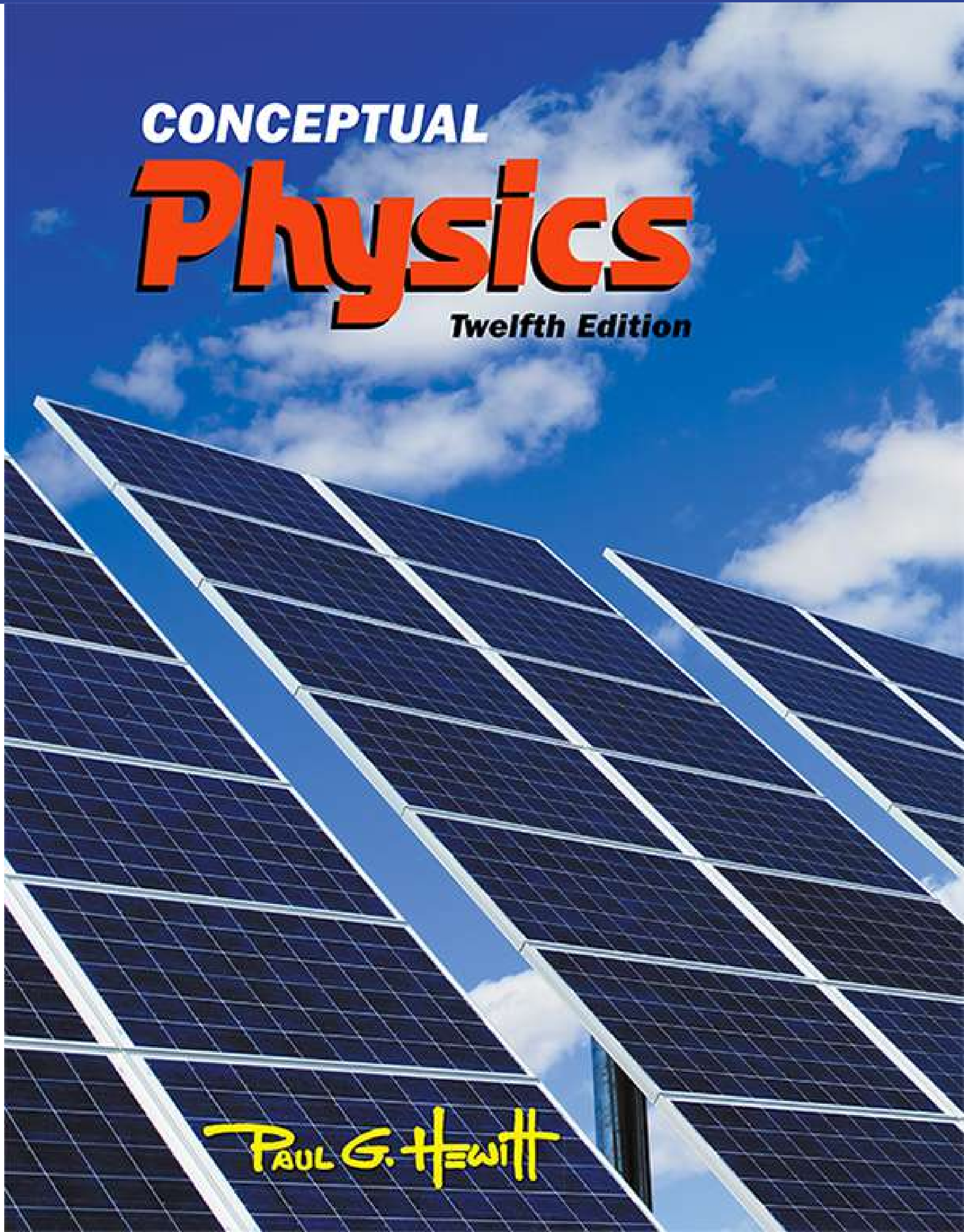
Lecture Outline

Chapter 32: The Atom and the Quantum



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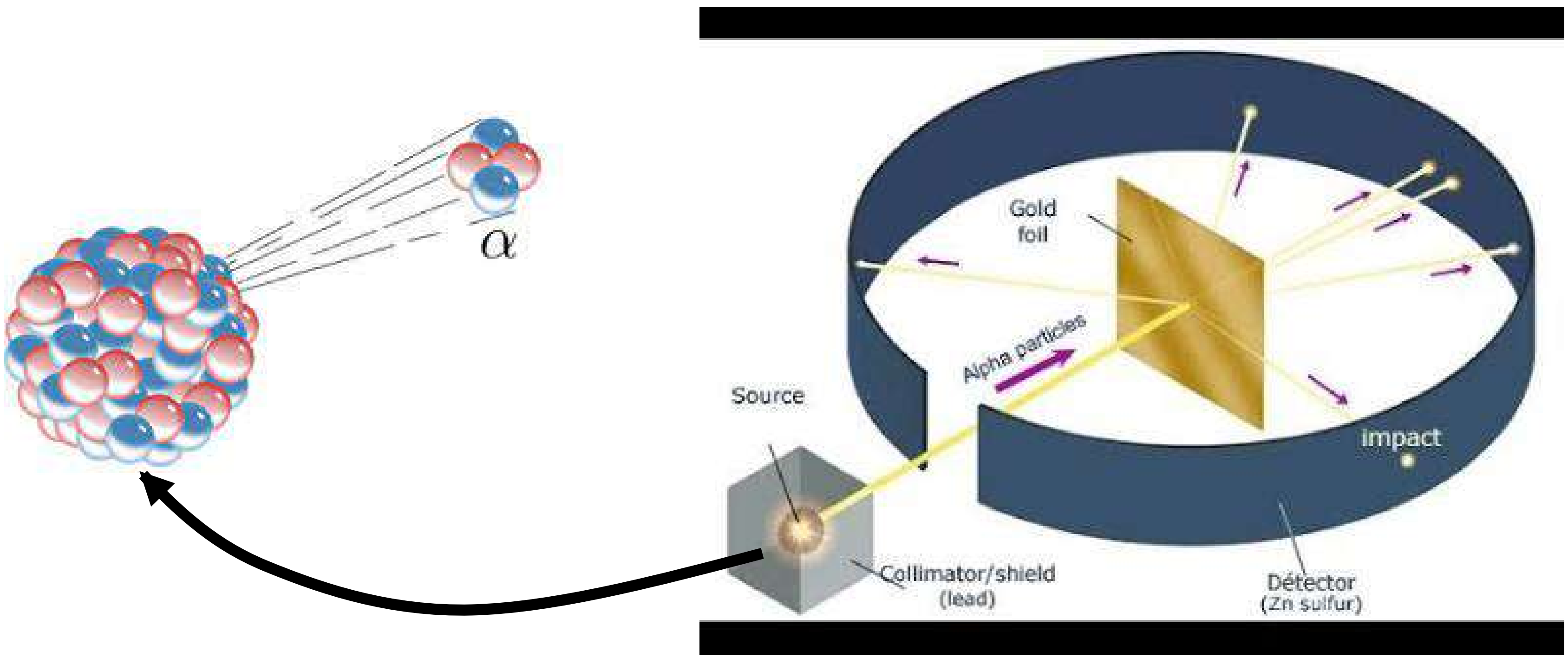
Discovery of the Atomic Nucleus

- In 1924, Ernest Rutherford performed an experiment in which a beam of positively charged **alpha particles** (the nucleus of a helium atom) was directed through a thin gold (Au) foil.



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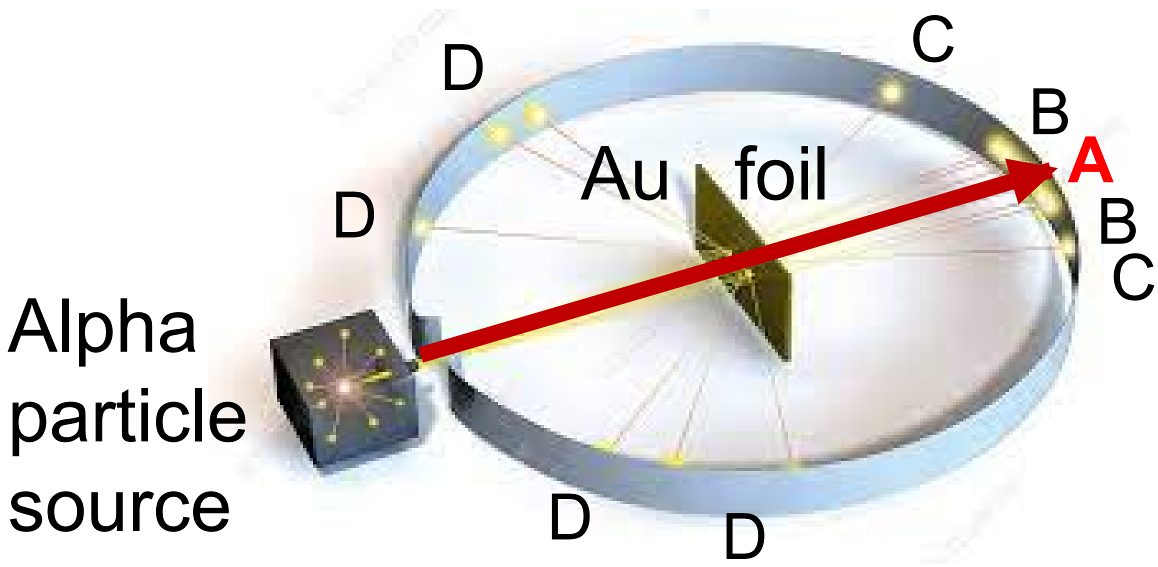
Results:



Most alpha particles passed through with little or no deflection (see A).

Some particles were deflected from their paths (B).

A few alpha particles were widely deflected (C)

A very small number were scattered backward (D)!

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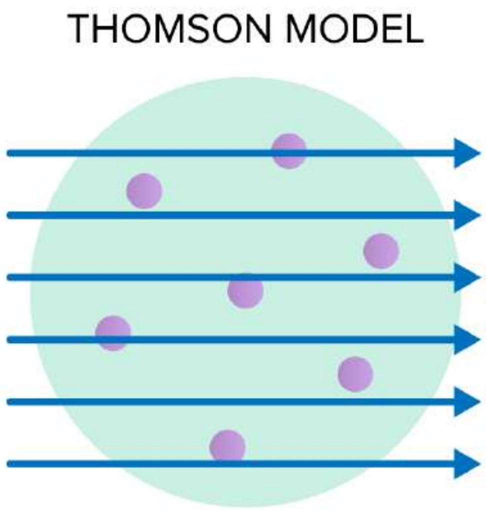
What did this mean?

Before this experiment, it was thought that the positive charge on the atom as like a pudding, with the electrons like plums in the pudding.

→ If true, the alpha particles should pass straight through the pudding.

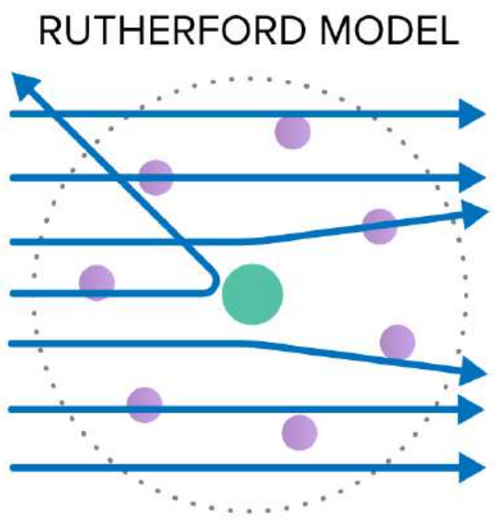


But they didn't. So, Rutherford reasoned:

1/ the undeflected particles pass through the mostly empty space in regions of the foil;

2/ the small number of deflected particles were repelled from extremely dense, positively charged central core.

3/ Every atom must contain a core, which he named the **atomic nucleus.**



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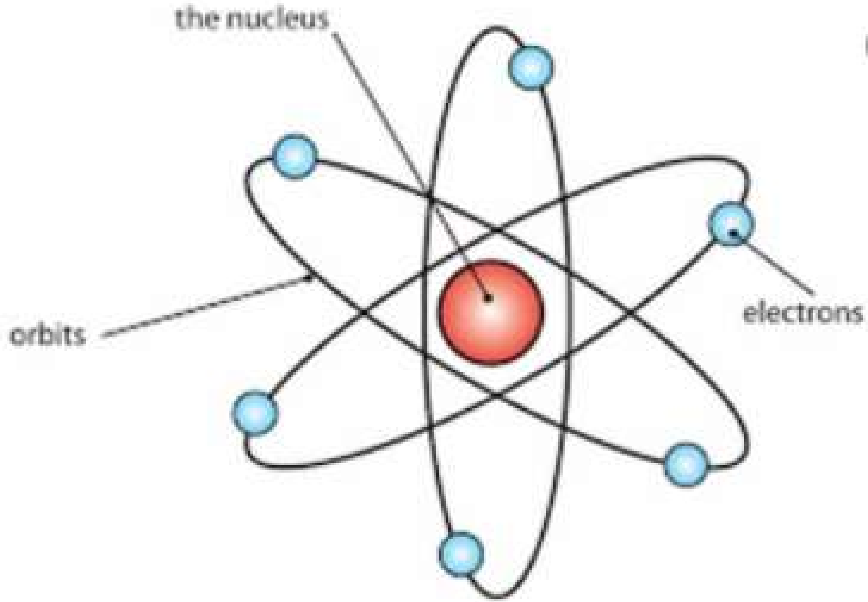
The Rutherford model of the atom:

- *Most the atom is empty space.*
- The positive charge is concentrated in a tiny volume compared to the rest of the atom.
- This is called the **nucleus**.

The negative charge moves around nucleus in orbits, just like the planets orbit the Sun.



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Classwork

1. Why do most alpha particles fired through a piece of gold foil emerge almost undeflected?

2. What did Rutherford discover about the atomic nucleus?

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Electron: A Short History

Greek for *amber* (fossilized tree sap):

Rubbed pieces attracted dust and paper.

Benjamin Franklin:

Postulated the electricity was like a fluid that flowed.

Objects with an excess were electrically *positive*.

Objects with a lack of it were electrically *negative*.

Kite experiment showed that lightning is electricity.

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Classwork

3. What did Benjamin Franklin postulate about electricity?

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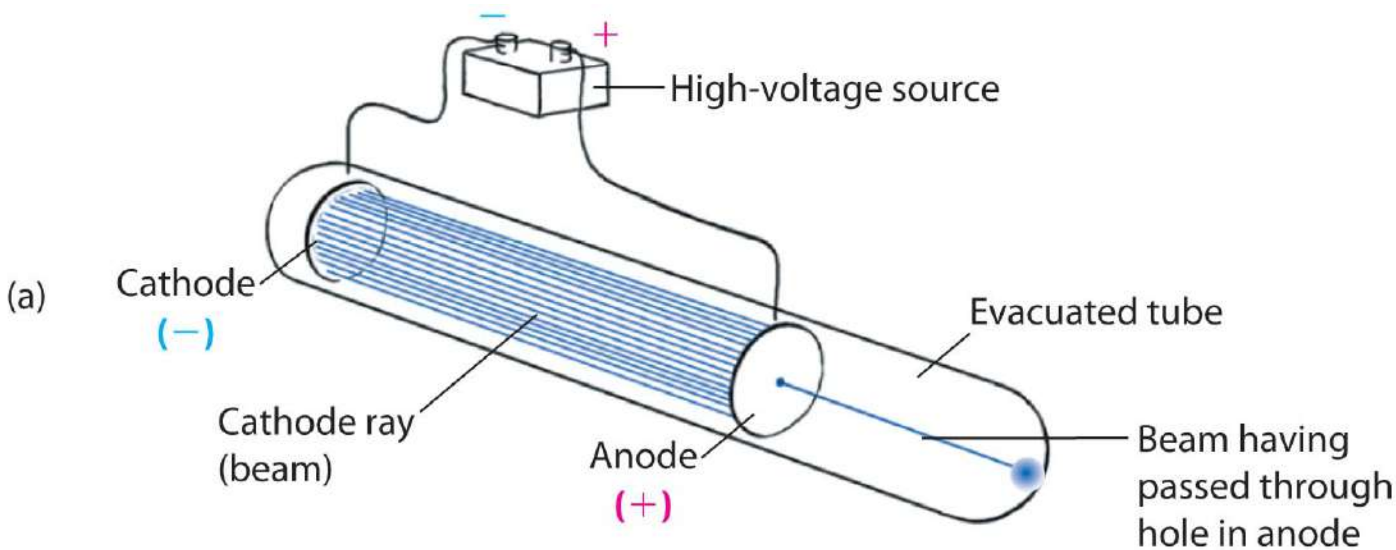
Discovery of the Electron

Electrodes in a glass tube under low pressure are connected to a voltage source, causing the gas to glow due to a "ray" emerging from the negative terminal (the *cathode*).

This is a cathode ray tube:



- Slits could make the ray narrow and plates could prevent the ray from reaching the positive terminal (the *anode*).
- The cathode “ray” is glowing gas emitted from the cathode.

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Classwork

4. What is a cathode ray?

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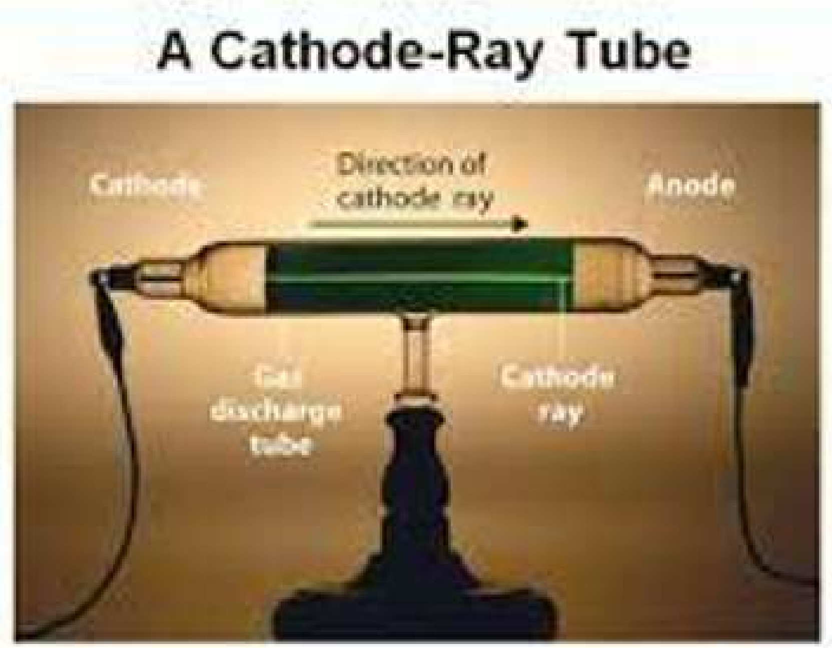
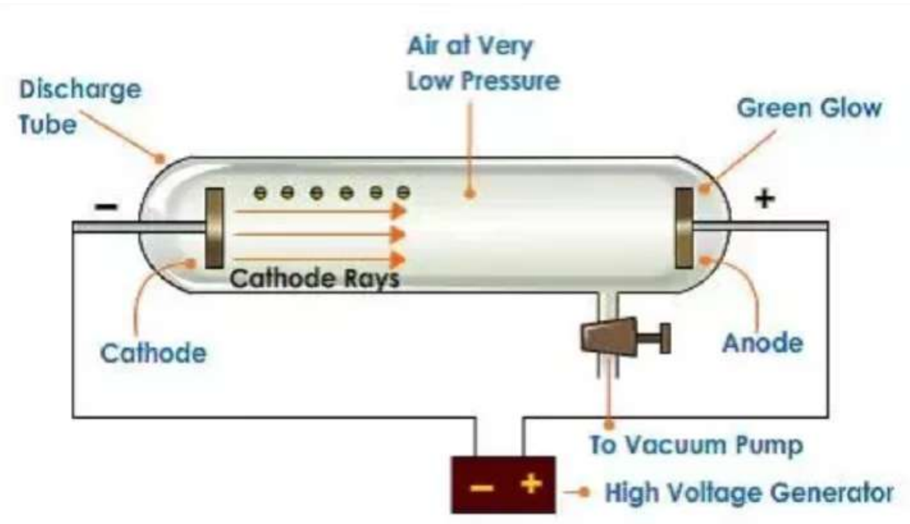
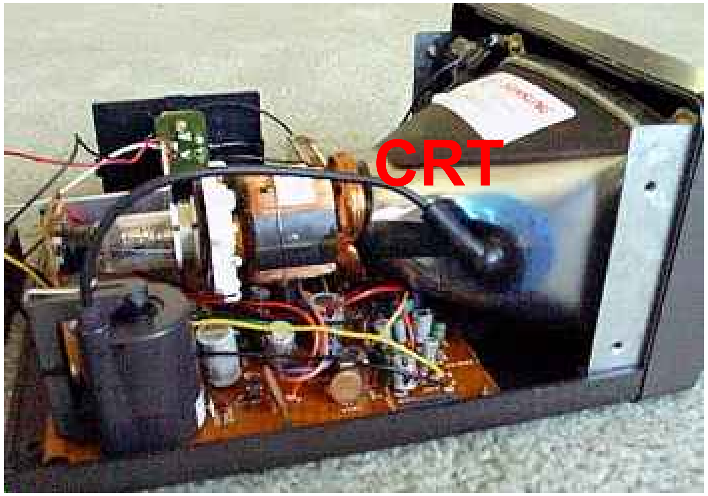
CRTs were in many early electronic devices:







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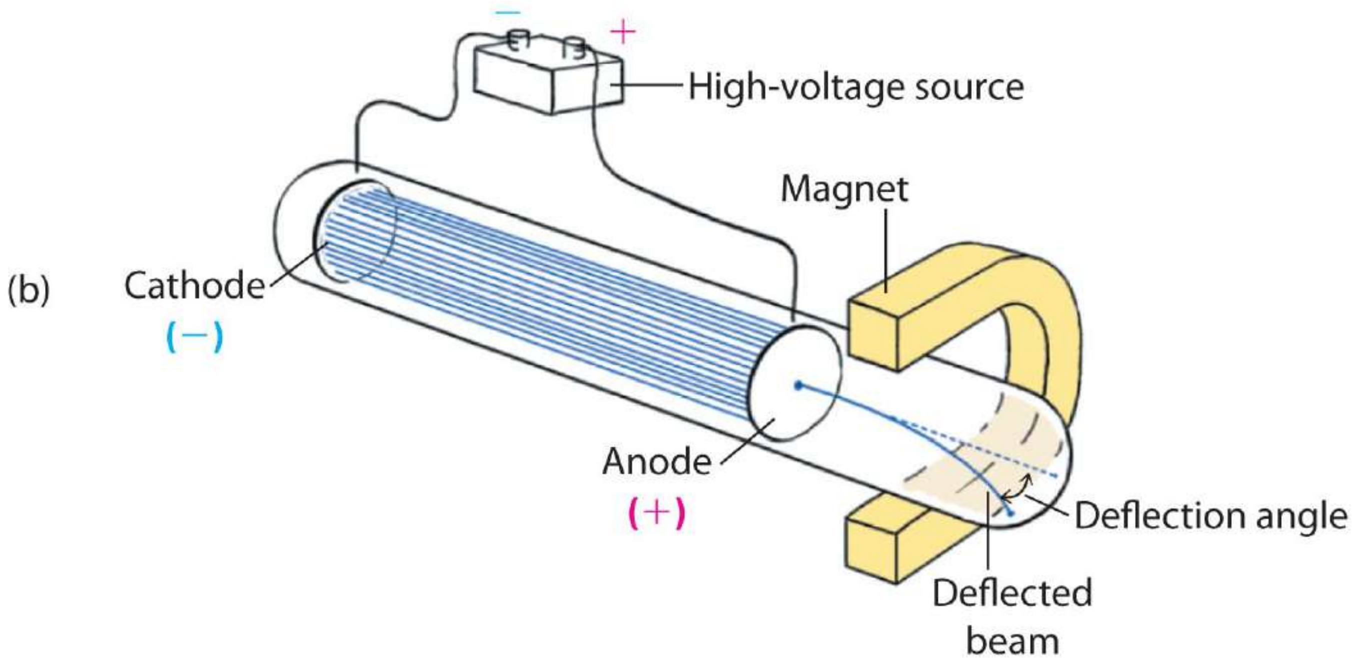
When electric charges were brought near the tube, the ray bent toward positive charges and away from negative charges.



- The ray was also deflected by the presence of a magnet.
- These findings indicated that **the cathode ray consisted of negatively charged particles.**

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Classwork

5. What property of a cathode ray is indicated when a magnet is brought near the tube?

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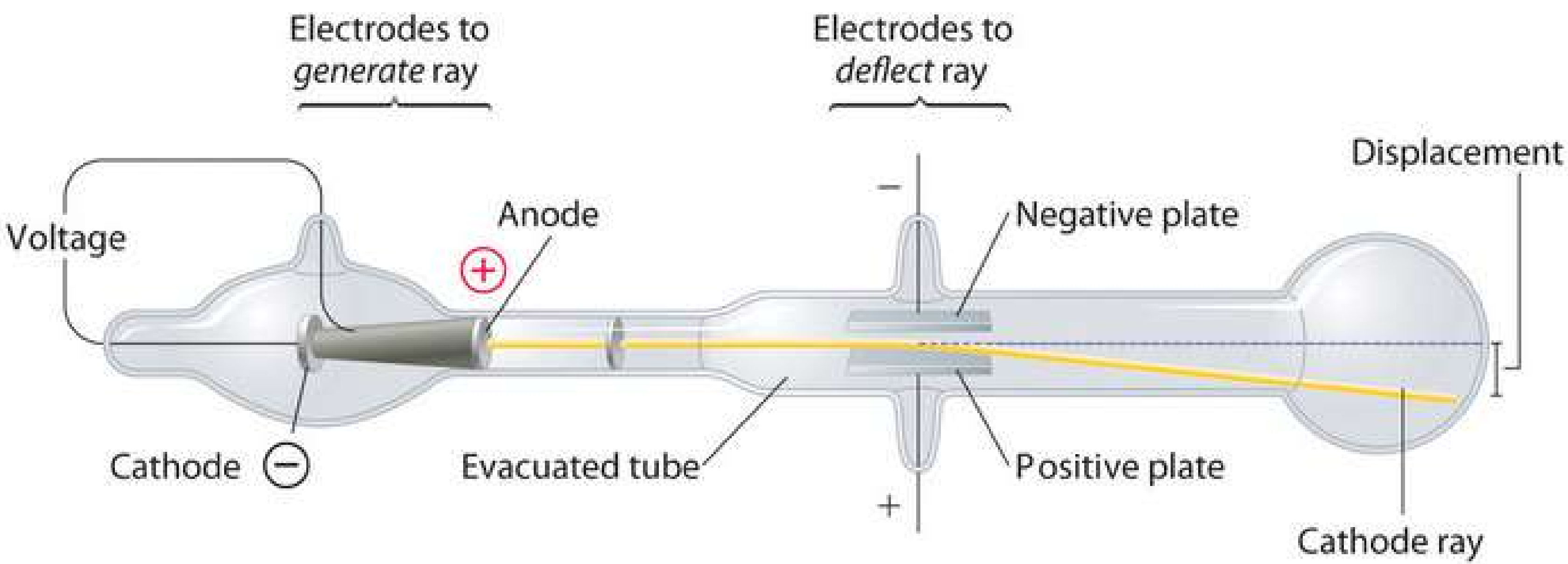
J.J. Thompson's Experiment:



- In 1897, he used a CRT to measure how much the cathode rays were deflected by electric and magnetic fields.

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- Thomson reasoned that the amount of the beam's deflection depended on the **mass** of the particles and their electrical **charge**.
  - The greater each particle's *mass*, the greater the inertia and the less the deflection.
  - The greater each particle's *charge*, the greater the force and the greater the deflection.

He discovered that the cathode rays were all lightweight, negative particles. He called them *electrons*.

→ He won the Nobel Prize for this!

→ He had measured the "charge-to-mass" ratio of the electrons. To find their mass, the charge had to be measured!

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Classwork

6. What discovery of J. J. Thomson won him the Nobel Prize?

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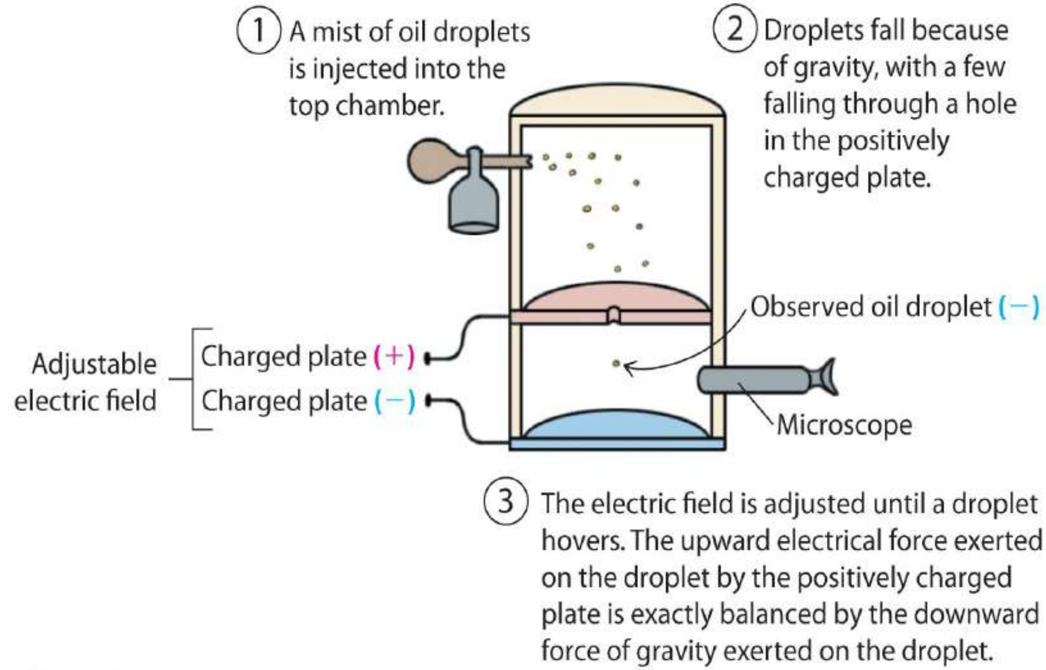
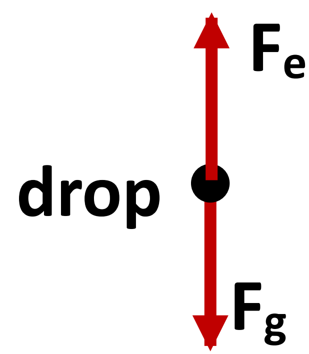
Robert Millikan

- Millikan sprayed tiny oil droplets into a chamber between electrically charged plates—into an *electric field*.



- He adjusted the field so that droplets hovered motionless, i.e., when downward force of gravity was balanced by upward electrical force.



- He used this to discover the charge on each drop.

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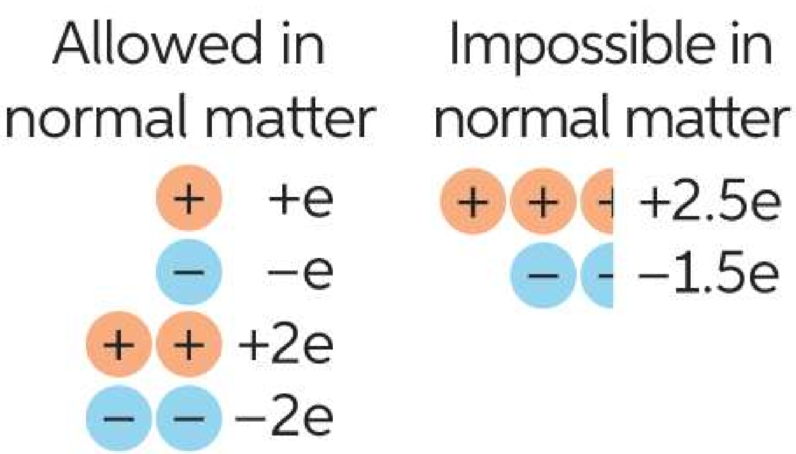
Quantized charge

Millikan found that the charge on each drop was always some multiple of a single value – the charge of each electron, e.

→ charge was quantized, and e was the quantum!

e = the smallest possible charge that can be found

= same charge as proton, but opposite.



Then, once he knew the charge, he used Thompson's charge-to-mass ratio to find the mass of the electron.

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Classwork

7. What did Robert Millikan discover about the electron?

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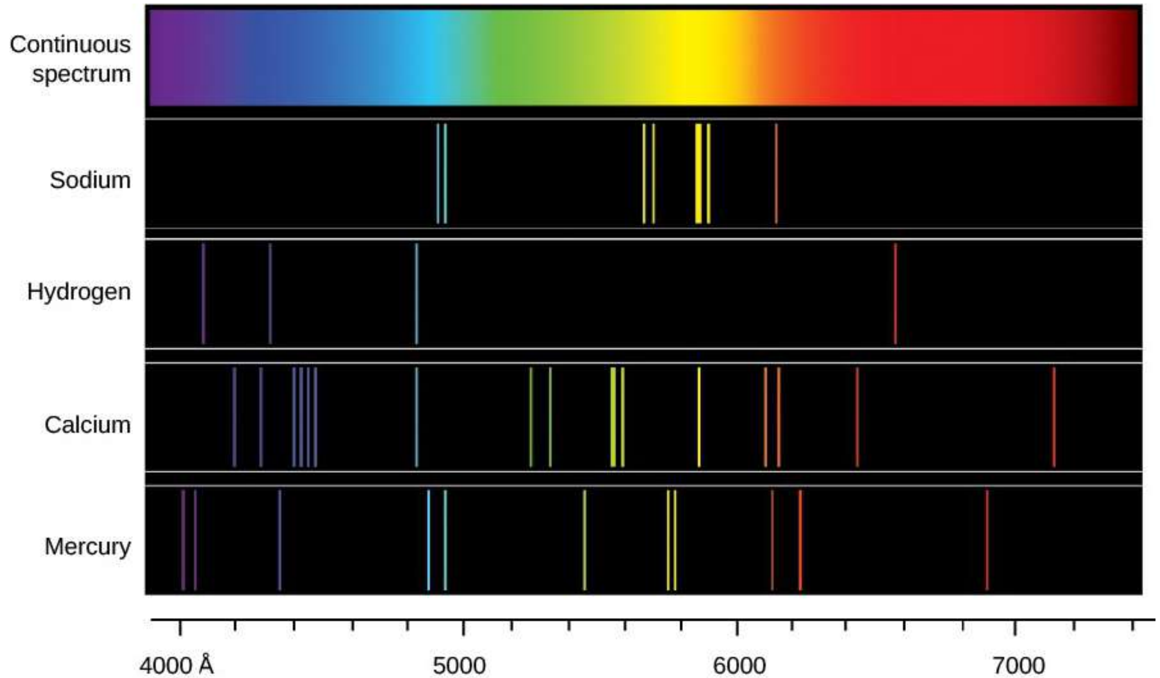
Atomic Spectra: Clues to Atomic Structure

When a gas is connected to a high voltage, it glows.

If the light is passed through slits, and through a prism (or diffraction grating), what emerges is a pattern of bright colors: **A bright-line spectrum.**

→ Each element produces a unique bright-line spectrum.



Physicists wondered:

Why isn't it a continuous spectrum?

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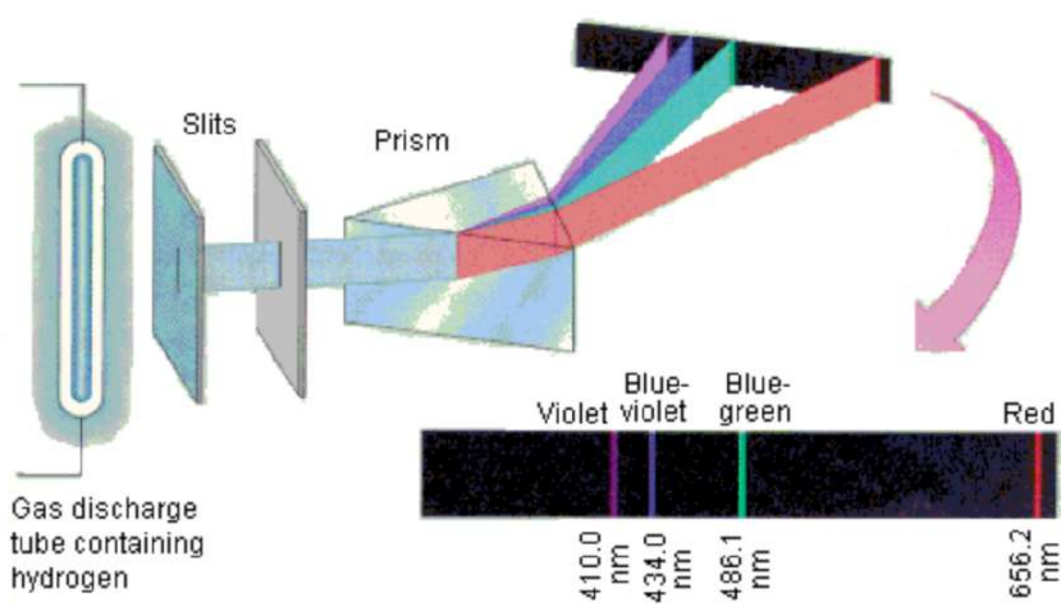
Hydrogen, H

Hydrogen is the simplest element. It has one proton and one electron. The bright-line spectrum of H is the simplest.



Understanding H would be a first step towards understanding what goes on inside of an atom.

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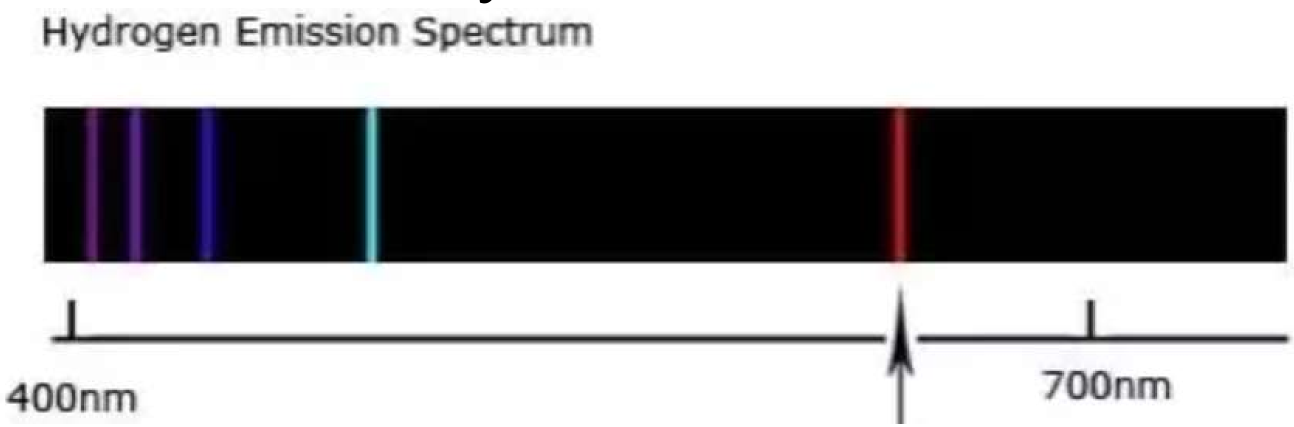
Atomic Spectra: Clues to Atomic Structure

- The hydrogen spectrum starts with a line in the red region, followed by one in the blue, then by several lines in the violet, and many in the ultraviolet.

plus more, closer lines in the UV



- Notice how the spacing between successive lines becomes smaller and smaller from the first in the red to the last in the ultraviolet, until the lines become so close that they seem to merge.

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- Jakob **Balmer** (a Swiss schoolteacher) first expressed the wavelengths of these lines in a mathematical formula.

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

n = 1, 2, 3, ….

Two other physicists, **Rydberg and Ritz**, found that if you look at the frequencies f of the lines:

Frequency of 1 line = sum of frequencies of 2 others:

$$f = f_1 + f_2$$

- What is going on inside the atom?
- → All were unable to provide a reason for these mathematical relationships.

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Classwork

8. What did Johann Jakob Balmer discover about the spectrum of hydrogen?

9. What did Johannes Rydberg and Walter Ritz discover about atomic spectra?

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Neils Bohr (1885-1962)

- Born in Copenhagen, Denmark
- Worked with J.J. Thompson
- Continued his work with Rutherford
- Won Nobel Prize in 1922 in physics
- Thought up complementarity:
  (wave particle duality)

- Helped refugees escape from the Nazis
- Worked on the Manhattan Project (atomic bomb)
- Called for international cooperation on nukes

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Bohr's Model of the Atom

- The one electron in H can "occupy" certain "stationary states," or orbitals of fixed energy.

These orbitals are numbered by a quantum number n:

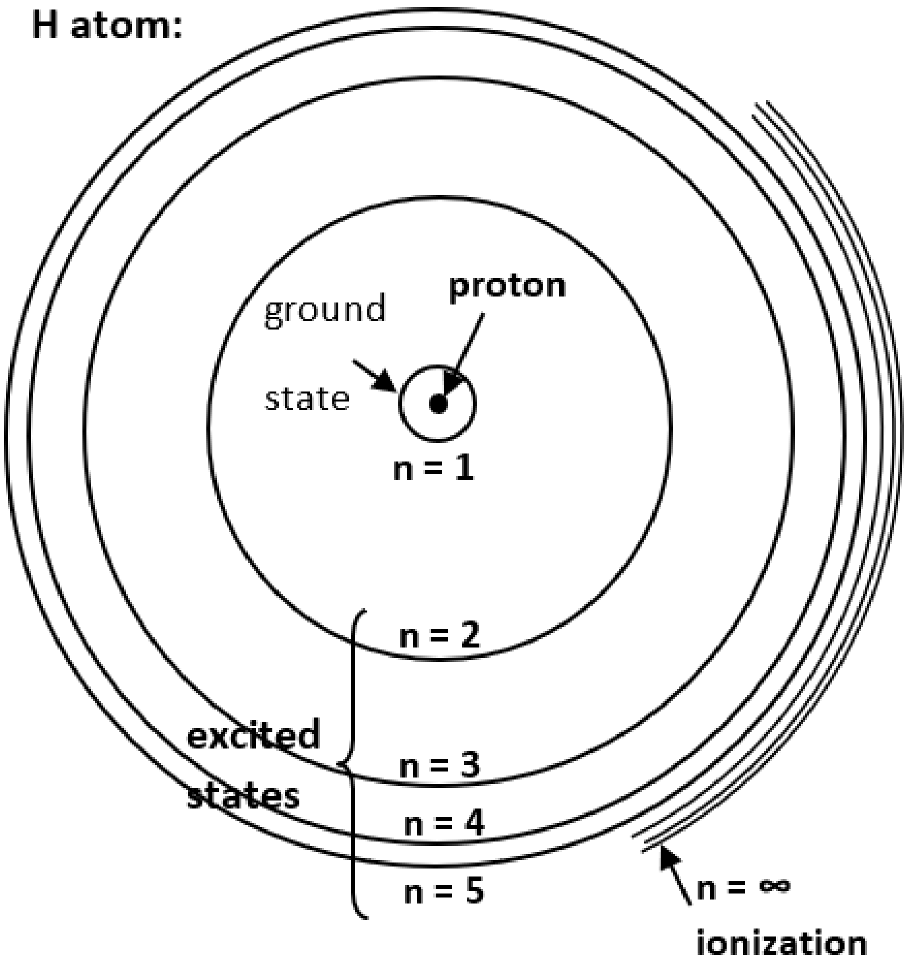
n= 1 is closest to the proton It has the lowest energy

and is the *ground state.*

n = 2, 3, 4, …are further away and higher energy

and are *excited states*

n = ∞ means the electron is ionized (removed from atom)



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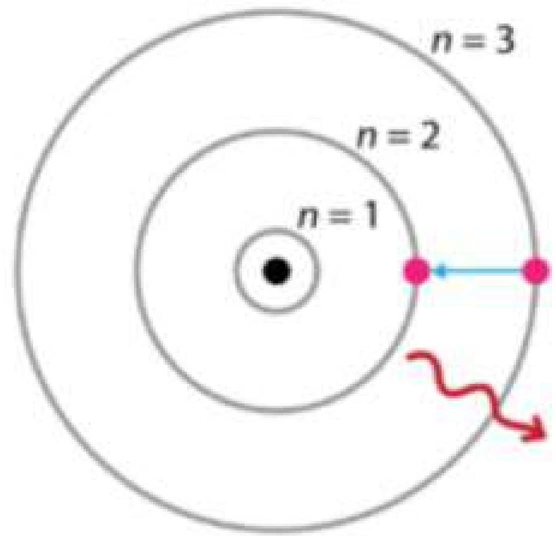
- Electrons can make quantum jumps (transitions), but only from one energy level to another.
- **Bohr postulated that:** *A photon of light is emitted when such an electron makes a jump from a higher to a lower energy state.*



The frequency f of the emitted photon is determined by:

$$\Delta E = hf$$

Here ΔE is the **difference** in the atom's energy when the electron is in the different states (orbitals).

- f = frequency of emitted photon

$\Delta E = hf$

or $f = \dfrac{\Delta E}{h}$

So a bigger ΔE means a higher f for the photon

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Classwork

10. What relationship between electron orbits and light emission did Bohr postulate?

12. What is the relationship between the energy differences of orbits in an atom and the light emitted by the atom?

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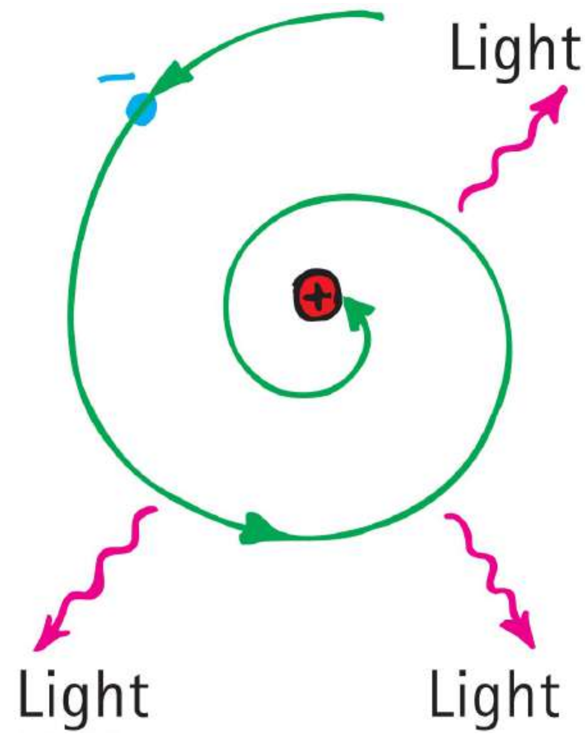
How Bohr Avoided “Classical” Problems:

- According to classical theory, an electron accelerating around its orbit should continuously emit radiation.
- This loss of energy should cause it to spiral rapidly into the nucleus.
- But this does not happen.
- Bohr avoided this problem by saying that the electron does not orbit, it just “sits there,” occupying its stationary state (or orbital).
- Radiation only occurs when an electron jumps levels.



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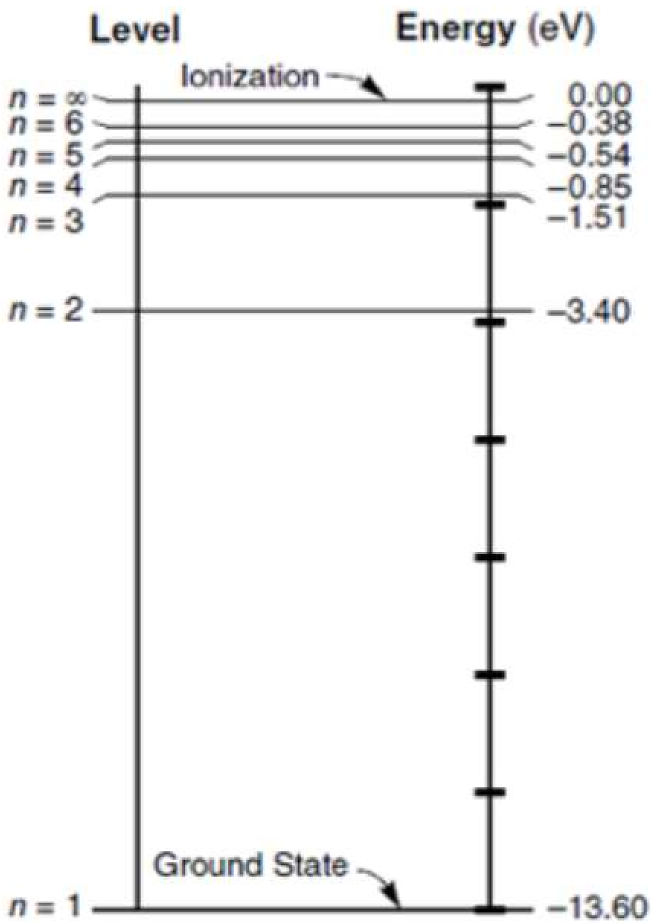
Energy levels in hydrogen H:



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Each stationary state has its own special energy.

The electron can only jump from one to another.

These levels were quantized (Bohr didn't really know why!)

Bigger jumps→ bigger E

Because E = hf, this meant:

Bigger jumps → higher f

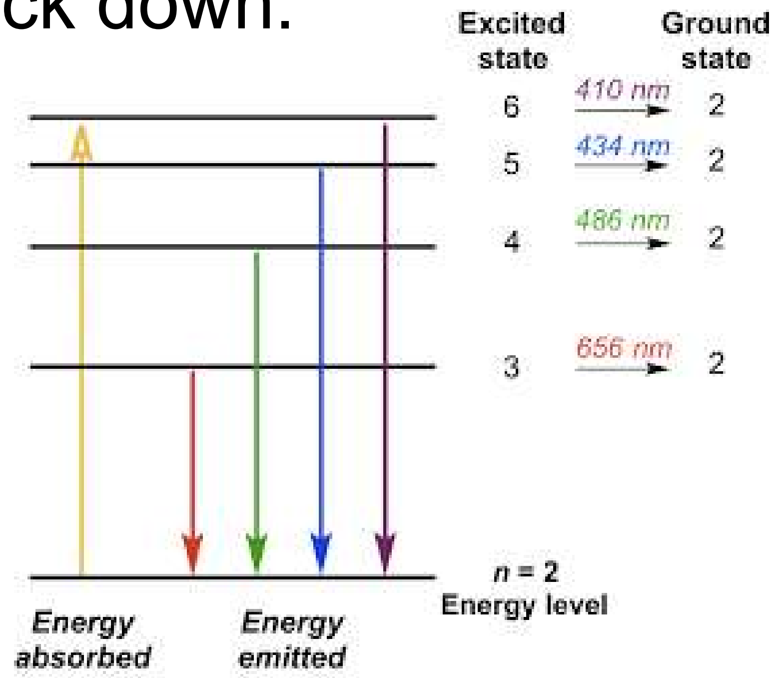
Only certain E levels were OK,

…so only certain frequencies

…and only certain wavelengths

Why the *distinct* lines in the spectrum?

- Putting the gas under high voltage excited the electron to a higher energy by absorbing energy.
- Then the electron fell back down.
- Light was Bo(h)rn!
- The jump from
- n = 3 to 2: **red** line
- n = 4 to 2: **cyan** line
- n =5 to 2: **blue** line
- n = 6 to 2: **violet** line



- Each line corresponds to one electron jump
- Energy levels are quantized, so frequencies are, too.

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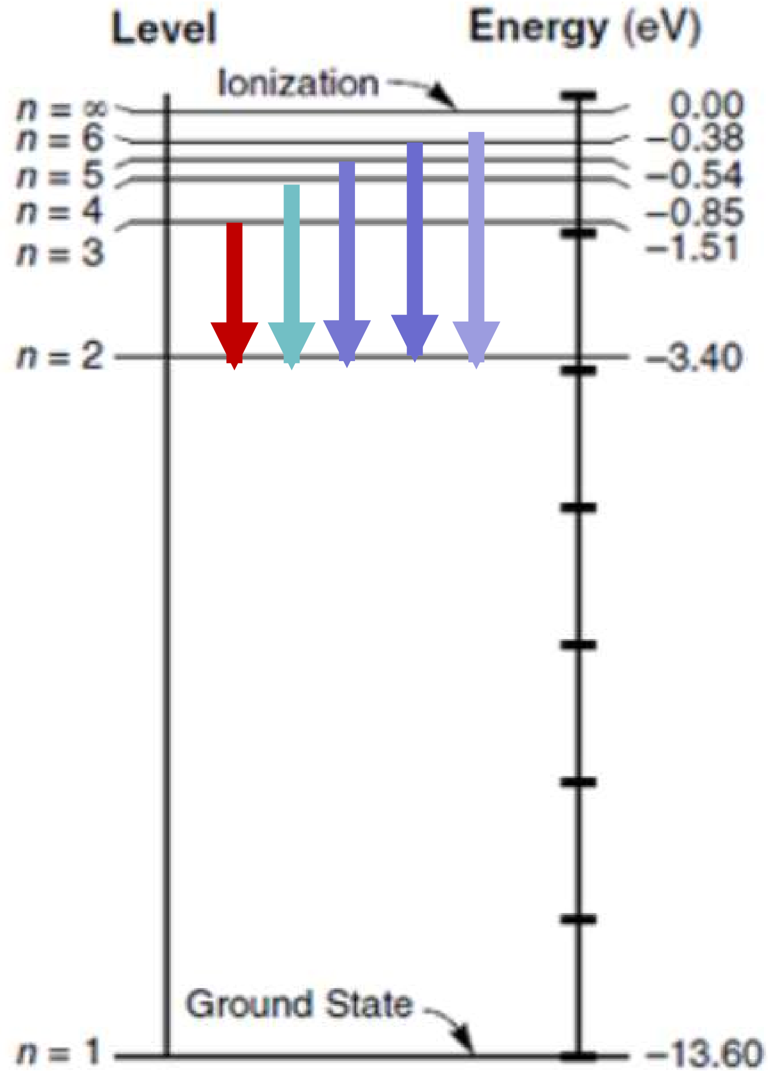
Closer and closer…

Notice how the energy levels get closer and closer near the top.

That is why the spectral lines got closer and closer, too.

Electrons were jumping from levels that were closer and closer together!



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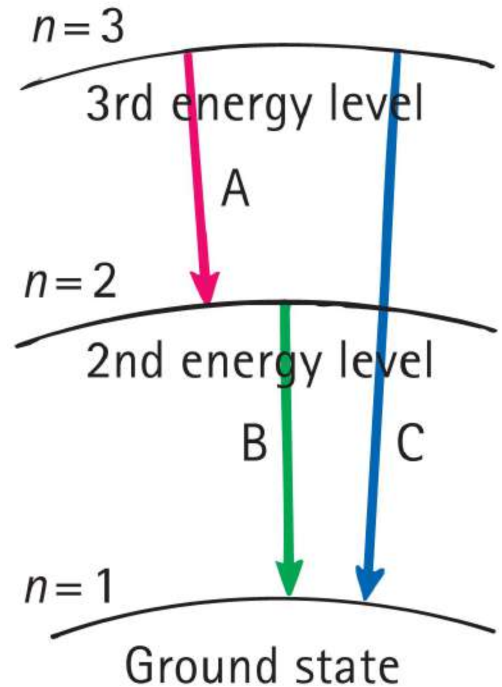
And remember that….



….Rydberg and Ritz had found: frequency of one line = sum of frequencies of 2 others:

$$f = f_1 + f_2$$

Why?

An electron in the n= 3 level can….

…. jump directly to n=1 (C):

This produces a photon with frequency $f = \Delta E_{3\text{-}1}/h$

…or it can jump to n=2 (A), and then to n=1(B):

This creates 2 photons with $f = \Delta E_{3\text{-}2}/h$ and $f = \Delta E_{2\text{-}1}/h$

Because the ΔE's add up, so do the frequencies!

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Classwork

11. According to Niels Bohr, can a single electron in one excited state give off more than one photon when it jumps to a lower energy state?

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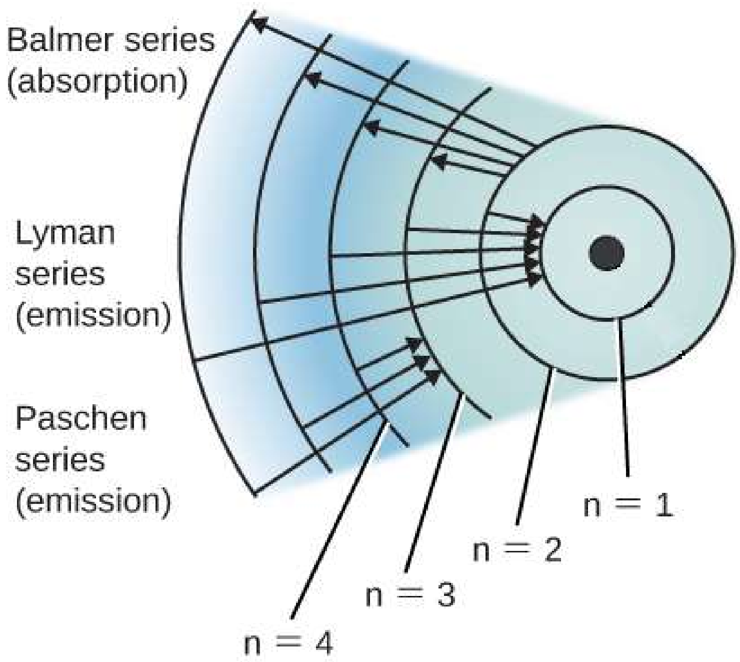
The Theory made new predictions:

All electron jumps to n = 2 created visible light.

Electron jumps to n = 1 are bigger → emit UV

Electron jumps to n= 3 are smaller → emit IR



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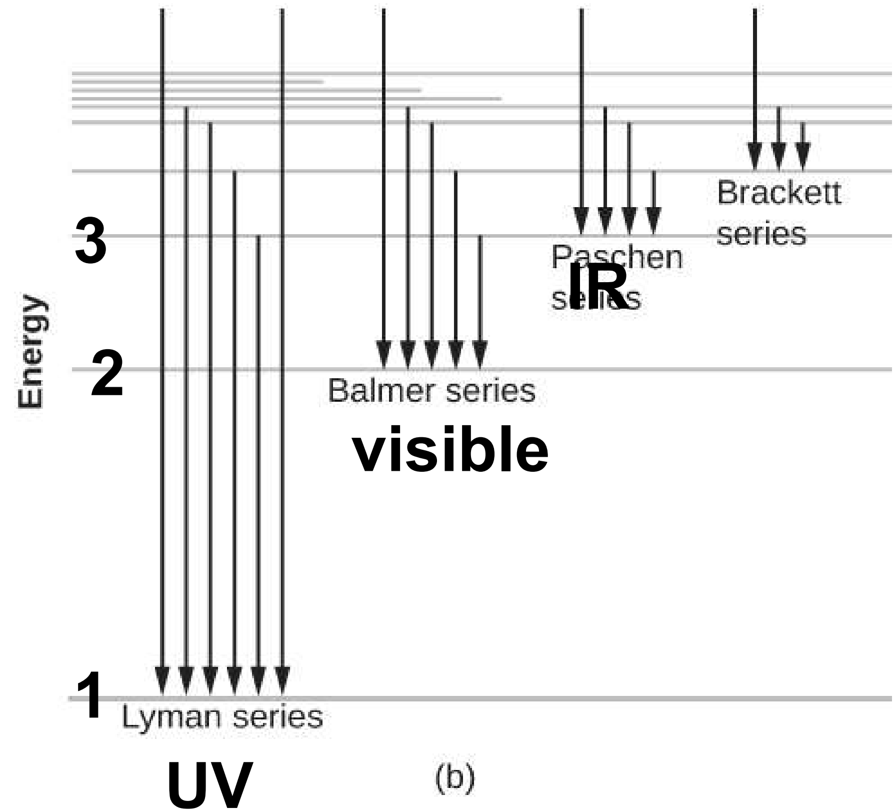

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But why were the energy levels quantized?

Why only certain rungs on the energy ladder?

That brings us back to deBroglie and his Nobel-Prize winning hypothesis that electrons can act like waves….
….next lesson!

Don't forget—Test on Chapters 31
and 32 on Friday.

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