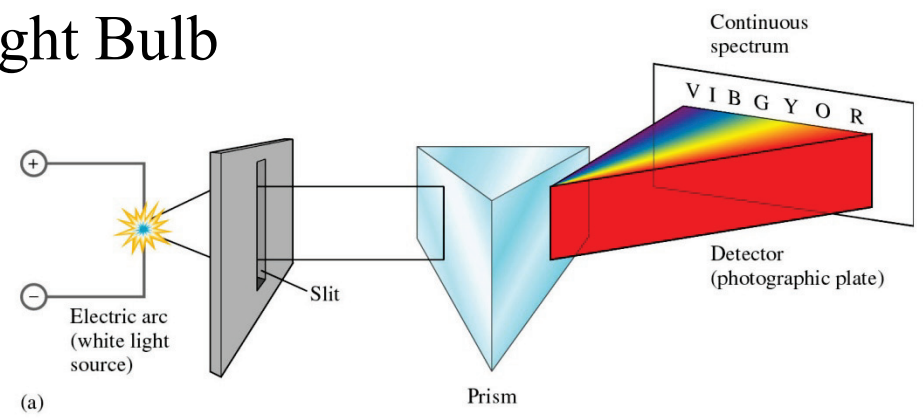
Each color has a different wavelength.

Red has the longest wavelength and violet has the shortest wavelength.

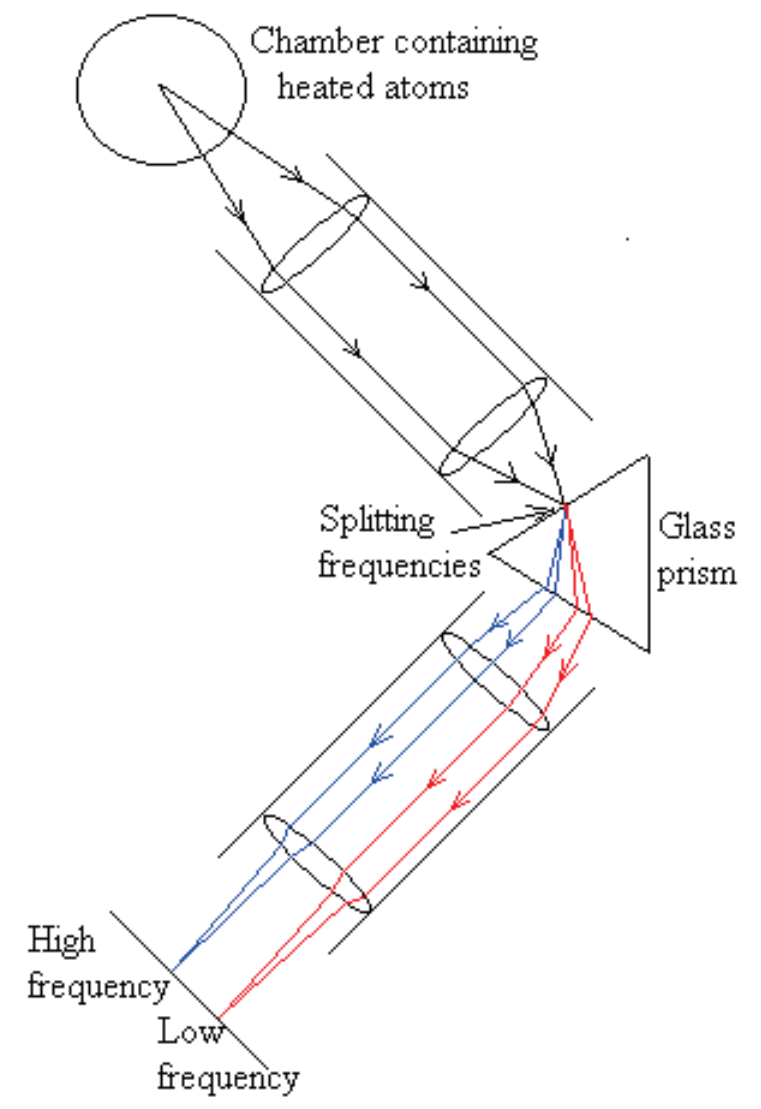
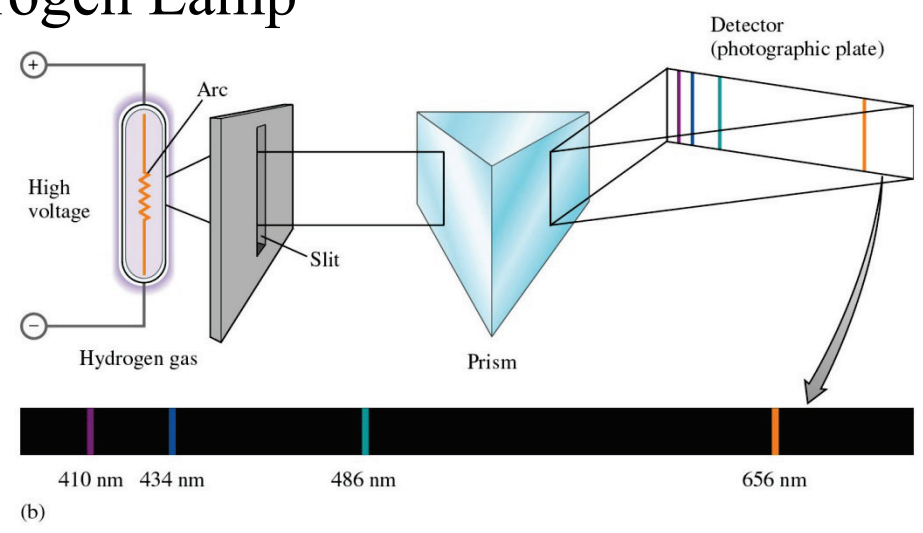
When all the waves are seen together, they make white light.

# Spectrum of hydrogen

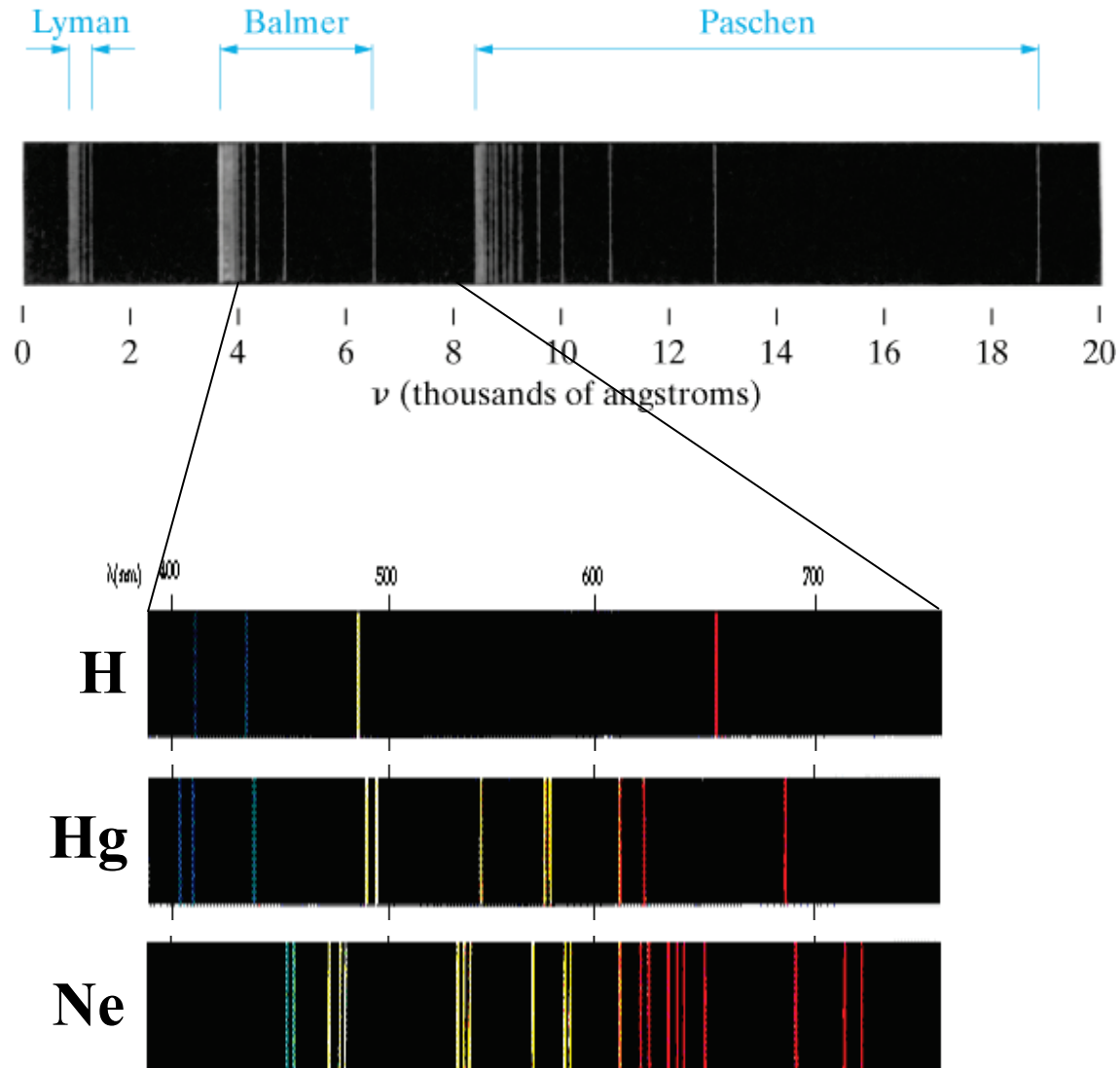
## Light Bulb



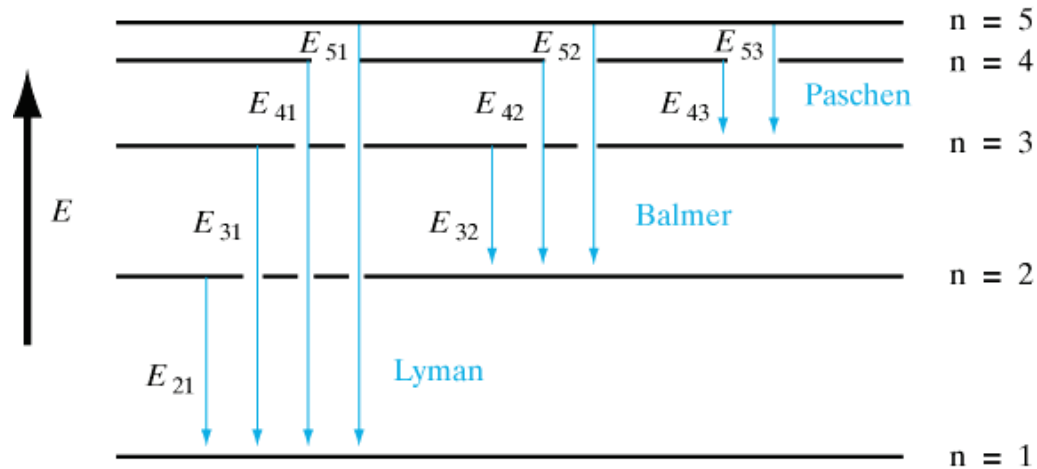
## Hydrogen Lamp



# Emission spectrum of gases



# Electron transitions in hydrogen atom



The series in the hydrogen spectrum follow *empirical* forms as:

$$\nu = cR \left( \frac{1}{1^2} - \frac{1}{n^2} \right) \quad \text{for Lyman (ultraviolet), where } n = 2, 3, 4, \dots$$

$$\nu = cR \left( \frac{1}{2^2} - \frac{1}{n^2} \right) \quad \text{for Balmer (visible), where } n = 3, 4, 5, \dots$$

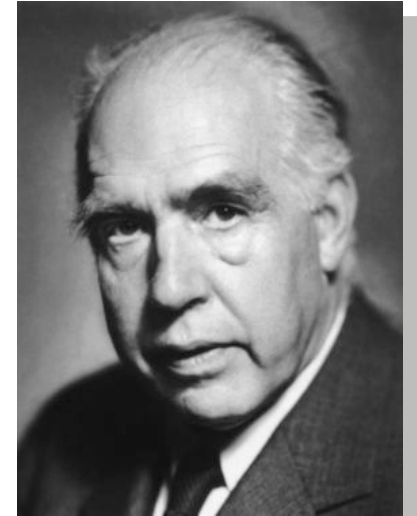
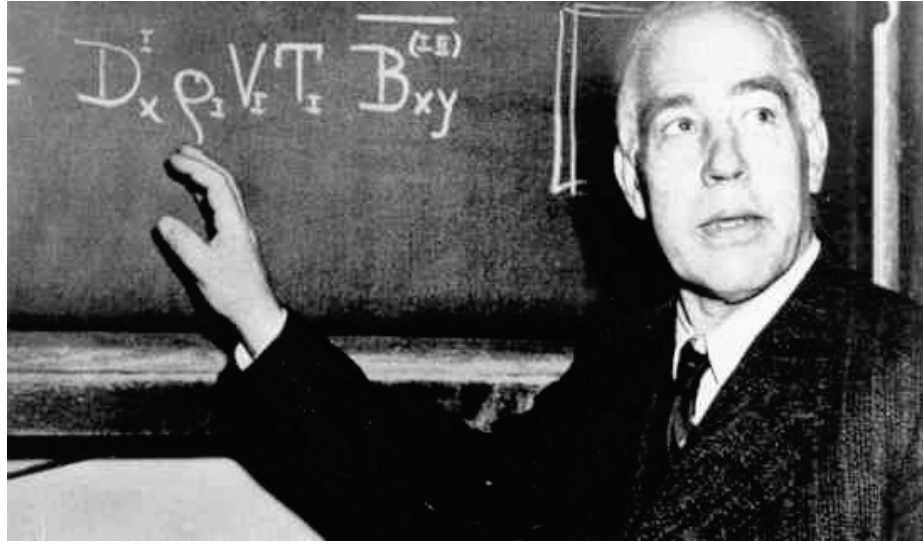
$$\nu = cR \left( \frac{1}{3^2} - \frac{1}{n^2} \right) \quad \text{for Paschen (infrared), where } n = 4, 5, 6, \dots$$

$\nu$  [nju:] – light wavelength;  $R \approx 1.1 \times 10^5 \text{ cm}^{-1}$  – Rydberg constant;  $c$  – speed of light.



# Bohr's model of atom

The model was developed by Niels Bohr (1885 – 1962)



In 1913, Bohr combined the concept of the nuclear atom (Rutherford) with the quantum theory (Planck) in a way that accounted quantitatively for the lines in the emission spectrum of the hydrogen atom. He postulated that the single valence electron of a hydrogen atom revolves around the nuclear proton in only certain allowed circular orbits, and that radiation emitted by the electron, if it moves from one orbit to another, has a frequency proportional to the energy difference of the two orbits. For more complicated atoms, Bohr developed, by ingenious methods, a scheme for atomic structures that corresponded with the periodic table. **Bohr's work on atomic structure was recognized by the 1922 Nobel Prize in Physics.** His Institute for Theoretical Physics in Copenhagen became a center for the study of atomic physics. He promoted peaceful uses for atomic energy and, in 1957, received the first U.S. Atoms for Peace Award.



# Bohr's postulates

To develop the model of atom, Bohr made several *postulates*:

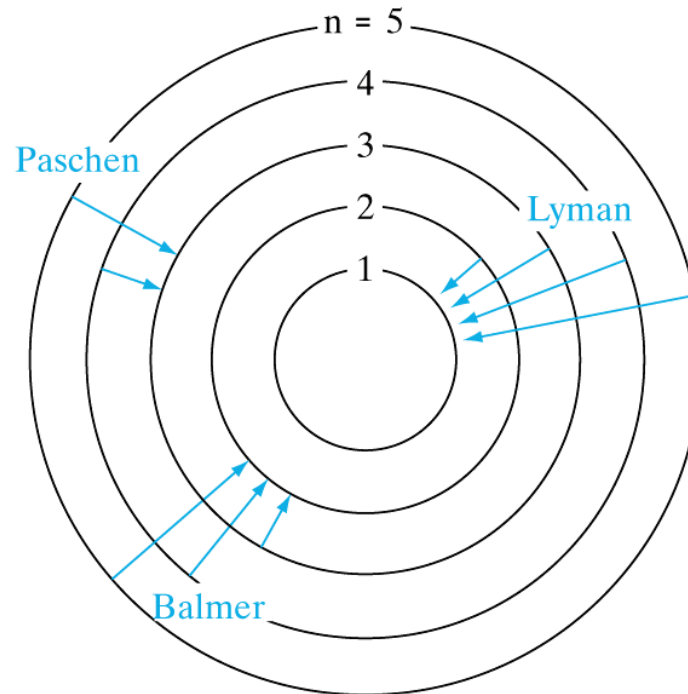
1. Electrons exist in certain stable, circular orbits about the nucleus (without radiation of light).
2. The electron may shift to an orbit of higher or lower energy with gaining or losing energy equal to the difference in the energy levels by absorption or emission of a photon of energy  $h\nu$
3. The *angular momentum*  $p_\theta$  of the electron in an orbit is always an integer multiple of Planck's constant divided by  $2\pi$  ( $h/2\pi = \hbar$ ):  $p_\theta = n\hbar$ , where  $n = 1, 2, 3, 4, \dots$

# Bohr's postulates – 2

Thus, according to Bohr, electron in the atom obeys the following rules:

1. angular momentum is quantized (i.e. cannot have just any value)
2. angular momentum is restricted to values for which “ $n$ ” is a positive integer
3. angular momentum can change only by discrete amounts (i.e. integral multiples of  $h/2\pi$ )

# Bohr's model – 1



Electron orbits and transitions  
in the Bohr's model of the hydrogen atom

# Bohr's model

Energy difference of electron when it is in orbits  $n_1$  and  $n_2$  :

$$E_{n_1} = -\frac{mq^4}{2K^2 n_1^2 \hbar^2} \quad E_{n_2} = -\frac{mq^4}{2K^2 n_2^2 \hbar^2}$$

$$\begin{aligned} E_{n_2} - E_{n_1} &= \left( -\frac{mq^4}{2K^2 n_2^2 \hbar^2} \right) - \left( -\frac{mq^4}{2K^2 n_1^2 \hbar^2} \right) \\ &= -\frac{mq^4}{2K^2 \hbar^2} \left( \frac{1}{n_2^2} - \frac{1}{n_1^2} \right) \end{aligned}$$

This expression looks similar to empirical formula for hydrogen spectrum:

$$\nu = cR \left( \frac{1}{1^2} - \frac{1}{n^2} \right)$$

$$\nu = cR \left( \frac{1}{2^2} - \frac{1}{n^2} \right)$$

$$\nu = cR \left( \frac{1}{3^2} - \frac{1}{n^2} \right)$$

for Lyman, where  $n = 2, 3, 4, \dots$  for Balmer, where  $n = 3, 4, 5, \dots$  for Paschen, where  $n = 4, 5, 6, \dots$

# Bohr's model

If an electrons drops from higher orbit to lower one, it emits light – a photon.

Transition of an electron from lower orbit to the higher is possible when it absorbs a photon.

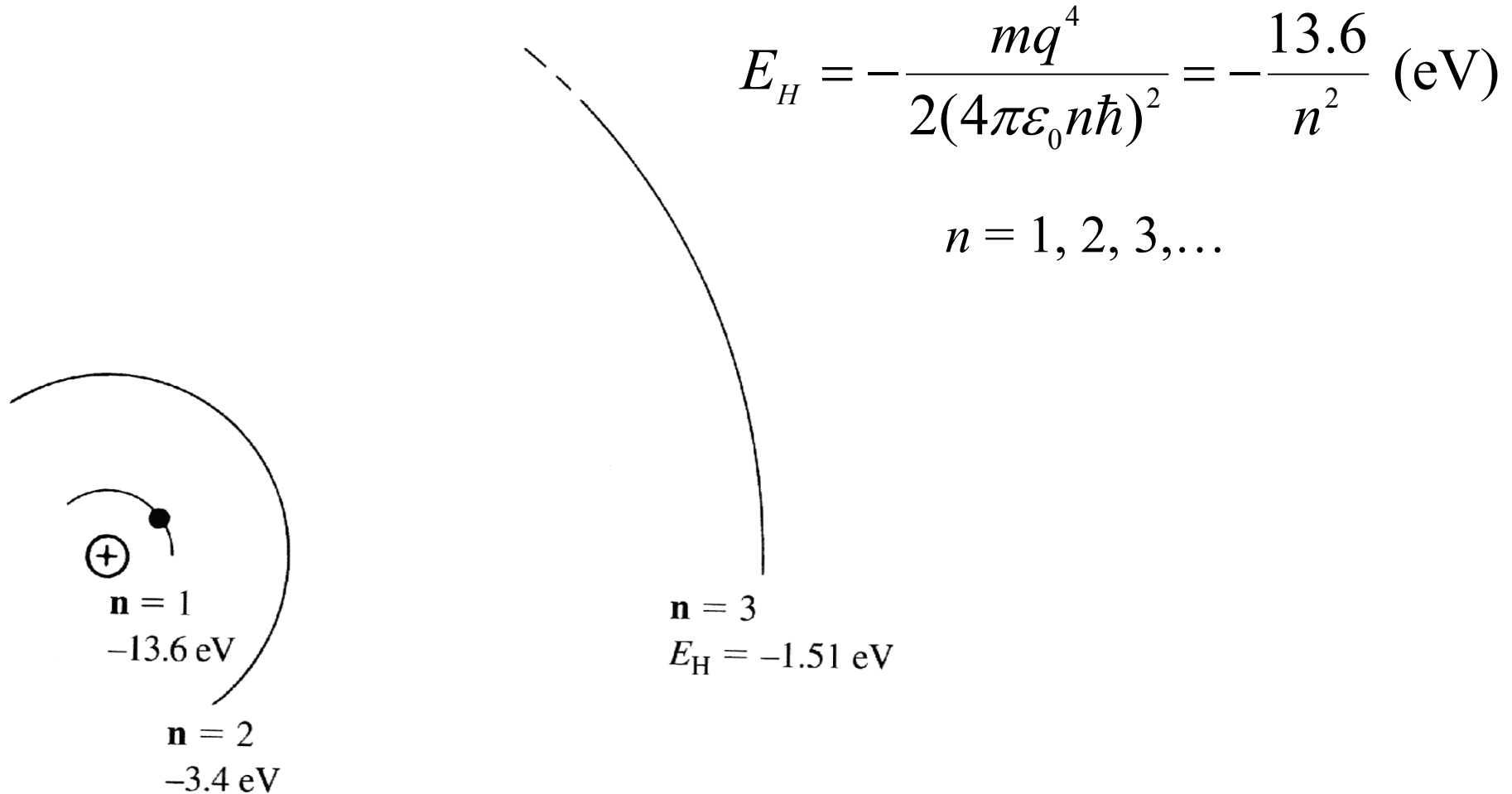
Energy of emitted photon is  $E = h\nu$

Frequency of emitted light when electron drops from higher orbit to lower one:

$$\nu = \frac{mq^4}{2K^2\hbar^2h} \left( \frac{1}{n_2^2} - \frac{1}{n_1^2} \right)$$

where  $K = 4\pi\epsilon_0$

# Energy levels of Hydrogen atom

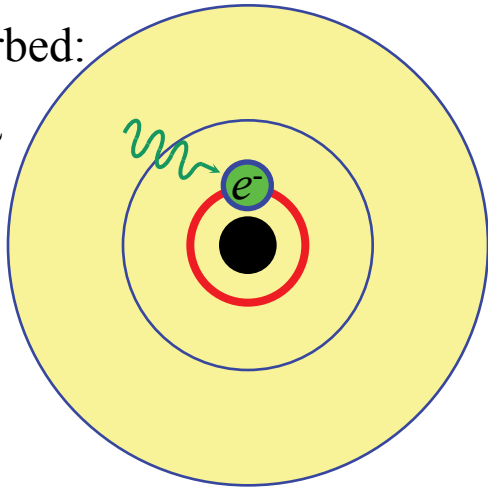


[The electron volt (eV) is a unit of energy equal to  $1.6 \times 10^{-19}$  joules]

# Atomic excitation

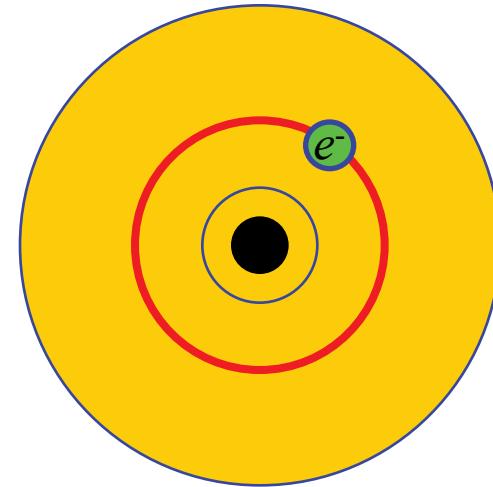
Energy absorbed:

$$E = hc/\lambda$$



Atom in *normal* state

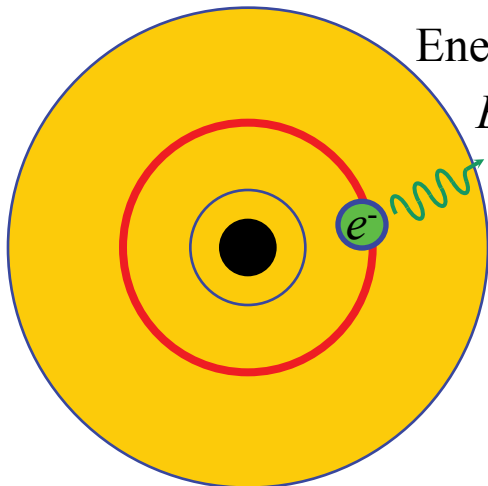
*Atomic  
excitation*



Atom in *excited* state

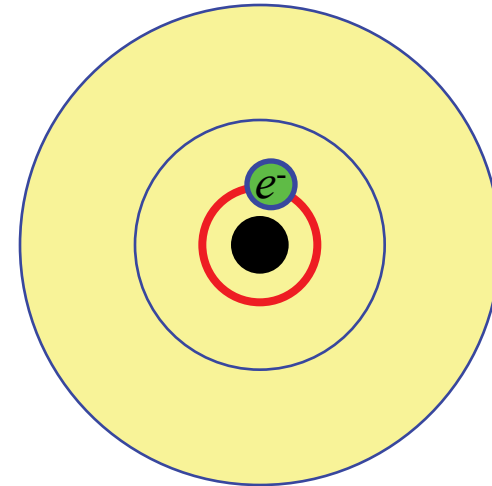
Energy released:

$$E = hc/\lambda$$



Atom in *excited* state

*Atomic  
relaxation*



Atom in *normal* state

*Not to scale*



# Deficiencies of the Bohr model

- The Bohr model could not explain the experimentally observed splitting of levels in addition to the levels predicted by the theory.
- It is difficult to extend the Bohr model to atoms more complex than hydrogen.

Although the Bohr model was a substantial step in understanding nature of atoms, a more comprehensive theory was needed.

# Spectrum of light emitted by hydrogen atom

