Discovery of electron

- J.J. Thomson measured e/m of a “tiny negatively charged particle”, 1897
- R. Millikan measured e, 1910 and Planck’s constant in 1912-1915
- The first discovered particle! - quite elementary until today

Structure? The major experimental facts:
1. the periodic table of elements - Dmitri Mendeleev, 1869
2. the stable matter is electrically neutral

Major hypothesis (J.J. Thomson):
- Positively charged spheres orbited by electrons - from the side raisin pudding positively charged stuff (pudding) stuffed with electrons - raisin
Discovery of radioactivity

- **Becquerel, 1896** – discovery of natural radioactivity – some matter emits invisible radiation (Uranium salt)
- The emitted rays are not X-rays discovered earlier by Wilhelm Roentgen
- Marie and Pierre Curie explored newly found radiation, separated polonium and radium
- Discovery of radon by F.E. Dorn, 1900
- F. Soddy, A. Fleck, Antonius Van den Broek - Becquerel’s found radioactivity is due to $\alpha$ particle (charge +2) - helium nucleus
- Final contribution Moseley - X-ray characteristic spectra: charge of the nucleus $=$ its atomic number
Rutherford experiment

- Rutherford scattered $\alpha$-particles on gold foil – first scattering experiment 1906-1909
- J.J. Thomson model “plum pudding” – there is certain density of matter – given enough energy a particles should get through being scattered by certain angles and having lost certain energy

The results are quite unexpected:
- most of a particles go through hardly scattered at all, not losing energy
- Some, very few, $\alpha$-particles are scattered backwards
Consequences: structure - matter consists of extremely dense and small positively charged nuclei and electrons orbiting them. The distances between the nuclei are many times larger than their sizes (by a factor of about $10^5$).

Rutherford continued scattering experiments after WW I - 1919

He coined the word "proton", 1920.

Was looking for the structure of $\alpha$-particle

Predicted neutron, discovered by James Chadwick, 1932

Nuclear Physics: Physics of the nucleus itself - 1921 - strong interactions!
Masses of components

\[
m_{\text{proton}} = 1.673 \times 10^{-27} \text{ kg}
\]

\[
m_{\text{neutron}} = 1.675 \times 10^{-27} \text{ kg}
\]

\[
m_{\text{electron}} = 9.11 \times 10^{-31} \text{ kg}
\]

Because of the mass – energy equivalence, it is convenient to introduce different units for masses:

\[
E_0 = mc^2 \quad \Rightarrow \quad m = E_0 / c^2
\]

atomic mass unit = 1/12 of the most abundant C isotope

1u : \[ E_0 = 1.66 \times 10^{-27} \times 9 \times 10^{16} / 1.6 \times 10^{-19} = 932.6 \text{ MeV} \]

\[
m_{\text{proton}} = 938.3 \text{ MeV/c}^2 = 1.00730u
\]

\[
m_{\text{neutron}} = 939.6 \text{ MeV/c}^2 = 1.00869u
\]

\[
m_{\text{electron}} = 0.511 \text{ MeV/c}^2 = 0.00055u
\]
Nuclear Definitions

Nucleus: “made of” protons & neutrons = nucleons

Mass number, \( A \): Number of nucleons in nucleus

Atomic Number, \( Z \): Number of protons in nucleus, amount of positive charge, position on periodic table

Neutron Number, \( N \): Number of neutrons in nucleus

\[ A = Z + N \]

Isotopes: Nuclei with same \( Z \) (same element), but different \( N \) & \( A \).

Isobars: Nuclei with same \( A \) (roughly same mass), but different \( Z \) (element) and \( N \).
Notation for nuclei and particles:

\[ \frac{A}{Z} X \]

Examples: Carbon:

Proton: \( ^1_1 p \)
Neutron: \( ^1_0 n \)
Electron: \( ^0_{-1} e \)

Two different isotopes of Carbon:

\[ ^{12}\text{C} \quad ^{14}\text{C} \]
Several remarks about nuclei

- We say that nuclei are “made of” nucleons - protons and neutrons.
- This is not quite so - the nucleons (although are the building blocks) are not the same as bare protons and neutrons: a bare neutron is not stable - it decays in about 887 seconds!
- There are more effects like magic numbers, stable and unstable isotopes that are not just a straight consequence of the protons and neutrons being together.
- The strong force is needed to keep the nucleus together and overcome electrostatic repulsion.
RADIOACTIVITY

= Radioactive Decay

Some isotopes are unstable: too many neutrons, too few neutrons, too heavy.

These nuclei will transform into more stable nucleus.

In the process the nucleus will emit particles:
Alpha (\(\alpha\)): Helium nucleus, \(\frac{4}{2}\text{He}\)
Beta (\(\beta\)): Electron, \(-1e\)
Gamma (\(\gamma\)): electromagnetic radiation, gamma photon
Penetration of radiation

Radiation loses energy (scatters) and is then absorbed.

In general, the larger the energy is, the smaller is the cross section.

The damage is done in interaction - at smaller energies.
Alpha Decay

Very heavy nuclei (Z>82) decay by emitting an alpha particle.

Example:

\[
\begin{align*}
\text{Pu}^{242}_{94} & \rightarrow \text{U}^{238}_{92} + \text{He}^{4}_{2} \\
\end{align*}
\]
Question:
Radium-226 decays via an alpha decay. What does it decay to?

\[ ^{226}_{88} \text{Ra} \rightarrow ? + ^4_2 \text{He} \]

1. Radon (Rn 222), Z = 86
2. Radon (Rn 230), Z = 86
3. Thorium (Th 222), Z = 90
4. Thorium (Th 230), Z = 90
Beta Decay

In Beta decay a neutron is spontaneously converted to a proton and an electron.

Example:

\[ ^{14}_6 \text{C} \rightarrow ^{14}_7 \text{N} + ^0_{-1} \text{e} \]

NOTE: A neutron is not a proton and an electron stuck together.
QUESTION:
Consider the following reaction. Which isotope are we starting with?

\[
? \rightarrow _{54}^{131}\text{Xe} + _{-1}^0\text{e}
\]

1. Cesium (Cs), Z=55, A=130
2. Cesium (Cs), Z=55, A=131
3. Cesium (Cs), Z=55, A=132
4. Iodine (I), Z=53, A=132
5. Iodine (I), Z=53, A=131
6. Iodine (I), Z=53, A=132

\[
_{53}^{131}\text{I} \rightarrow _{54}^{131}\text{Xe} + _{-1}^0\text{e} + \bar{\nu}
\]
What’s that?

- In order to ensure energy conservation, another particle has been predicted by W. Pauli in 1930 (before the discovery of neutron). It has been discovered only in 1955 by F. Reines and C. Cowan.
- This particle is a neutrino. It is almost massless, has no charge and moves with almost the speed of light, very weakly interacts with matter...
Gamma Decay

Nuclei can be excited, just like electrons in an atom. They will emit a gamma photon and revert back to the ground state.

\[ ^{87}_{38} \text{Sr}^* \rightarrow ? + \gamma \]

\[ ^{87}_{38} \text{Sr}^* \rightarrow ^{87}_{38} \text{Sr} + \gamma \]
Beta\(^+\) decay

- Positron - the anti-particle of an electron - same mass and spin, but the charge is the same, but opposite sign.
- Positron was predicted by P.A.M. Dirac in 1930 and discovered by C. Anderson in 1932.
- Many elements undergo a so-called \(\beta^+\) decay emitting a positron (e\(^+\)).

\[
\begin{align*}
\text{Na}^{22}_{11} & \rightarrow \text{Ne}^{22}_{10} + e^0_{1} + \nu
\end{align*}
\]
Radioactivity

There are many modes of radioactive decay, and a particular isotopic nucleus may decay by more than one mode. The most common decay modes are:

- $\alpha$ Emission, loss of a helium nucleus
- $\beta$ Emission, loss of an electron
- Electron capture, gain of an electron
- Positron emission, loss of an anti-electron

These modes of decay are associated with particular parts of the Segré chart. All types of radioactivity are also associated with simultaneous $\gamma$ ray emission due to a rearrangement of the nuclear energy levels. The Segré chart does not show meta-stable isotopes which only emit a $\gamma$ ray.

A plot of $N$ vs $Z$ to display isotopes is called a Segré Chart.
Radioactivity and Energy

Particles emitted during radioactive decay have kinetic energy ⇒ Heat

Responsible for keeping the earth’s core molten ⇒ continental drift, volcanism
Used in some thermoelectric generators for space missions.

But where does this energy come from?

- Binding energy is negative!
- Each spontaneous decay works in such a way that the binding energy of the products is larger than the BE of the initial nucleus - the total energy of the nuclei is reduced and an excess of energy is expelled as kinetic energy of products
Half-Life

Not all radioactive isotopes decay at the same rate.

Measured by half-life: Time in which half of original material has decayed.

Note: “decaying” isotopes don’t disappear, they just transform into a different isotope.
Half-life is a constant for a given isotope.

Example: Half-life = 1 day

1 g radioactive isotope initially

How much is left after one day?

Answer: $\frac{1}{2}$ gram

QUESTION: How much is left after 1 additional day?

1. Nothing, since the other $\frac{1}{2}$ g has now decayed as well.
2. $\frac{1}{4}$ gram
3. $\frac{1}{2}$ gram
Answer: $\frac{1}{4}$ gram.

The half-life is always the time it takes for $\frac{1}{2}$ of the original amount to decay, whatever the initial amount maybe.

How is this possible?

Quantum mechanics: We can not predict how long a single nucleus will be stable. We can only predict the probability that it will decay in a certain time.

Half-life: Time interval during which nucleus has 50% chance to decay.
Half-lifes vary over a **HUGE** range:

<table>
<thead>
<tr>
<th>Element</th>
<th>Isotope</th>
<th>Half-Life</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrogen</td>
<td>$^1\text{H}$</td>
<td>...</td>
</tr>
<tr>
<td></td>
<td>$^2\text{H}$ deuterium</td>
<td>...</td>
</tr>
<tr>
<td></td>
<td>$^3\text{H}$ tritium</td>
<td>12.3 yr</td>
</tr>
<tr>
<td>Helium</td>
<td>$^3\text{He}$</td>
<td>...</td>
</tr>
<tr>
<td></td>
<td>$^4\text{He}$</td>
<td>...</td>
</tr>
<tr>
<td></td>
<td>$^5\text{He}$</td>
<td>$2 \times 10^{-21}$ s</td>
</tr>
<tr>
<td></td>
<td>$^6\text{He}$</td>
<td>0.805 s</td>
</tr>
<tr>
<td></td>
<td>$^8\text{He}$</td>
<td>0.119 s</td>
</tr>
<tr>
<td>Carbon</td>
<td>$^{12}\text{C}$</td>
<td>...</td>
</tr>
<tr>
<td></td>
<td>$^{13}\text{C}$</td>
<td>...</td>
</tr>
<tr>
<td></td>
<td>$^{14}\text{C}$</td>
<td>5,730 yr</td>
</tr>
<tr>
<td></td>
<td>$^{15}\text{C}$</td>
<td>24 s</td>
</tr>
<tr>
<td>Silver</td>
<td>$^{107}\text{Ag}^\text{a}$</td>
<td>44.2 s</td>
</tr>
<tr>
<td></td>
<td>$^{108}\text{Ag}$</td>
<td>2.42 min</td>
</tr>
<tr>
<td></td>
<td>$^{109}\text{Ag}^\text{b}$</td>
<td>39.8 s</td>
</tr>
<tr>
<td></td>
<td>$^{110}\text{Ag}$</td>
<td>24.6 s</td>
</tr>
<tr>
<td>Uranium</td>
<td>$^{232}\text{U}$</td>
<td>70 yr</td>
</tr>
<tr>
<td></td>
<td>$^{233}\text{U}$</td>
<td>159,000 yr</td>
</tr>
<tr>
<td></td>
<td>$^{234}\text{U}$</td>
<td>245,000 yr</td>
</tr>
<tr>
<td></td>
<td>$^{235}\text{U}$</td>
<td>704,000,000 yr</td>
</tr>
<tr>
<td></td>
<td>$^{236}\text{U}$</td>
<td>23,400,000 yr</td>
</tr>
<tr>
<td></td>
<td>$^{237}\text{U}$</td>
<td>6.75 d</td>
</tr>
<tr>
<td></td>
<td>$^{238}\text{U}$</td>
<td>4,470,000,000 yr</td>
</tr>
<tr>
<td></td>
<td>$^{239}\text{U}$</td>
<td>23.5 min</td>
</tr>
</tbody>
</table>

\( ^\text{a} \)Not every isotope of the element is given.

\( ^\text{b} \)Asterisks (*) indicate that the nuclei are in an excited state.
Radioactive Dating

Since half-lives are fixed they can be used to date things as long as we know the initial ratio of isotopes.

Example: Carbon dating

C-14 is produced in the upper atmosphere by bombardment of nitrogen by cosmic rays:

\[
\begin{array}{c}
\begin{array}{c}
\text{1} \\
\text{0} \\
\text{n}
\end{array}
\end{array}
\begin{array}{c}
\begin{array}{c}
\text{14} \\
\text{7} \\
\text{N}
\end{array}
\end{array}
\rightarrow
\begin{array}{c}
\begin{array}{c}
\text{14} \\
\text{6} \\
\text{C}
\end{array}
\end{array}
\begin{array}{c}
\begin{array}{c}
\text{1} \\
\text{p}
\end{array}
\end{array}
\end{array}
\]

C-14 decays with a half-life of 5,730 years back into nitrogen:

\[
\begin{array}{c}
\begin{array}{c}
\text{14} \\
\text{6} \\
\text{C}
\end{array}
\end{array}
\rightarrow
\begin{array}{c}
\begin{array}{c}
\text{14} \\
\text{7} \\
\text{N}
\end{array}
\end{array}
\begin{array}{c}
\begin{array}{c}
\text{0} \\
\text{e}
\end{array}
\end{array}
\end{array}
\]

Carbon Dating

As we breath, we continuously add carbon to our body that has a certain (very small) percentage of C-14.

Therefore the C-14/C-12 ratio is fixed as long as an organism is alive.

Once the organism dies, no new carbon is added and C-14 content goes down.

Half of the C-14 will be gone after 5,700 years, $\frac{3}{4}$ will be gone after 11,400 years etc.
Radioactive dating:

Carbon dating good for up to 40,000 years on organic materials (bones, wood).

Dating of rocks: Uranium-Lead, Potassium-Argon, Rubidium-Strontium, can date rocks back to billions of years.

Note: you do not need to know how much of the original isotope was there in the first place. Example: Rubidium-Strontium “isochrones”.

Whole-rock rubidium-strontium isochron for a set of samples of a Precambrian granite body exposed near Sudbury, Ontario.
Time scales

- Age of an average human: $8 \times 10^1$ years
- Age of human civilization: $5 \times 10^3$ years
- Age of upright walking human species: $2 \times 10^6$ years
- Age of first known life: $3.7 \times 10^9$ years
- Age of the Earth: $4.55 \times 10^9$ years
- Age of universe: $1.37 \times 10^{10}$ years
Artificial nuclear reactions

Radioactive isotopes occur naturally, but they can also be made artificially by bombarding nuclei with particles:

\[
\begin{align*}
^{238}_{92}\text{U} + ^{0}_{1}\text{n} & \rightarrow ^{239}_{92}\text{U} \\
^{239}_{92}\text{U} & \rightarrow ^{239}_{93}\text{Np} + ^{0}_{-1}\text{e} \\
^{239}_{93}\text{Np} & \rightarrow ^{239}_{94}\text{Pu} + ^{0}_{-1}\text{e}
\end{align*}
\]

Making nuclear fuel for reactors.

Irene and Frederic Joliot - Curie, 1934
Making use of binding energy

Mass and Energy are equivalent:

\[ E_0 = mc^2 \]

Binding energy - mass difference

\[
m_{\text{proton}} = 938.3 \text{ MeV/c}^2
\]
\[
m_{\text{neutron}} = 939.6 \text{ MeV/c}^2
\]
\[
m_{\text{electron}} = 0.511 \text{ MeV/c}^2
\]

\[
m_{\alpha} = 4.00 \text{ u} = 3,730.3 \text{ MeV/c}^2 < 3,755.8 \text{ MeV/c}^2
\]

Binding Energy = 25.5 MeV

\[
m_{\text{Fe}} = 55.85 \text{ u} = 52,085.7 \text{ MeV/c}^2 < 52,583.8 \text{ MeV/c}^2
\]

Binding Energy = 498.1 MeV
Fe has the largest binding energy per nucleon - the most desired position for a nucleus: lower fusing, higher decay.
Nuclear Fission:

Very heavy nuclei can be broken up into more stable (larger binding energy), smaller nuclei if bombarded by neutrons.
Each time a U-235 nucleus undergoes fission, it releases three more neutrons.

These neutrons can hit other U-235 nuclei and split them, releasing 9 more neutrons... 27 neutrons ... 81 neutrons ... 243 ... 729 ... 2,187 ... 6,561 ... 19,683 ...

Chain reaction!
Atomic bombs

- Explosive
- Plutonium
- Explosive
- Uranium-235
Fission can also be controlled...

Use low concentration of U-235. U-238 does not fission, but is much more abundant.

Fission of U-235 is more efficient if neutrons are slow. Use “moderator” (carbon, water) to slow down neutrons.
Some materials (cadmium, boron) absorb neutrons: Use as “control rods“:
Nuclear fusion

If we combine a proton and a neutron they form a deuteron.

They bind together (binding energy) and their combined mass is reduced.

Binding energy = “lost” mass $\times c^2$.

Fusion in stars

\[ H \rightarrow He \rightarrow C, O, Ne \rightarrow O, Ne, Na \rightarrow Si, S \rightarrow Fe \]

Fe is the heaviest element produced by fusion; more heavy elements are produced during supernovae explosions - similarly to the production of elements in neutron irradiation.
Nuclear Fusion

\[ \text{H} + \text{H} \rightarrow {^2}\text{H} + e^+ + \nu \]

\[ {^2}\text{H} + \text{H} \rightarrow {^3}\text{He} + \gamma \]

\[ {^3}\text{He} + {^3}\text{He} \rightarrow {^4}\text{He} + \text{H} + \text{H} \]

“proton-proton chain”
Tokamak

Temperature required is $1-3 \times 10^8$ K

ITER - 2005