ATOMIC STRUCTURE

RUTHERFORD SCATTERING:
Rutherford scattering experiment-> alpha particle projected on gold foil, most of them are not deflected -> most of the space in atom is empty
some are deflected-> that area is nucleus which contains +ve particles called
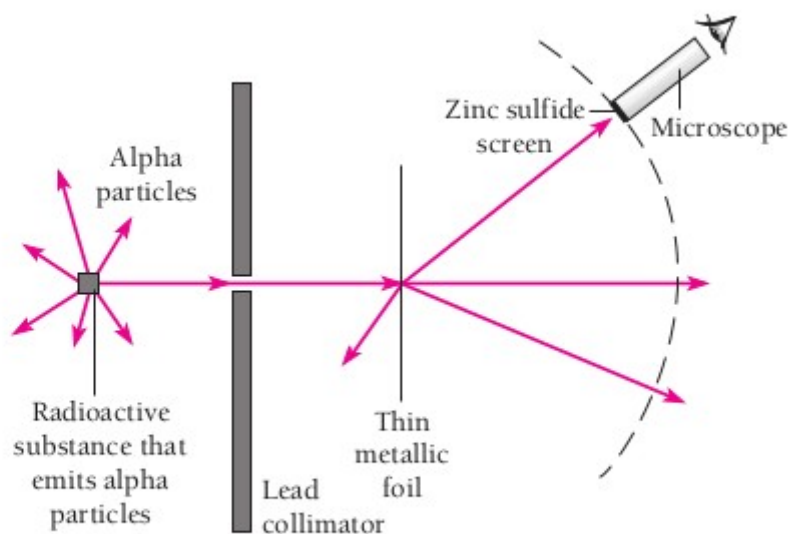contradict to Thompson.

Figure 4.2 The Rutherford scattering experiment.

RUTHERFORD MODEL:
Rutherford's model explained these results by proposing that the positive charge and mass of the atom were concentrated in a very small, dense nucleus. Alpha particles were deflected when they came close to this nucleus.

The **Rutherford model** explained that the large deflections occurred when alpha particles passed near the dense nucleus, which could exert enough force to change their path dramatically.

Failures and Limitations of the Rutherford Model:
1)As the electron loses energy, it would spiral inward toward the nucleus, eventually collapsing into it. This would mean that atoms should not be stable, but in reality, atoms are stable. This problem could not be explained by Rutherford's model.

2)No Explanation for Atomic Spectra:
3)Couldn't explain the quantization of energy level

$N(\theta)$

$N(180°)$

0° 20° 40° 60° 80° 100° 120° 140° 160° 180°
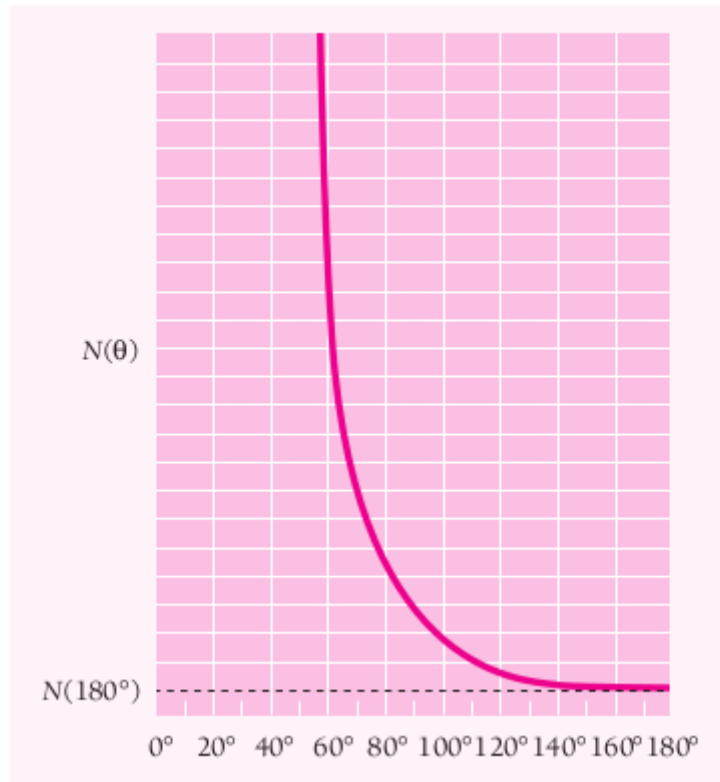
Rutherford Scattering Formula

The formula that Rutherford obtained for alpha particle scattering by a thin foil on the basis of the nuclear model of the atom is

Rutherford scattering formula
$$N(\theta) = \frac{N_i n t Z^2 e^4}{(8\pi\epsilon_0)^2 r^2 \, KE^2 \, \sin^4(\theta/2)}$$
(4.1)

This formula is derived in the Appendix to this chapter. The symbols in Eq. (4.1) have the following meanings:

$N(\theta)$ = number of alpha particles per unit area that reach the screen at a scattering angle of θ
N_i = total number of alpha particles that reach the screen
n = number of atoms per unit volume in the foil
Z = atomic number of the foil atoms
r = distance of the screen from the foil
KE = kinetic energy of the alpha particles
t = foil thickness

1. Spectral Lines and Atomic Behavior:

- **Spectral lines** are specific wavelengths of light that are either emitted or absorbed by atoms. Each element has its own unique set of spectral lines, which act like a "fingerprint" for that element.
- When a gas is **excited** (for example, by passing electricity through it), the atoms emit light at certain specific wavelengths, creating a pattern called an **emission spectrum**. This is important because it shows that atoms do not emit light at all wavelengths randomly, but only at very specific ones.
- On the other hand, when white light (which contains all wavelengths) passes through a gas, the gas **absorbs** certain wavelengths of light, leaving dark lines in the spectrum. This is called an **absorption spectrum**.

2. Emission vs. Absorption Spectra:s

- **Emission spectra**: These consist of bright lines on a dark background and are produced when atoms in a gas release energy and emit light at specific wavelengths.
- **Absorption spectra**: These consist of dark lines on a bright background and are created when atoms absorb light at certain wavelengths, leaving gaps in the spectrum.

Spectral Series

A century ago the wavelengths in the spectrum of an element were found to fall into sets called **spectral series**. The first such series was discovered by J. J. Balmer in 1885 in the course of a study of the visible part of the hydrogen spectrum. Figure 4.10 shows the **Balmer series.** The line with the longest wavelength, 656.3 nm, is designated H_α, the next, whose wavelength is 486.3 nm, is designated H_β, and so on. As the wave-length decreases, the lines are found closer together and weaker in intensity until the **series limit** at 364.6 nm is reached, beyond which there are no further separate lines but only a faint continuous spectrum. Balmer's formula for the wavelengths of this series is

Balmer
$$\frac{1}{\lambda} = R\left(\frac{1}{2^2} - \frac{1}{n^2}\right) \quad n = 3, 4, 5, \ldots \qquad (4.6)$$

The quantity R, known as the **Rydberg constant,** has the value

Rydberg constant
$$R = 1.097 \times 10^7 \text{ m}^{-1} = 0.01097 \text{ nm}^{-1}$$

The H_α line corresponds to $n = 3$, the H_β line to $n = 4$, and so on. The series limit corresponds to $n = \infty$, so that it occurs at a wavelength of $4/R$, in agreement with experiment.

The Balmer series contains wavelengths in the visible portion of the hydrogen spectrum. The spectral lines of hydrogen in the ultraviolet and infrared regions fall into several other series. In the ultraviolet the **Lyman series** contains the wavelengths given by the formula
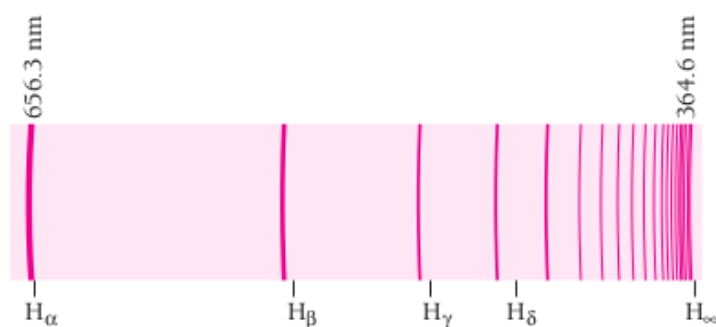

Figure 4.10 The Balmer series of hydrogen. The H_α line is red, the H_β line is blue, the H_γ and H_δ lines are violet, and the other lines are in the near ultraviolet.

Bohr's Atomic Model

Niels Bohr proposed his atomic model in 1913, refining the earlier Rutherford model. Bohr introduced the concept of **quantized energy levels** for electrons orbiting the nucleus, which helped address some of the limitations of Rutherford's model. Here's a detailed explanation of Bohr's atomic model:

Key Features of Bohr's Atomic Model:

1. **Electron Orbits are Quantized:**

   - Bohr postulated that electrons revolve around the nucleus in specific circular orbits or energy levels without radiating energy. These orbits are called **stationary orbits**.
   - Each orbit corresponds to a specific energy level, and the electron remains stable in these orbits as long as it doesn't absorb or emit energy.

2. **Quantization of Angular Momentum:**

   - The angular momentum of the electron in these orbits is quantized, i.e., it can only take certain discrete values. This is expressed by: $L = n\hbar = n \times 2\pi\text{th}$ where L is the angular momentum, n is the principal quantum number (a positive integer), h is Planck's constant, and $\hbar$ is the reduced Planck constant.

3. **Energy Levels:**

   - Each orbit corresponds to a fixed energy level, and the energy of an electron depends on the orbit it occupies. Electrons in higher orbits have more energy.
   - When an electron jumps from one orbit to another, it either **absorbs** energy (if it moves to a higher orbit) or **emits** energy (if it moves to a lower orbit), which results in the emission or absorption of electromagnetic radiation (like photons). The energy of the emitted or absorbed photon is given by: $\Delta E = E_2 - E_1 = h\nu$

An electron can circle a nucleus only if its orbit contains an integral number of de Broglie wavelengths.

Condition for orbit
stability
n*lambda $= 2\pi r_n$
n $= 1, 2, 3, \ldots$
(4.12)
where r_n designates the radius of the orbit that contain n wavelengths. The integer n
is called the quantum number of the orbit. Substituting for #, the electron wavelength

Orbital radii in Bohr atom
$$r_n = \frac{n^2 h^2 \epsilon_0}{\pi m e^2} \qquad n = 1, 2, 3, \ldots \tag{4.13}$$

The radius of the innermost orbit is customarily called the **Bohr radius** of the hydrogen atom and is denoted by the symbol a_0:

Bohr radius
$$a_0 = r_1 = 5.292 \times 10^{-11} \, \text{m}$$

The other radii are given in terms of a_0 by the formula

$$r_n = n^2 a_0 \tag{4.14}$$

4.5 ENERGY LEVELS AND SPECTRA

A photon is emitted when an electron jumps from one energy level to a lower level

The various permitted orbits involve different electron energies. The electron energy E_n is given in terms of the orbit radius r_n by Eq. (4.5) as

$$E_n = -\frac{e^2}{8\pi\epsilon_0 r_n}$$

Substituting for r_n from Eq (4.13), we see that

Energy levels
$$E_n = -\frac{me^4}{8\epsilon_0^2 h^2}\left(\frac{1}{n^2}\right) = \frac{E_1}{n^2} \qquad n = 1, 2, 3, \ldots \tag{4.15}$$

$$E_1 = -2.18 \times 10^{-18} \, \text{J} = -13.6 \, \text{eV}$$

The energies specified by Eq. (4.15) are called the **energy levels** of the hydrogen atom and are plotted in Fig. 4.15. These levels are all negative, which signifies that the electron does not have enough energy to escape from the nucleus. An atomic electron can have only these energies and no others. An analogy might be a person on a ladder, who can stand only on its steps and not in between.

If the quantum number of the initial (higher-energy) state is n_i and the quantum number of the final (lower-energy) state is n_f, we are asserting that

$$\text{Initial energy} - \text{final energy} = \text{photon energy}$$

$$E_i - E_f = h\nu \qquad (4.16)$$

where ν is the frequency of the emitted photon. From Eq. (4.15) we have

$$E_i - E_f = E_1\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right) = -E_1\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right)$$

We recall that E_1 is a negative quantity (-13.6 eV, in fact), so $-E_1$ is a positive quantity. The frequency of the photon released in this transition is therefore

$$\nu = \frac{E_i - E_f}{h} = -\frac{E_1}{h}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \qquad (4.17)$$

Since $\lambda = c/\nu$, $1/\lambda = \nu/c$ and

Hydrogen spectrum

$$\frac{1}{\lambda} = -\frac{E_1}{ch}\left(\frac{1}{n_f^2} - \frac{1}{n_i^2}\right) \qquad (4.18)$$

FREQUENCY;

$$v = \frac{e}{\sqrt{4\pi\epsilon_o mr}}$$

according to Eq. (4.4), where r is the radius of its orbit. Hence the frequency of revolution f of the electron is

$$f = \frac{\text{electron speed}}{\text{orbit circumference}} = \frac{v}{2\pi r} = \frac{e}{2\pi\sqrt{4\pi\epsilon_0 mr^3}}$$

The radius r_n of a stable orbit is given in terms of its quantum number n by Eq. (4.13) as

$$r_n = \frac{n^2 h^2 \epsilon_0}{\pi m e^2}$$

Under what circumstances should the Bohr atom behave classically? If the electron
orbit is so large that we might be able to measure it directly, quantum effects ought
not to dominate. An orbit 0.01 mm across, for instance, meets this specification. As
we found in Example 4.3, its quantum number is n =435.
The requirement that quantum physics give the same results as classical physics in
the limit of large quantum numbers was called by Bohr the correspondence
principle. It has played an important role in the development of the quantum theory of
matter.
Bohr himself used the correspondence principle in reverse, so to speak, to look for
the condition for orbit stability. Starting from Eq. (4.19) he was able to show that stable
orbits must have electron orbital angular momenta of
   mvr=nh/2*pi

4.7
NUCLEAR MOTION
The nuclear mass affects the wavelengths of spectral lines

Thus far we have been assuming that the hydrogen nucleus (a proton) remains
stationary while the orbital electron revolves around it. What must actually
happen, of
course, is that both nucleus and electron revolve around their common center of
mass,
which is very close to the nucleus because the nuclear mass is much greater
than that
of the electron (Fig. 4.17). A system of this kind is equivalent to a single
particle of
mass m# that revolves around the position of the heavier particle.

demonstrated in Sec. 8.6.) If m is the electron mass and M the nuclear mass, then m'
is given by

Reduced mass $$m' = \frac{mM}{m + M} \qquad (4.22)$$

The quantity m' is called the **reduced mass** of the electron because its value is less
than m.

To take into account the motion of the nucleus in the hydrogen atom, then, all we
need do is replace the electron with a particle of mass m'. The energy levels of the
atom then become

**Energy levels
corrected for
nuclear motion** $$E'_n = -\frac{m'e^4}{8\epsilon_0^2 h^2}\left(\frac{1}{n^2}\right) = \left(\frac{m'}{m}\right)\left(\frac{E_1}{n^2}\right) \qquad (4.23)$$

Owing to motion of the nucleus, all the energy levels of hydrogen are changed by the
fraction

$$\frac{m'}{m} = \frac{M}{M + m} = 0.99945$$

There are two main ways in which an atom can be excited to an energy above its ground state and thereby become able to radiate. One of these ways is by a collision with another particle in which part of their joint kinetic energy is absorbed by the atom. Such an excited atom will return to its ground state in an average of 10^{-8} s by emitting one or more photons.

Another excitation mechanism is involved when an atom absorbs a photon of light whose energy is just the right amount to raise the atom to a higher energy level. For example, a photon of wavelength 121.7 nm is emitted when a hydrogen atom in the
n=2 state drops to the n=1 state. Absorbing a photon of wavelength 121.7 nm by a hydrogen atom initially in the n=1 state will therefore bring it up to the n=2 state (Fig. 4.19).

The **Franck-Hertz experiment** was conducted by James Franck and Gustav Hertz in 1914 and played a key role in confirming the quantum nature of atoms and their energy levels, supporting the Bohr model of the atom.

Purpose of the Experiment:

The experiment aimed to investigate how electrons interact with mercury atoms and how energy is exchanged during collisions between free electrons and atoms.

Experimental Setup:

- **Electron tube:** The experiment used a glass tube filled with mercury vapor.
- **Cathode and Anode:** Electrons were emitted from a heated cathode and accelerated towards a positively charged anode, creating a current.
- **Grid:** A grid between the cathode and anode allowed the acceleration of electrons by controlling their energy.
- **Variable Voltage:** The accelerating voltage between the grid and the anode was gradually increased to control the energy of the electrons.

Key Observations:

As the accelerating voltage increased, Franck and Hertz observed periodic drops in the current flowing through the tube. These drops occurred at specific intervals of energy (around 4.9 eV), corresponding to a characteristic energy of the mercury atoms.

Interpretation:

- The experiment showed that when the energy of the electrons reached 4.9 eV, they transferred this energy to the mercury atoms, exciting them to a higher energy state.
- The drop in current occurred because the electrons lost their kinetic energy after the inelastic collisions with mercury atoms.

- When the energy of the electrons was below 4.9 eV, the collisions were elastic, meaning no energy was transferred to the mercury atoms.

Quantum Significance:

The experiment demonstrated that atoms have discrete energy levels, validating the idea that electrons could only transfer specific amounts of energy to atoms, consistent with the quantum theory.