Atomic and Nuclear Physics

Topic 7.1: The Atom

The Atom

The earliest references to the concept of atoms date back to ancient India in the 6th century BCE.

In approximately 450 BCE, Democritus coined the term átomos (Greek: ἄτομος), which means "uncuttable" or "the smallest indivisible particle of matter", i.e., something that cannot be divided further.

Although the term initially referred not only to matter but also to spiritual elements, it was later adopted in when modern Science started to develop.

The Atom

• In 1661, natural philosopher Robert Boyle suggested that matter was composed of various combinations of different "corpuscles" or atoms, rather than the classical elements of air, earth, fire and water.

Robert Boyle (1627-1691)

• In 1789 the term element was defined by the French nobleman and scientific researcher Antoine Lavoisier to mean basic substances that could not be further broken down by the methods of chemistry.

Antoine Lavoisier

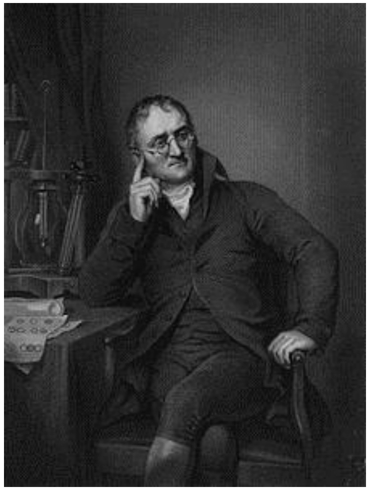
(1743-1794)

A replica of Lavoisier's laboratory at the Deutsches Museum in Munich, Germany. The large lens in the center of the picture was used to focus sunlight in order to ignite samples during combustion studies.

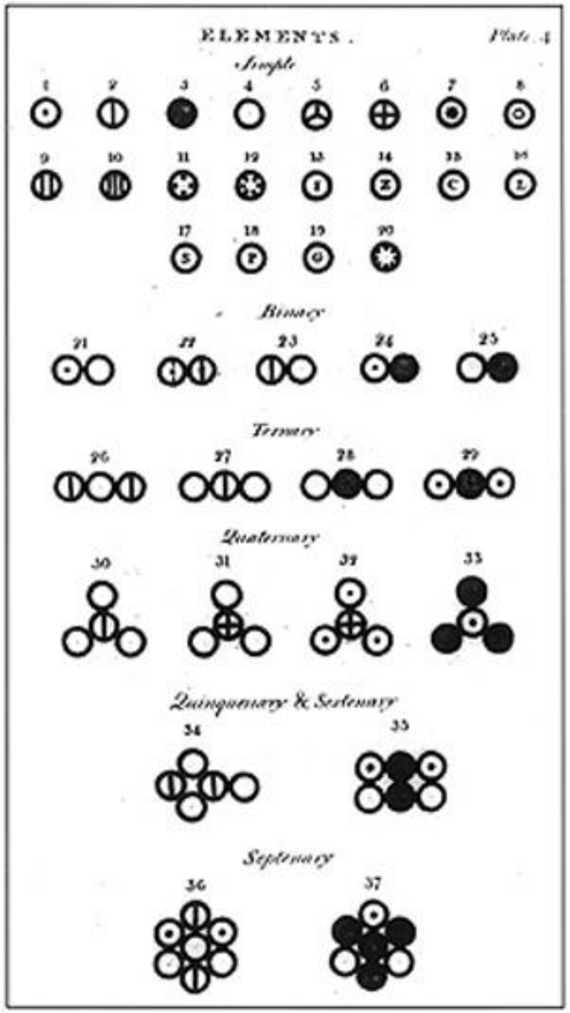
The Atom

• In 1803, English instructor and natural philosopher John Dalton proposed that each element consists of atoms of a single, unique type, and that these atoms can join together to form chemical compounds.

•Dalton used the concept of atoms to explain why elements always react in a ratio of small whole numbers - the law of multiple proportions - and why certain gases dissolve better in water than others.



John Dalton (1766-1844)



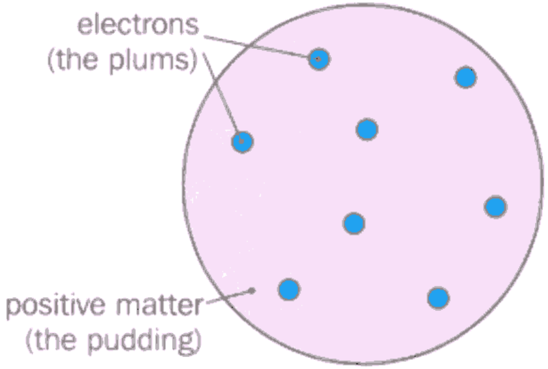
The Atom

• In 1897, the physicist J. J. Thomson, through his work on cathode rays, discovered the electron and its subatomic nature, which destroyed the concept of atoms as being indivisible units.

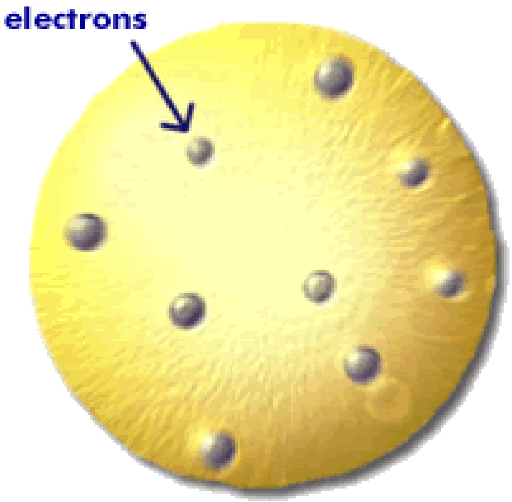
J. J. Thomson

(1856-1940)

• Thomson believed that the electrons were distributed throughout the atom, with their charge balanced by the presence of a uniform sea of positive charge (the plum pudding model).



Thomson's 'plum-pudding' model



The Atom

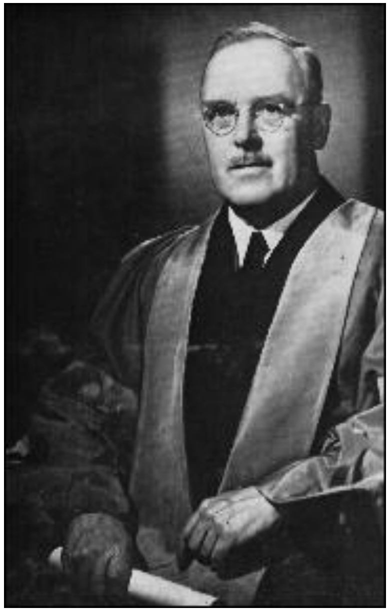
• However, in 1909, Geiger and Marsden, two researchers under the direction of physicist Ernest Rutherford, bombarded a sheet of gold foil with helium ions and discovered that a small percentage were deflected through much larger angles than was predicted using Thomson's proposal.

Ernest Rutherford

(1871-1937)

Hans Geiger

(1882-1945)

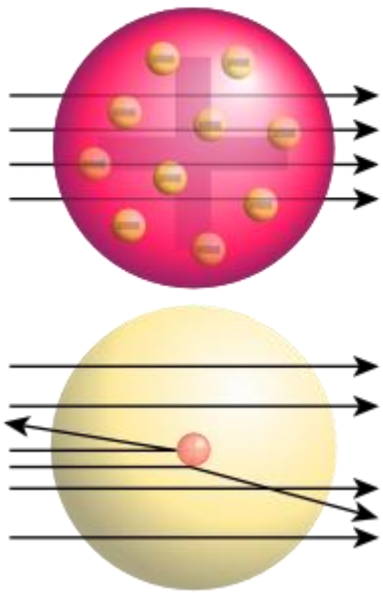


Ernest Marsden

(1889-1970)

The Atom

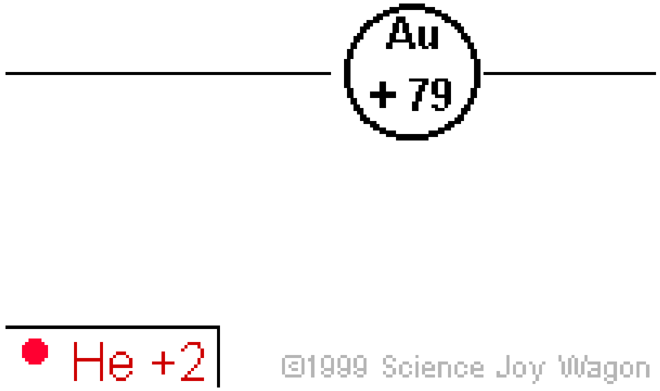
• Rutherford interpreted the gold foil experiment as suggesting that the positive charge of an atom and most of its mass was concentrated in a nucleus at the centre of the atom (the Rutherford model), with the electrons orbiting it like planets around a sun.



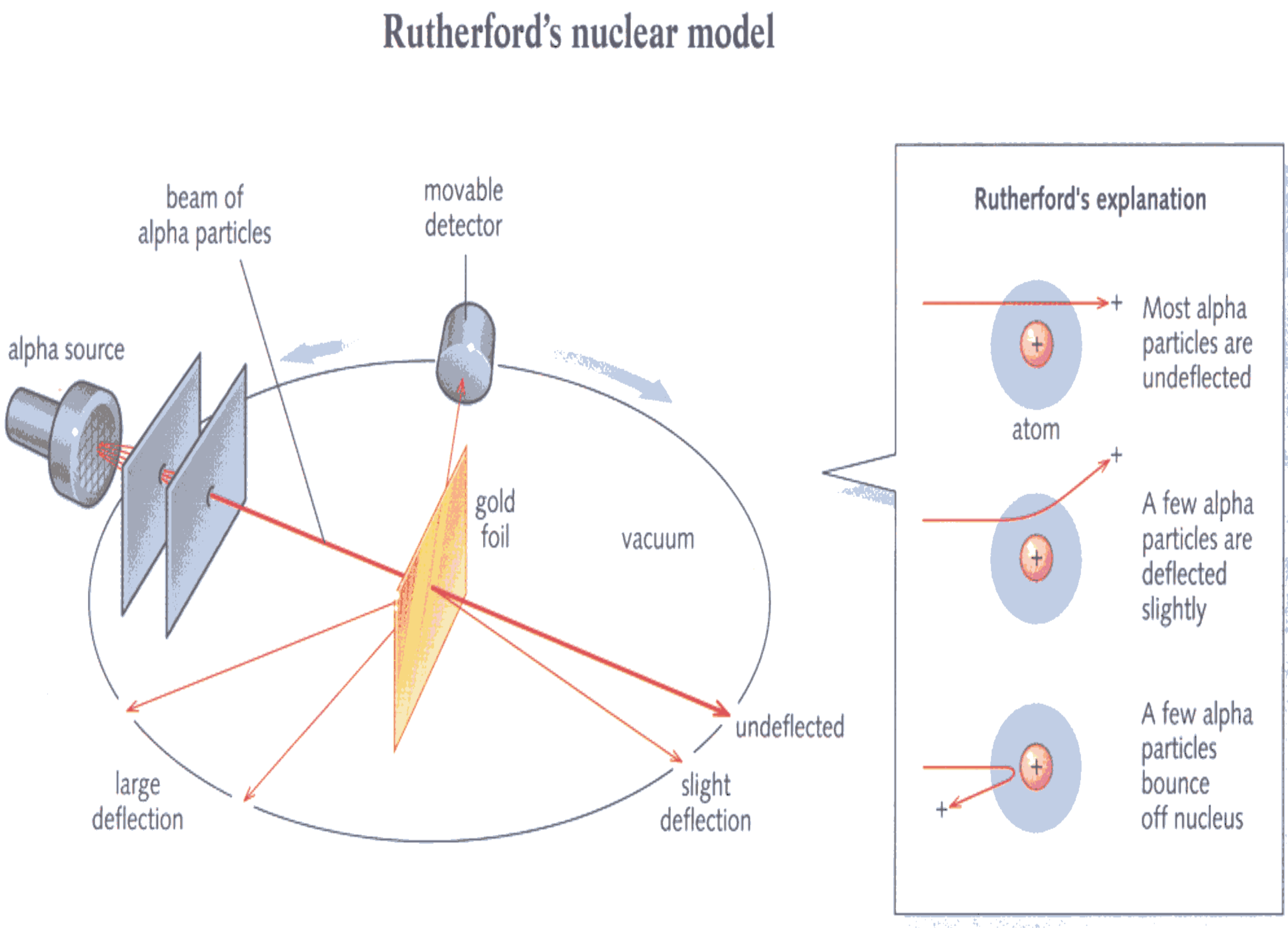
→ Expected results: alpha particles passing through the plum pudding model of the atom undisturbed.

→ Observed results: a small portion of the particles were deflected, indicating a small, concentrated positive charge.

•Positively charged helium ions passing close to this dense nucleus would then be deflected away at much sharper angles



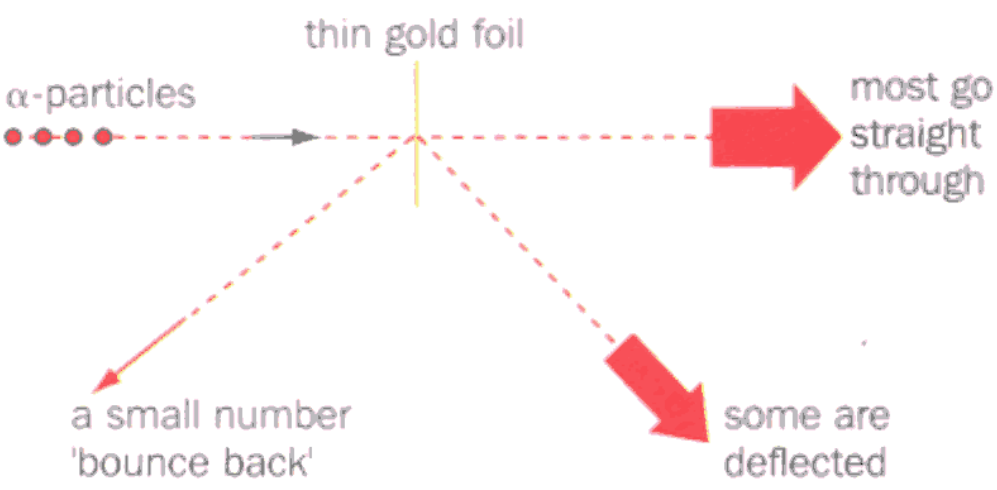
Rutherford's (or Geiger-Marsden) gold foil experiment

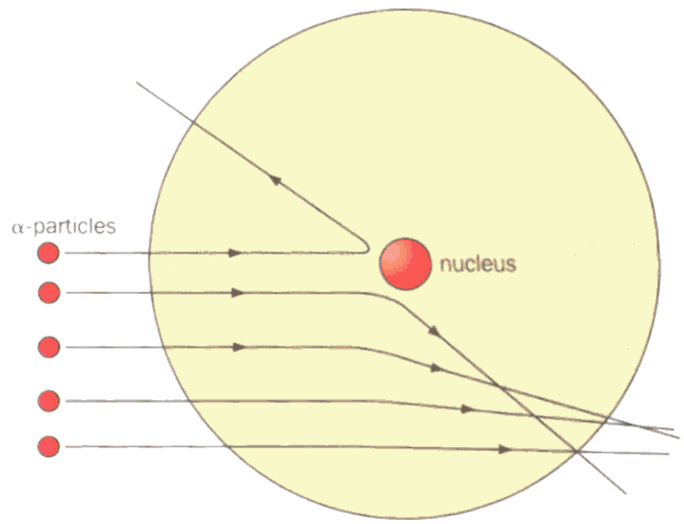


The Atom

• Most of the α-particles passed straight through the foil, but to Rutherford's surprise a few were scattered back towards the source.

• Rutherford said that this was rather like firing a gun at tissue paper and finding that some bullets bounce back towards you!



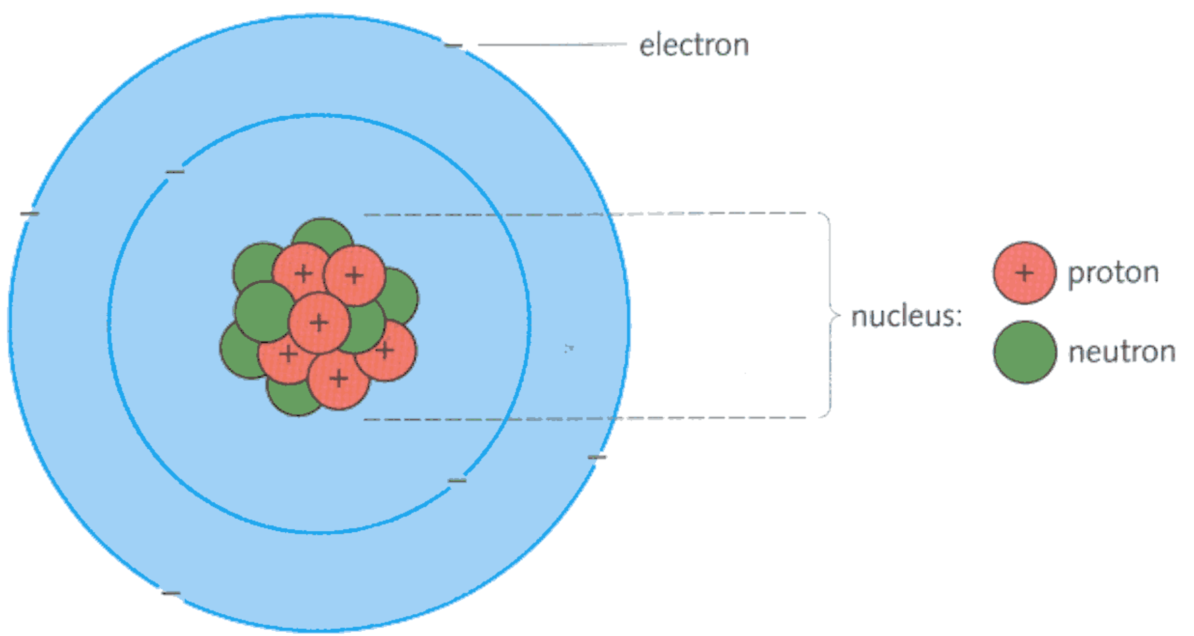


Consequences of the Rutherford's experiment

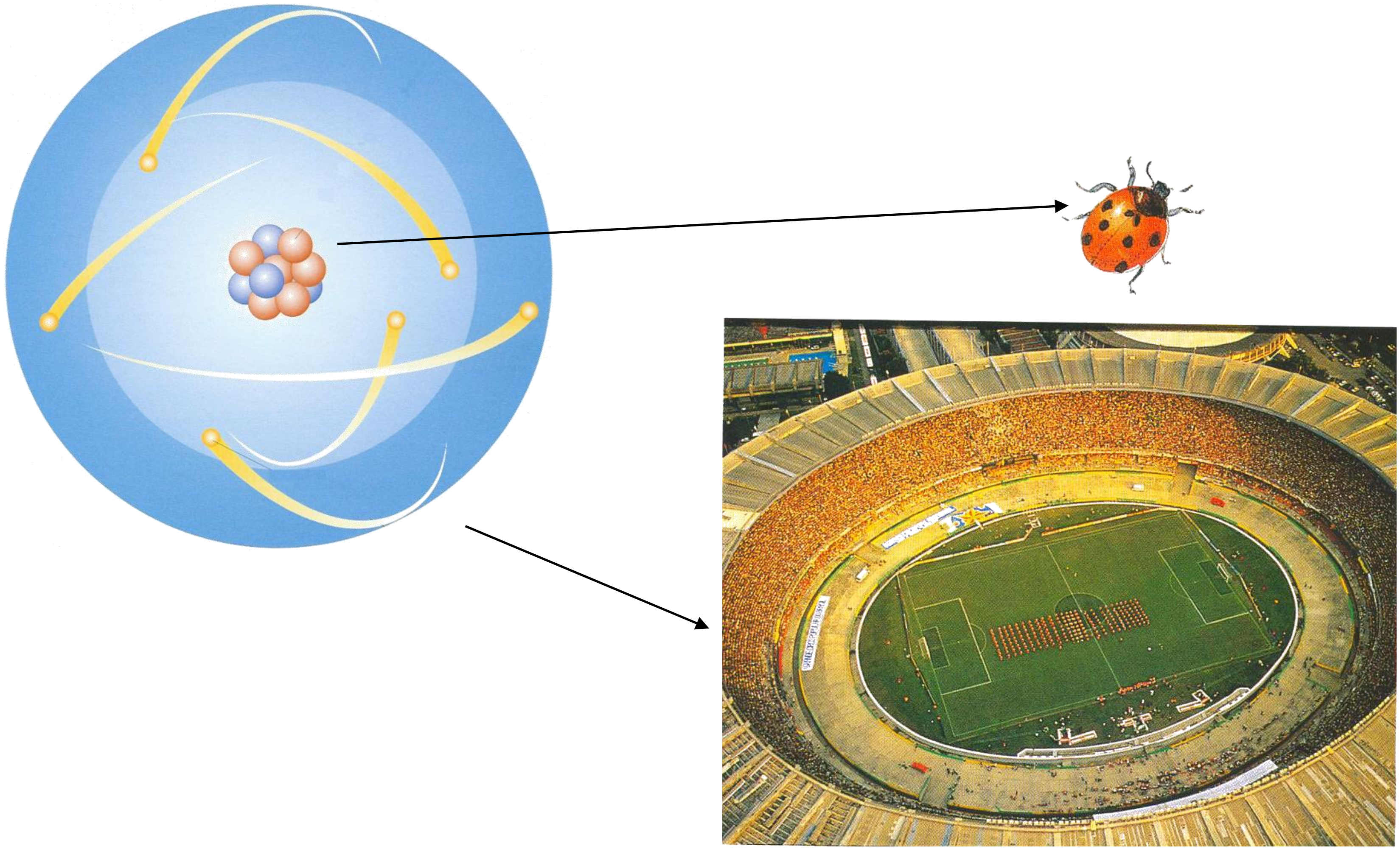
• All of an atom's positive charge and most of its mass is concentrated in a tiny core. Rutherford called this the nucleus.

• The electrons surround the nucleus, but they are at relatively large distances from it.

• The atom is mainly empty space!



Relative size of the nucleus and electric cloud



Rutherford's model of the atom

- Can we use Rutherford's model of the atom to explain the α-particle scattering?

- The concentrated positive charge produces an electric field which is very strong close to the nucleus.

- The closer the path of the α-particle to the nucleus, the greater the electrostatic repulsion and the greater the deflection.

- Most α-particles are hardly deflected because they are far away from the nucleus and the field is too weak to repel them much.

- The electrons do not deflect the α-particles because the effect of their negative charge is spread thinly throughout the atom.

Rutherford's model of the atom

• Using this model Rutherford calculated that the diameter of the gold nucleus could not be larger than 10^{-15} m.

• Other experiments confirmed the existence of a nucleus inside the atom – a small, massive object carrying the positive charge of the atom.

• The force that would keep the electrons in orbit was the electrical force between electrons and the positive nuclear charge – Coulomb's force.

Can you see any problems here?

Rutherford's model of the atom

• According to the electromagnetic theory an accelerated charge would radiate electromagnetic waves and thus lose energy.

• The electrons move in circular paths around the nucleus. But if they radiate and lose energy, then they would fall towards the nucleus.

• because of this, Rutherford's model cannot explain way matter is stable, i.e., why atoms exist.

The Bohr model

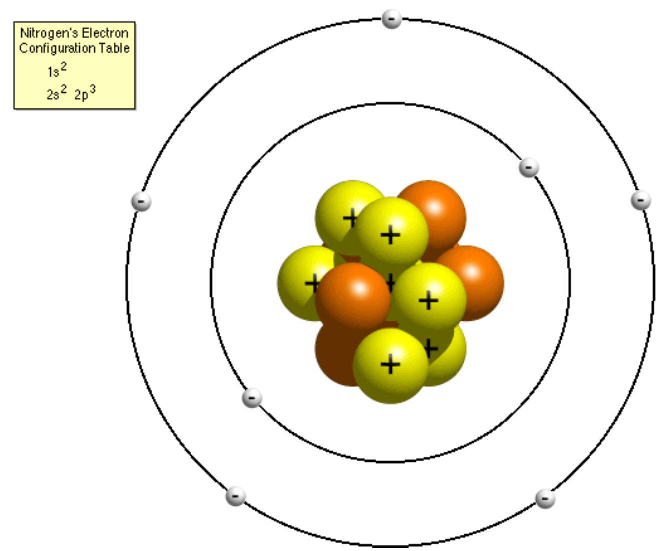
Bohr's model of the atom

• The first attempt to solve the problem with Rutherford's model came from Niels Bohr, a Danish physicist, in 1911.

• Bohr revised Rutherford's model by suggesting that the electrons were confined to clearly defined orbits, and could jump between these, but could not freely spiral inward or outward in intermediate states.

• An electron must absorb or emit specific amounts of energy to transition between these fixed orbits.

Niels Bohr (1885-1962)

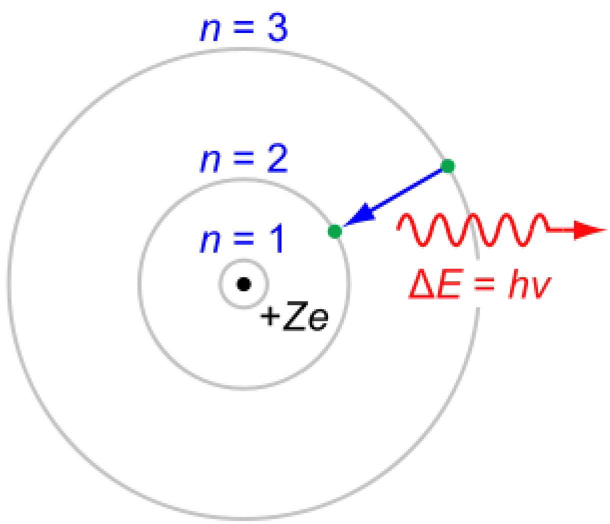


Bohr's postulates

• By examining the H atom, Bohr realized that the electron could exist in certain specific states of definite energy (**energy levels**) without radiating energy, if a certain condition was met by the orbit radius.

• The electron energy is thus **discrete** and not continuous.

• An electron can only lose energy when it makes a transition from one state to another of lower energy. The emitted energy is then the difference of energy between the initial and final states.



• The evidence for this is the **absorption** and **emission spectra**.

Energy levels - evidence

• Thomas Melvill (1726-1753) was the first to study the light emitted by various gases. He used a flame as a heat source, and passed the light emitted through a prism.

• Melvill discovered that the pattern produced by light from heated gases is very different from the continuous rainbow pattern produced when sunlight passes through a prism.

• The new type of spectrum consisted of a series of bright lines separated by dark gaps.

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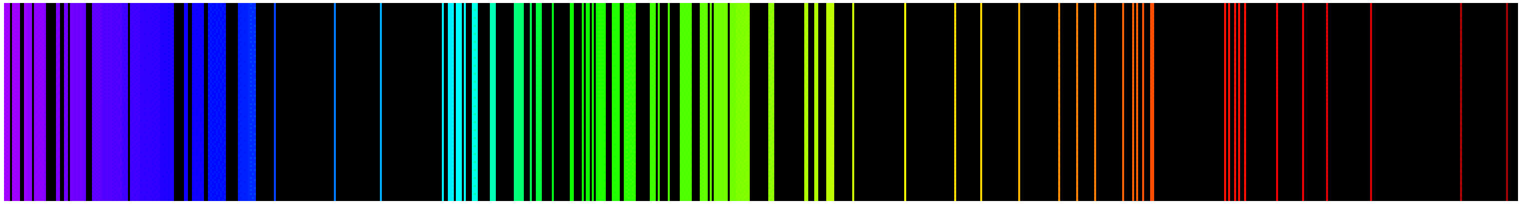


Emission spectrum of iron

Emission spectra

Individual atoms, free of the strong interactions that are present in a solid, emit only certain specific wavelengths that are unique to those atoms.

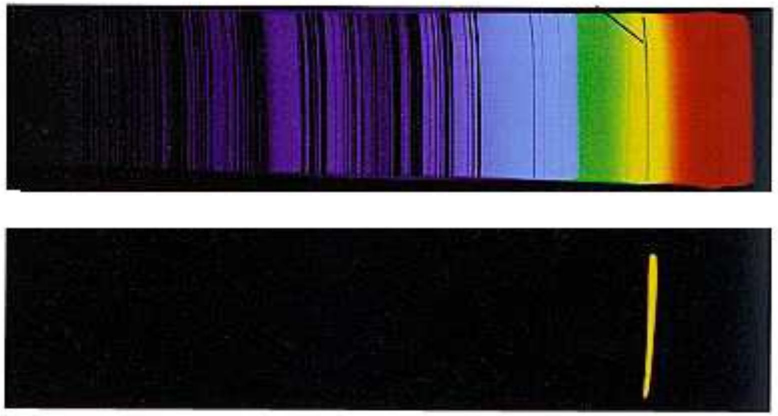
Energy levels - evidence

• This spectrum became known as a line spectrum.

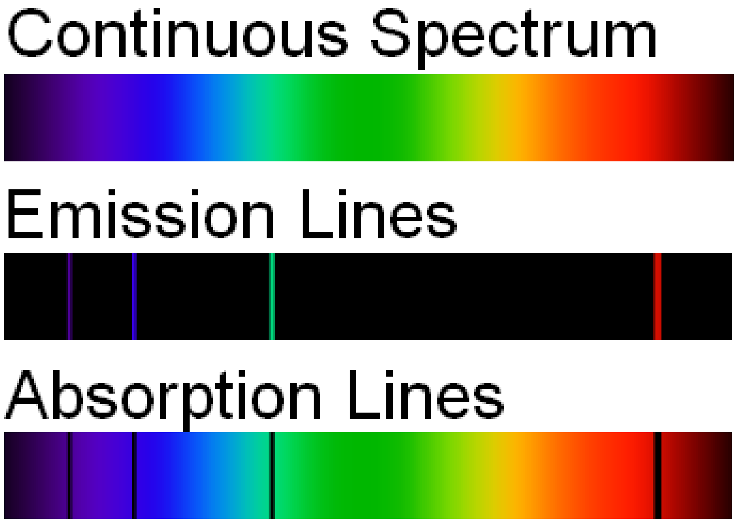
• Melvill also noted the line spectrum produced by a particular gas was always the same.

•In other words, the spectrum was characteristic of the type of gas, a kind of "fingerprint" of the element or compound.

• This was a very important finding as it opened the door to further studies, and ultimately led scientists to a greater understanding of the atom.

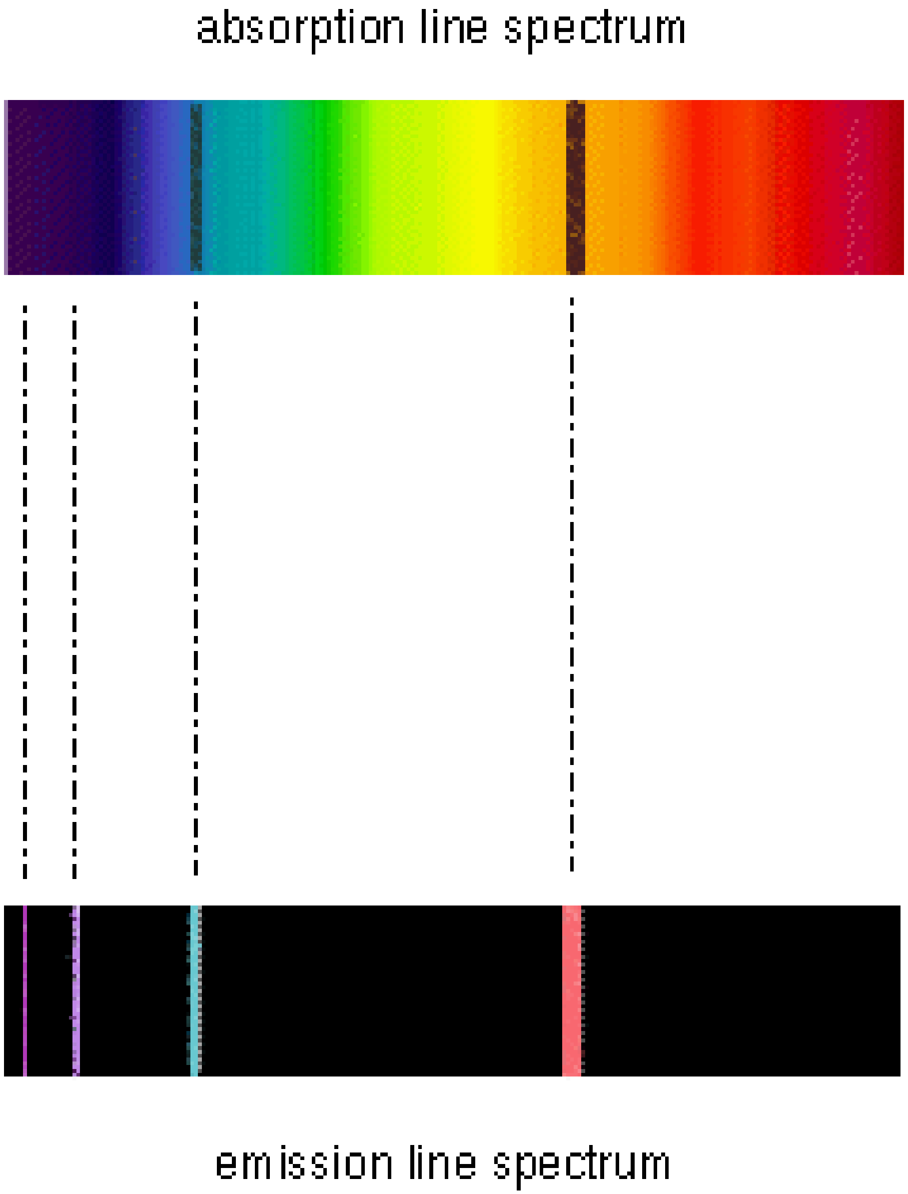


Emission and absorption spectra



- Spectra can be categorised as either emission or absorption spectra.
- An emission spectrum is, as the name suggests, a spectrum of light emitted by an element.
- It appears as a series of bright lines, with dark gaps between the lines where no light is emitted.
- An absorption spectrum is just the opposite, consisting of a bright, continuous spectrum covering the full range of visible colours, with dark lines where the element literally absorbs light.
- The dark lines on an absorption spectrum will fall in exactly the same position as the bright lines on an emission spectrum for a given element, such as neon or sodium.

Emission and absorption spectra for the same gas



Line spectra

What causes line spectra?

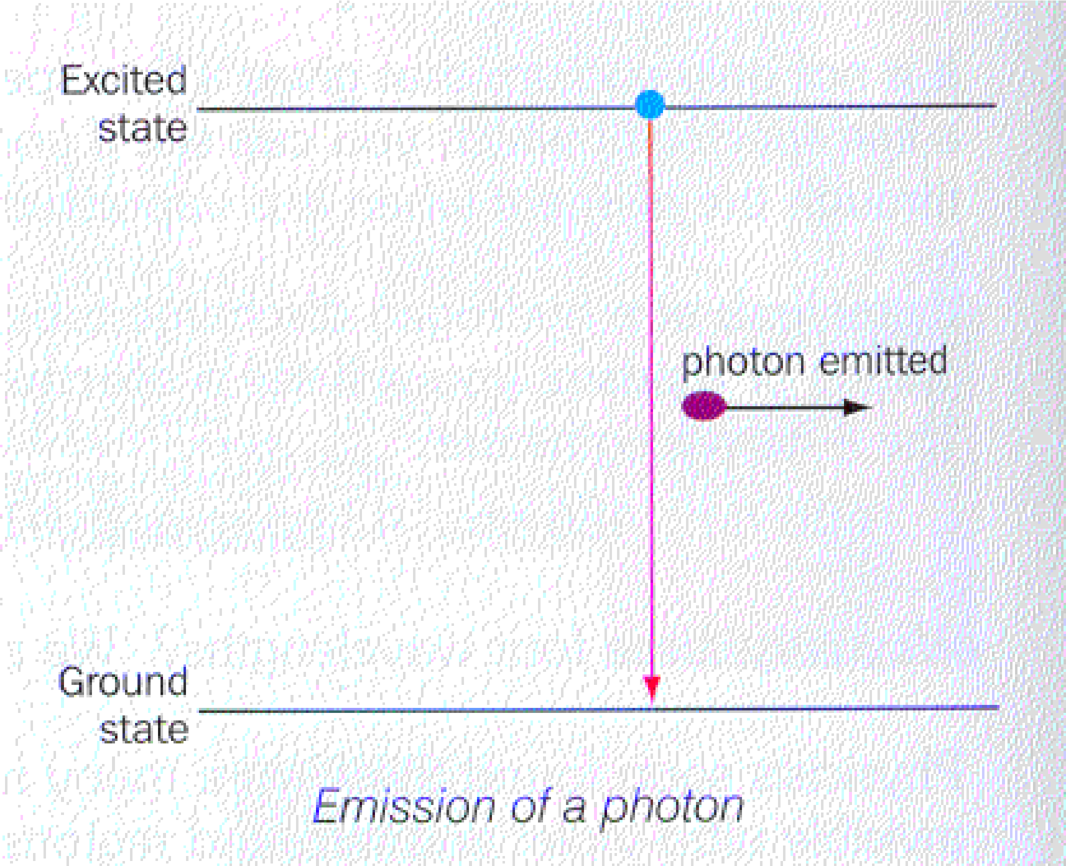
- You always get line spectra from atoms that have been excited in some way, either by heating or by an electrical discharge.
- In the atoms, the energy has been given to the electrons, which then release it as light.
- Line spectra are caused by changes in the energy of the electrons.
- Large, complicated atoms like neon give very complex line spectra, so physicists first investigated the line spectrum of the simplest possible atom, hydrogen, which has only one electron.

Line spectra

- Planck and Einstein's quantum theory of light gives us the key to understanding the regular patterns in line spectra.

- The photons in these line spectra have certain energy values only, so the electrons in those atoms can only have certain energy values.

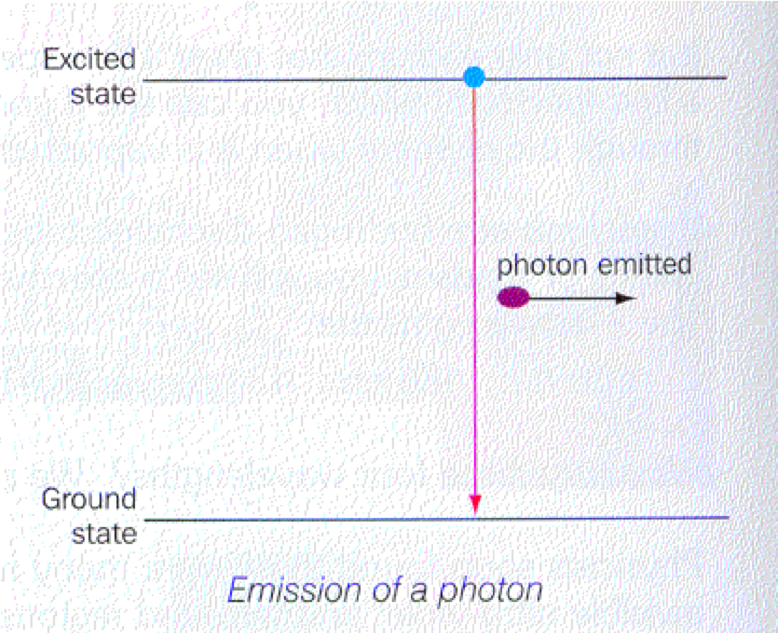
- This energy level diagram shows a very simple case. It is for an atom in which there are only two possible energy levels:



Line spectra

- The electron, shown by the blue dot, has the most potential energy when it is on the upper level, or **excited state**.

- When the electron is on the lower level, or **ground state**, it has the least potential energy.



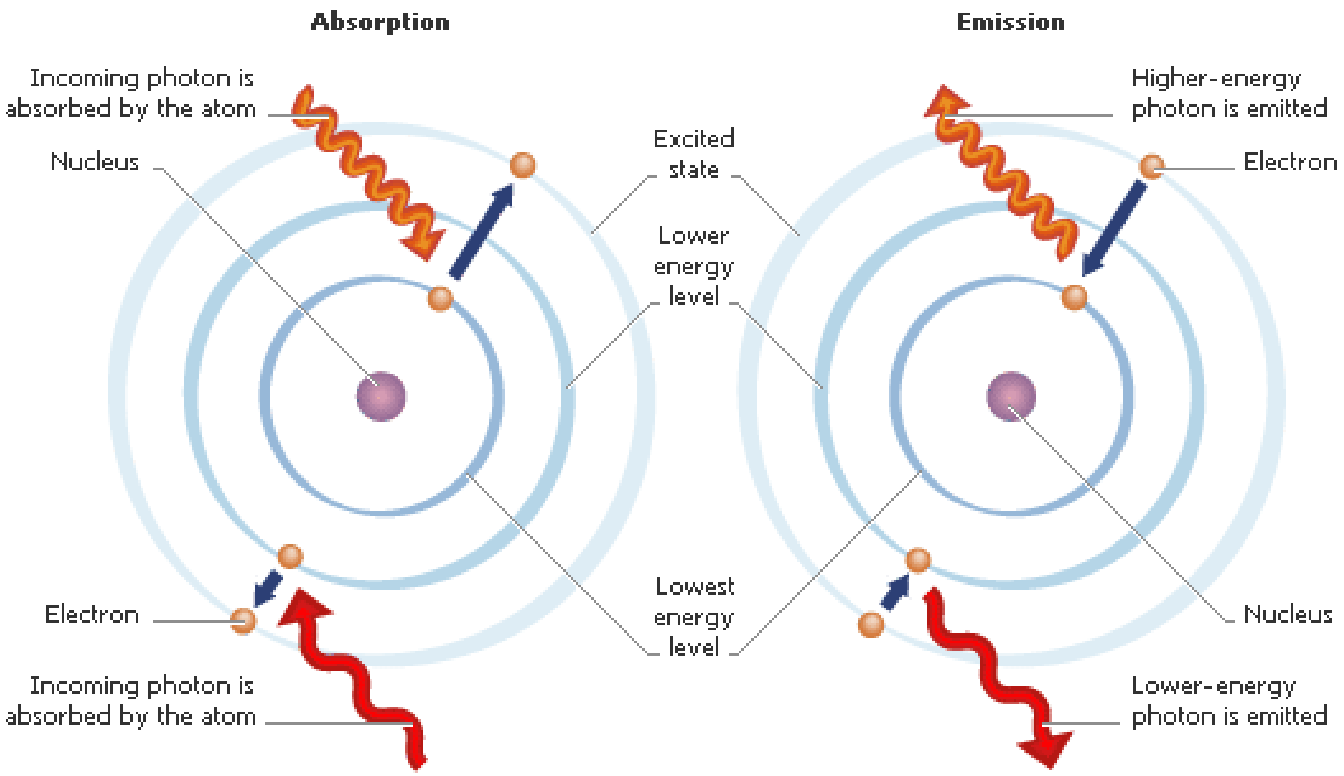
- The diagram shows an electron in an excited atom dropping from the excited state to the ground state.

- This energy jump, or **transition**, has to be done as one jump. It cannot be done in stages.

- This transition is the smallest amount of energy that this atom can lose, and is called a **quantum** (plural = quanta).

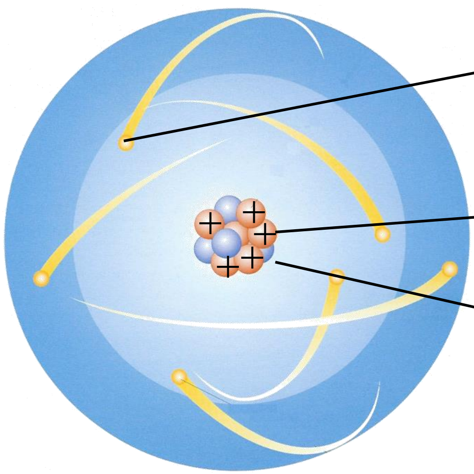
Line spectra

- The potential energy that the electron has lost is given out as a photon (particle of light).
- This energy jump corresponds to a specific **frequency** (or **wavelength**) giving a specific line in the line spectrum.
- This outlines the evidence for the existence of atomic energy levels.



Nuclear Structure

Atomic structure



Electrons (negative particles) - **e**

Protons (positive particles) - **p**

Neutrons (uncharged particles) - **n**

Particle	Relative Mass	Charge	Location
Proton	1	+1	Nucleus
Neutron	1	0	Nucleus
Electron	1/1800	-1	Electric cloud

Mass number and atomic number

A – Mass number

Z – Atomic number

$$^{A}_{Z}X \rightarrow \text{Element X}$$

Mass number = no. of protons + no. neutrons

= no. of nucleons

$A = p + n$

Atomic number = no. of protons

$Z = p$

Atoms have no charge. So, no. electrons = no. protons

$e = p$

Elements

- All materials are made from about 100 basic substances called **elements.**
- An atom is the smallest “piece” of an element you can have.
- Each element has a different number of protons in its atoms:
  - it has a different **atomic number** (sometimes called the **proton number).**
  - The atomic number also tells you the number of electrons in the atom.

All nuclides can be described using this format:

Nucleon number A

Chemical symbol of element

Proton number Z

Eg. $^{7}_{3}\mathbf{Li}$

number of nucleons (protons + neutrons)

number of protons

Eg. $^{9}_{4}Be$ Be = Beryllium

Isotopes

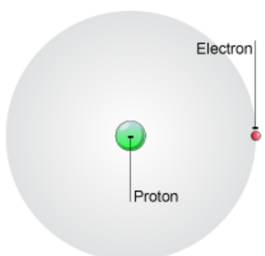
Isotopes are atoms that have the same number of protons but different number of neutrons.

Hydrogen	Deuterium	Tritium
$^{1}_{1}H$	$^{2}_{1}H$	$^{3}_{1}H$
1 proton	**1 proton**	**1 proton**
0 neutrons	**1 neutron**	**2 neutrons**

Isotopes

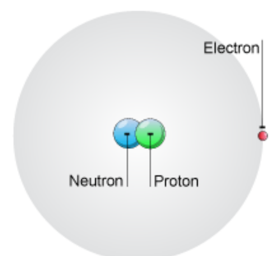
Hydrogen

$${}^{1}_{1}\mathrm{H}$$



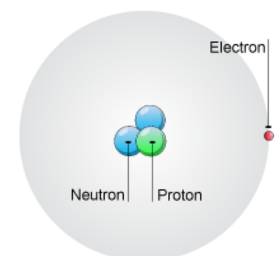
Deuterium

$${}^{2}_{1}\mathrm{H}$$



Tritium

$${}^{3}_{1}\mathrm{H}$$



- Since the isotopes of an element have the same number, of electrons, they must have the same chemical properties.
- The atoms have different masses, however, and so their physical properties are different.

Evidence for neutrons

- The existence of isotopes is evidence for the existence of neutrons because there is no other way to explain the mass difference of two isotopes of the same element.

- By definition, two isotopes of the same element must have the same number of protons, which means the mass attributed to those protons must be the same.

- Therefore, there must be some other particle that accounts for the difference in mass, and that particle is the neutron.

Interactions in the nucleus

- Electrons are held in orbit by the force of attraction between opposite charges.

- Protons and neutrons (nucleons) are bound tightly together in the nucleus by a different kind of force, called the **strong, short-range nuclear force.**

- It is this force that prevents the protons from repelling each other and breaking the nucleus apart.

- There are also **Coulomb interaction** between protons due to the fact that they are charged particles.