Topic 10: atomic physics

Content:

- atomic physics – basic historical developments
- models of atom (Thomson, Rutheford, Bohr)
- electrons, protons, neutrons
- first steps to quantum mechanics (Planck, Einstein)
- duality of particles
- nuclear physics
- natural radioactivity, nuclear reactions

atomic physics – historical developments (1/20)

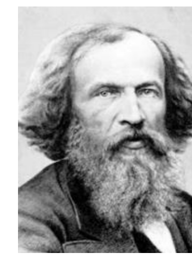
6th century BC - Democritus, along with Leucippus and Epicurus (ancient Greek pre-Socratic philosophers), proposed the earliest views on the idea of the atomic structure of the matter – first time the term atoma, (greek word for indivisible)

2th century BC - text of Vaisheshika school of philosophy (India),

18 centrury AD - british chemist and physicist John Dalton - introduced the term atom as the basic unit of a chemical element,

19 century: several important developments: atoms are stable objects, have inner structure, their dimensions have been estimated,

1869 - the periodic system of elements by Dmitri Mendeleev as a great step forward,



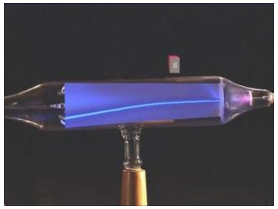
1891 – introduction of the term electron by G.J. Stoney, the Anglo-Irish physicist as the "fundamental unit quantity of electricity“, he predicted that they carry negative electrical charge and was able to estimate the value of this elementary charge *e* (but he believed these charges were permanently attached to atoms and could not be removed),

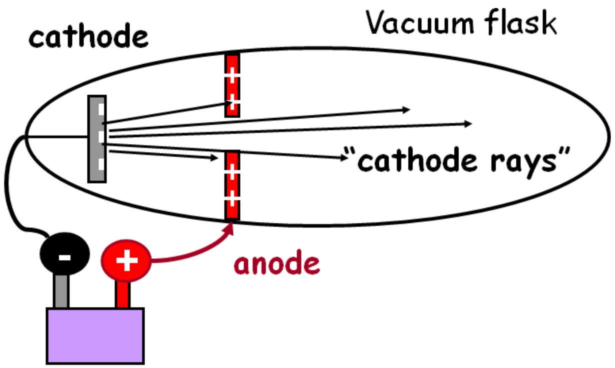
1897 - British physicist J.J. Thomson showed that the cathode rays were composed of a previously unknown negatively charged particle, which was named the electron (by G.F. Fitzgerald).

atomic physics – historical developments (2/20)

J. J. Thomson performed experiments indicating that cathode rays really were unique particles, rather than waves, atoms or molecules as was believed earlier

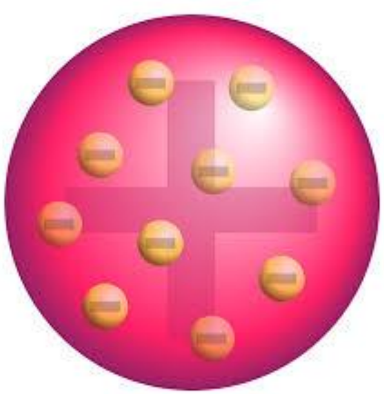
(he also estimated the mass of the electron, but called its as “corpuscle”).

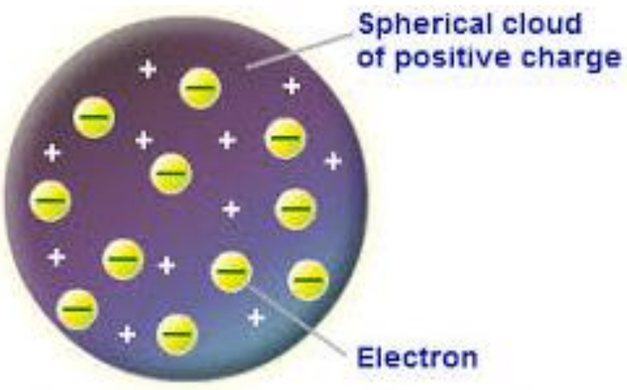




experiment:
https://www.youtube.com/watch?v=O9Goyscbazk

1904 – Thomson presented one of the first model of atom structure: an uniformly distributed substance - atom was made up of electrons scattered within an elastic sphere surrounded by a soup of positive charge to balance the electron's charge - like plums surrounded by pudding.



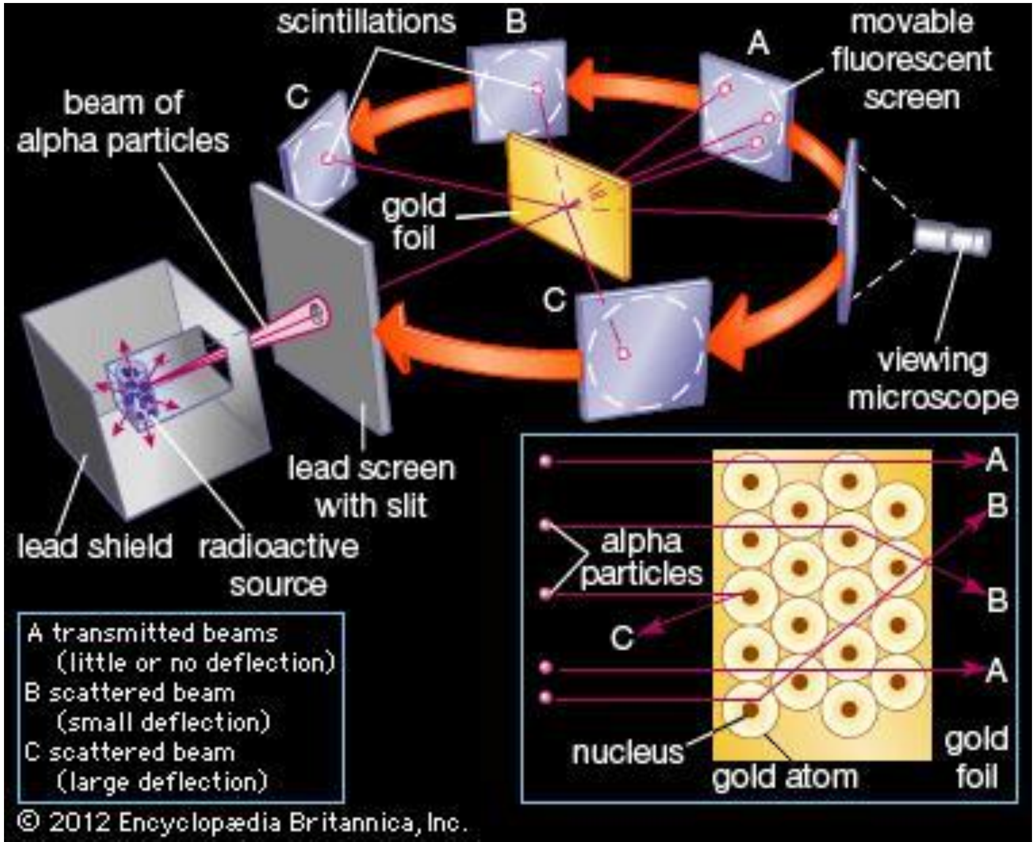


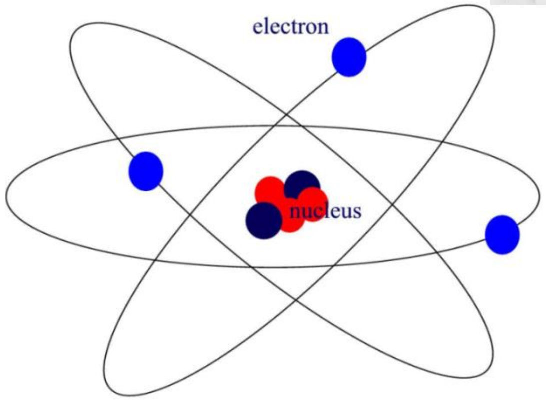
PLUM PUDDING MODEL

atomic physics – historical developments (3/20)

1911 Ernest Rutherford presented next important model of atom structure: called nuclear atom or planetary model of the atom.

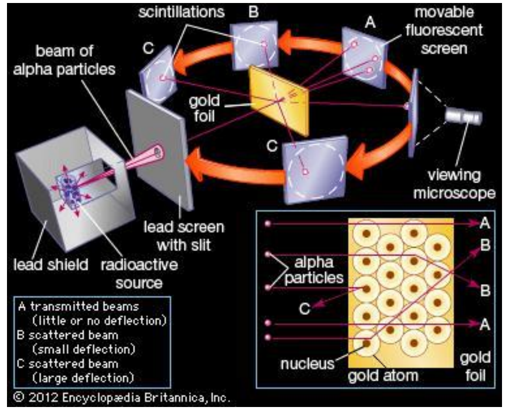
His gold-foil experiment: Rutherford disproved Thomson's model of the atom as a uniformly distributed substance. Because only very few of the alpha particles in his beam were scattered by large angles after striking the gold foil while most passed completely through, Rutherford knew that the gold atom's mass must be concentrated in a tiny heavy (dense) nucleus.

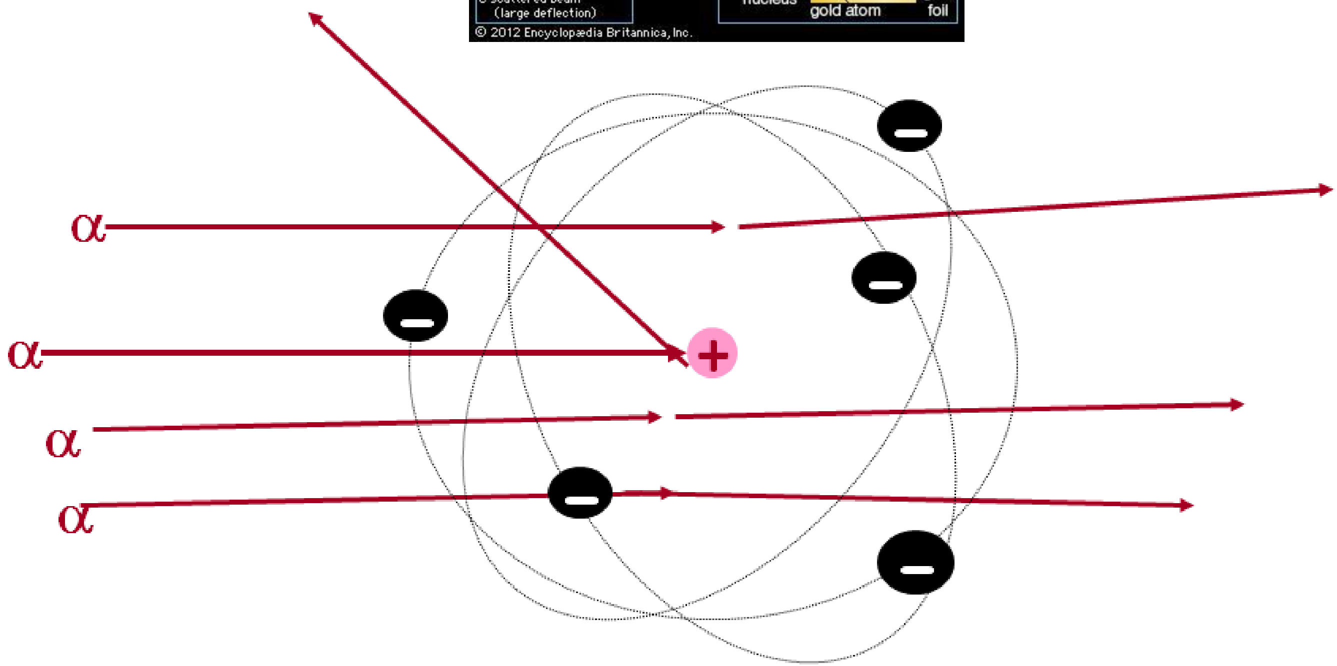


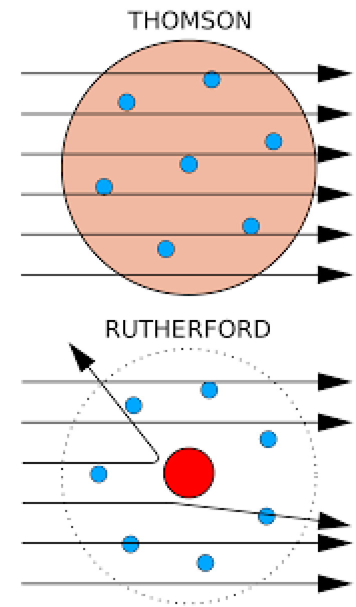


he suggested that the positive charge was all in a central nucleus.

1911 Ernest Rutherford - gold-foil experiment:





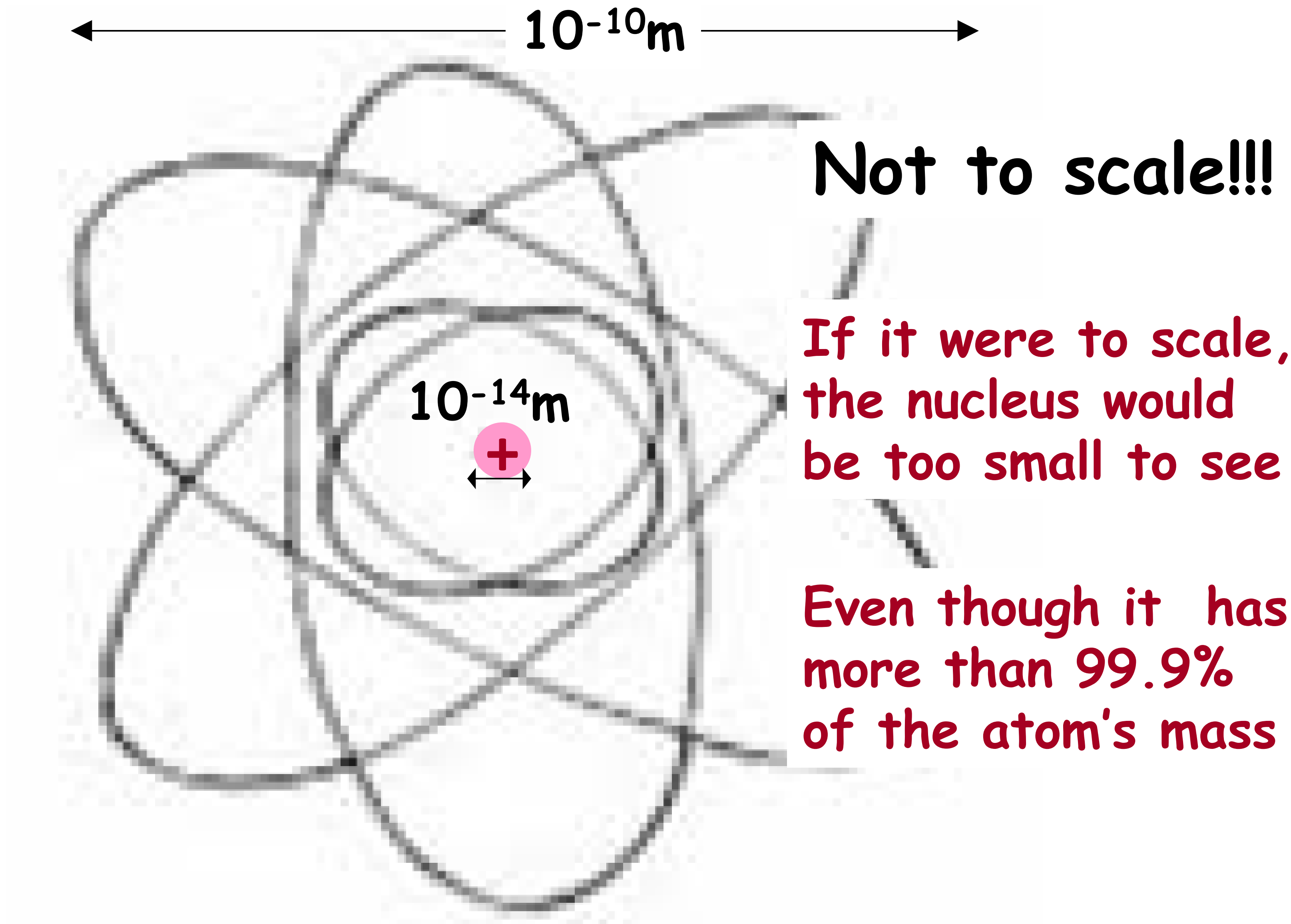


Rutherford saw ~1/10,000 α-rays scatter at wide angles from this he inferred a nuclear size of about 10^{-14} m.

modern reconstruction:
https://www.youtube.com/watch?v=XBqHkraf8iE

atomic physics – historical developments (4/20)

1911 Ernest Rutherford - nuclear model of the atom:



atomic physics – historical developments (5/20)

footbal stadium

golf ball

$\times 10^{-4}$

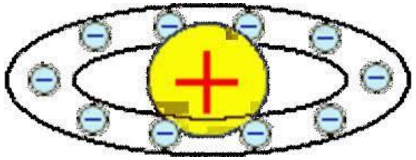
Atom

Nucleus
(99.97% of the atom mass)

atomic physics – historical developments (6/20)

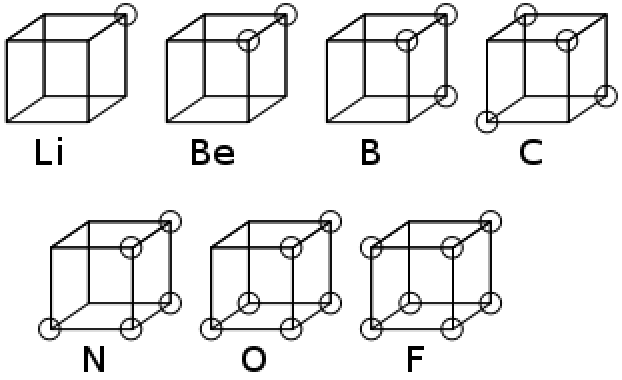
Comment: Other alternative (early) models of atom:

Saturnian model: in 1904, H. Nagaoka proposed an alternative planetary model of the atom in which a positively charged center is surrounded by a number of revolving electrons, in the manner of Saturn and its rings.



Nagaoka's model made two predictions:
- a very massive atomic center (in analogy to a very massive planet),
- electrons revolving around the nucleus, bound by electrostatic forces (in analogy to the rings revolving around Saturn, bound by gravitational forces).

Cubic model: Developed in 1902 by G.N. Lewis and published in 1916. Electrons were positioned at the eight corners of a cube, describing the atom.

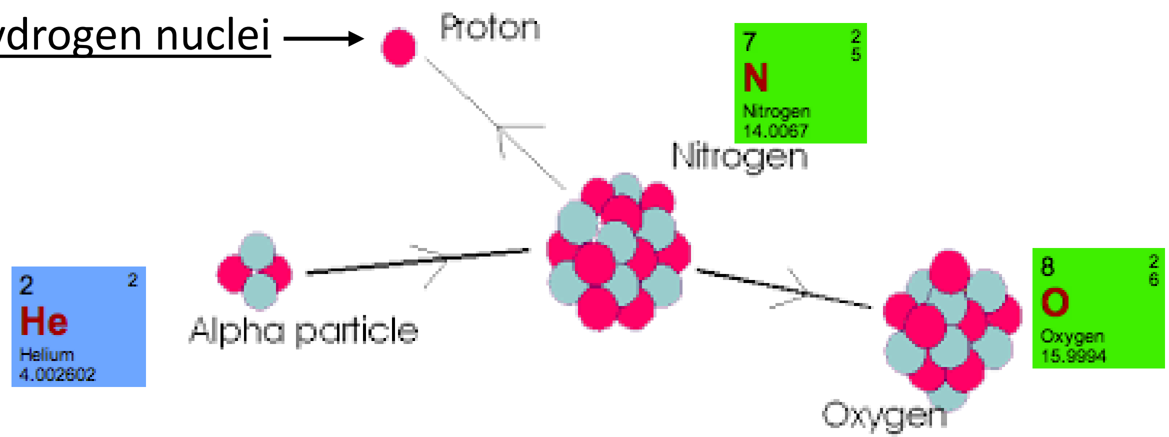


atomic physics – historical developments (7/20)

Nuclear model of the atom – protons:

In 1919 Rutherford and others discovered that they could change one element into another by striking it with energetic alpha particles (helium nuclei).
In the early 1920's Rutherford and other physicists made a number experiments, transmuting one atom into another. In every case, hydrogen nuclei were emitted in the process. It was apparent that the hydrogen nucleus played a fundamental role in atomic structure, and by comparing nuclear masses to charges, it was realized that the positive charge of any nucleus could be accounted for by an integer number of hydrogen nuclei. By the late 1920's physicists were regularly referring to hydrogen nuclei as 'protons'.

a hydrogen nuclei →

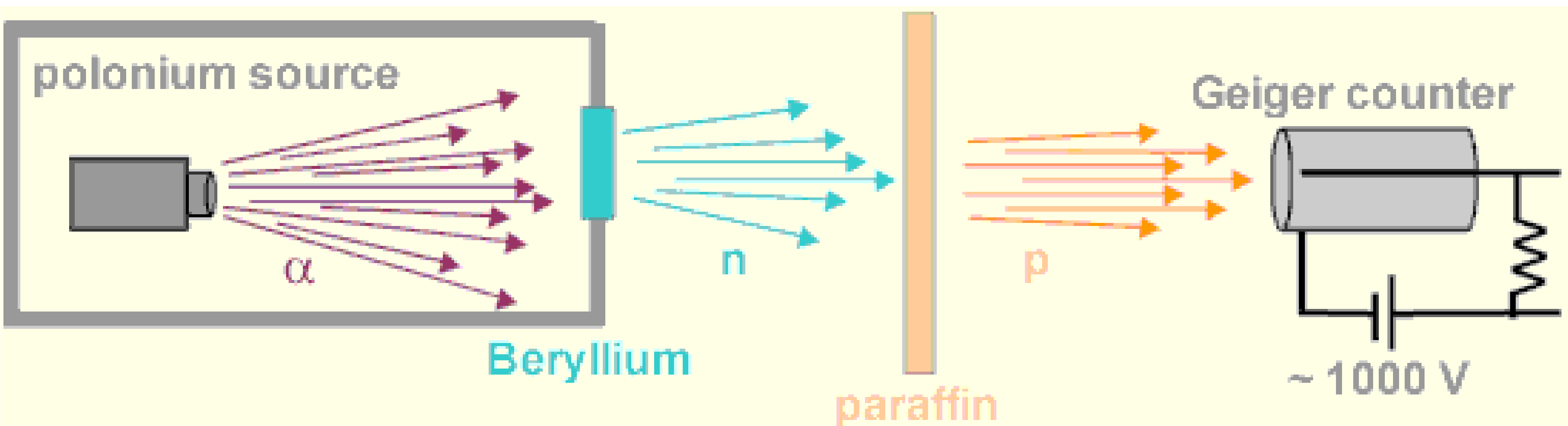


$$^{14}_{7}N + ^{4}_{2}He \rightarrow ^{17}_{8}O + ^{1}_{1}H$$

atomic physics – historical developments (8/20)

Nuclear model of the atom – neutrons:

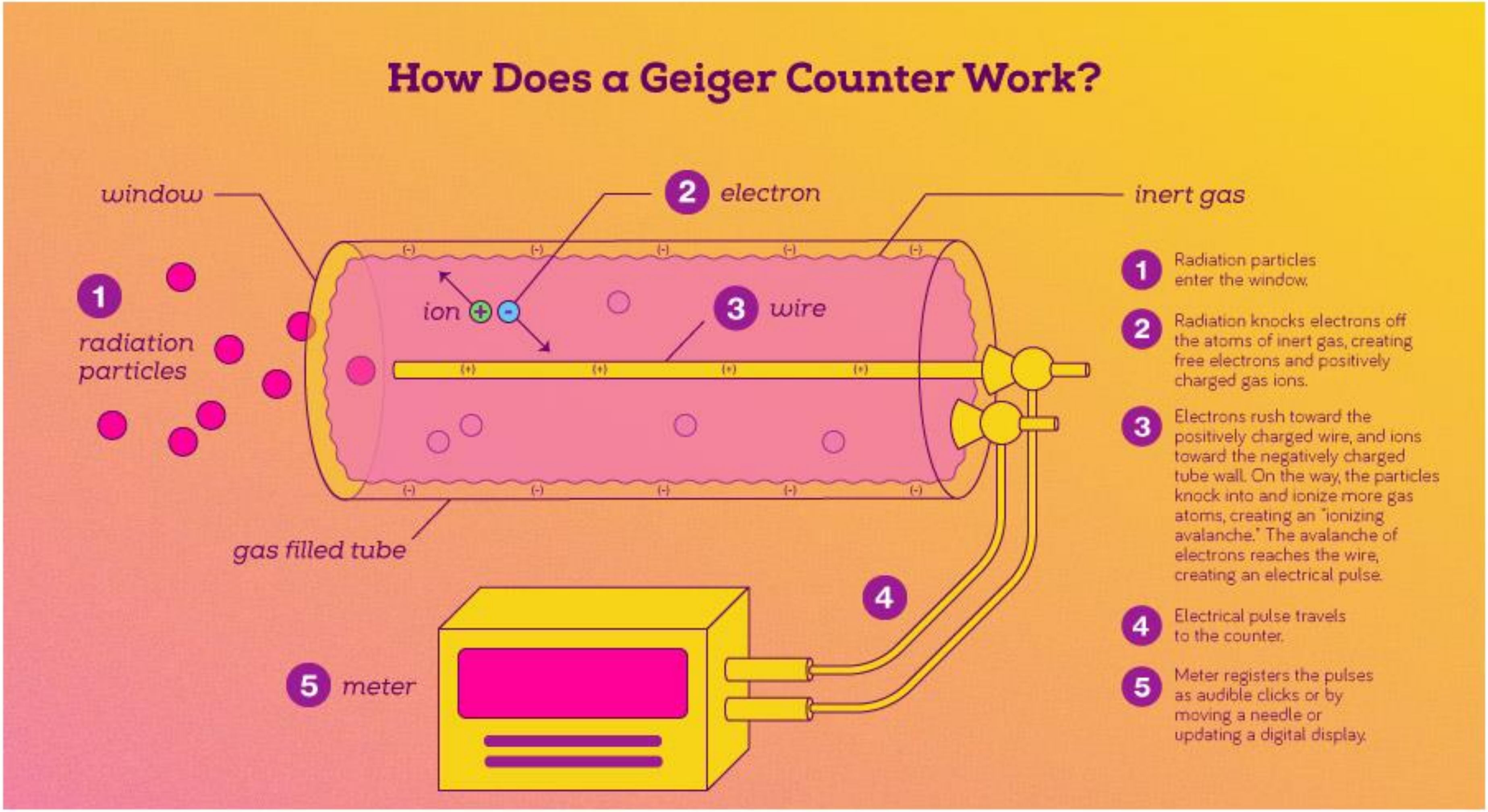
In 1920, Ernest Rutherford postulated that there were neutral, massive particles in the nucleus of atoms. This conclusion arose from the disparity between an element's atomic number (protons = electrons) and its atomic mass (usually in excess of the mass of the known protons present).
Experiments in 1930 showed that Beryllium, when bombarded by alpha particles, emitted a very energetic stream of radiation.
In 1932, J. Chadwick proposed that this particle was Rutherford's neutron.



$$ {}^{4}_{2}\alpha + {}^{9}_{4}Be \rightarrow {}^{12}_{6}C + {}^{1}_{0}n $$

atomic physics – historical developments (8b/20)

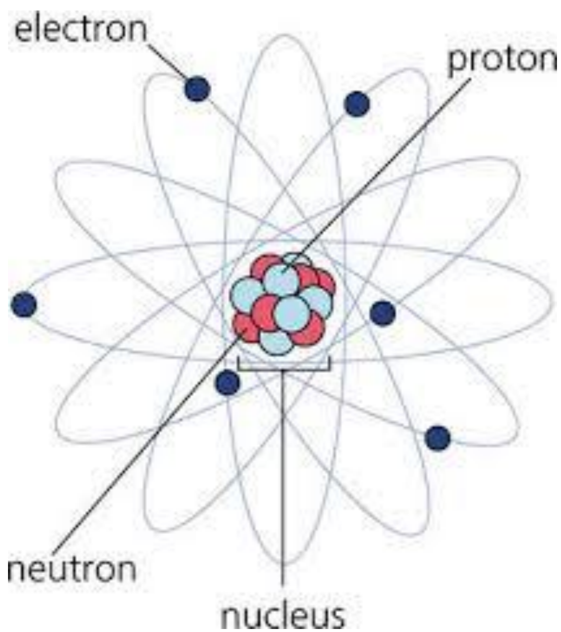
Comment: Geiger counter



atomic physics – historical developments (9/20)

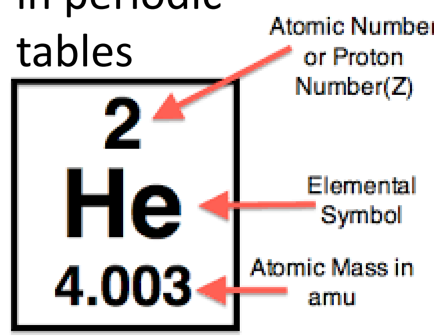
Nuclear model of the atom – nucleons:

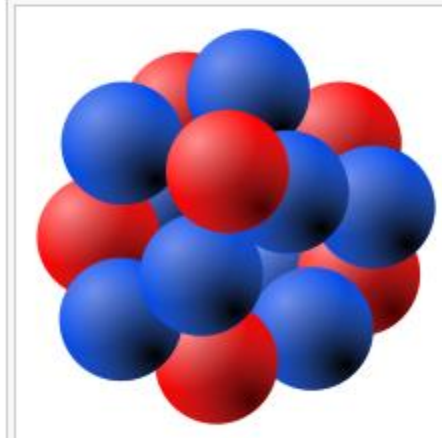
In 1932 – W. Heisenberg and D. Ivanenko proposed a hypothesis that nucleus of an atom is build by nucleons: protons (*p*) and neutrons (*n*).
The mass number of a given atomic isotope is identical to its number of nucleons.
Atomic number (also known as its proton number) is the number of protons found in the nucleus of an atom.



in chemical reactions

in periodic tables





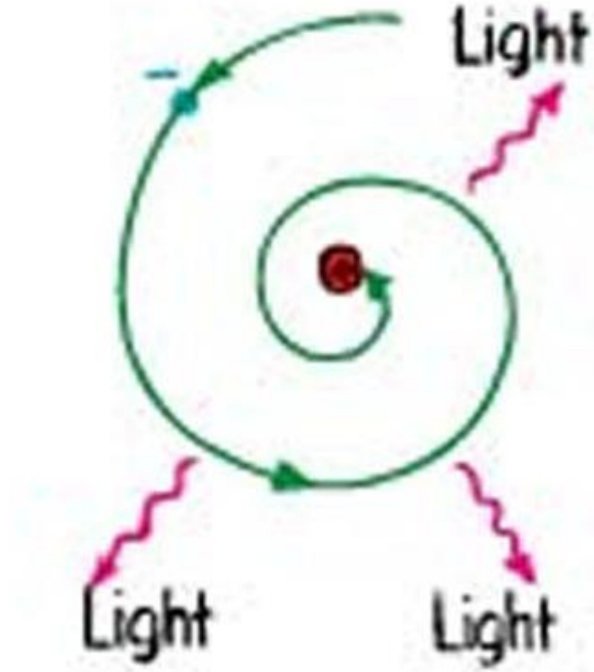
An atomic nucleus is a compact bundle of the two types of nucleons: Protons (red) and neutrons (blue). In this picture, the protons and neutrons look like little balls stuck together, but an actual nucleus, as understood by modern nuclear physics, does not look like this. An actual nucleus can only be accurately described using quantum mechanics. For example, in a real nucleus, each nucleon is in multiple locations at once, spread throughout the nucleus.

Until the 1960s, nucleons were thought to be elementary particles, each of which would not then have been made up of smaller parts. Now they are known to be composite particles, made of three quarks bound together by the so-called strong interaction.

atomic physics – historical developments (10/20)

Nuclear model of the atom – basic problem in classical EM theory:

After presenting the planetary model with tiny positive nucleus, classical physics predicted that orbiting electrons would spiral around the nucleus with increasing acceleration into the nucleus and radiate their energy away. This clearly did not happen.



according to Maxwell's theory, a Rutherford atom would only survive for only about 10^{-12} secs (!)

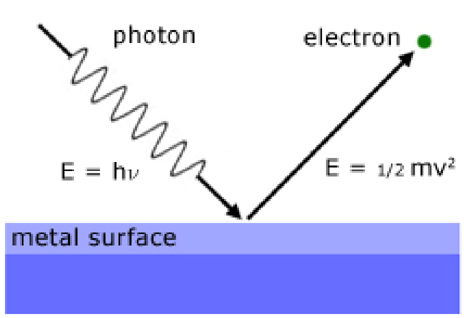
We will come to it back during the explanation of Bohr atom model, based on quantum mechanics interpretation.

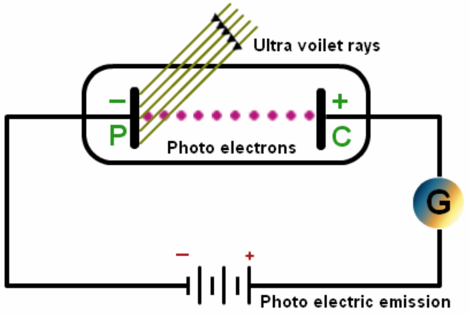
Physics had reached a turning point in 1900 with Planck's hypothesis of the quantum behavior of radiation, so a radical solution would be considered possible.
The successful atomic model, which came out of these attempts is the Bohr model – we will discuss it in more detail later on.

atomic physics – historical developments (11/20)

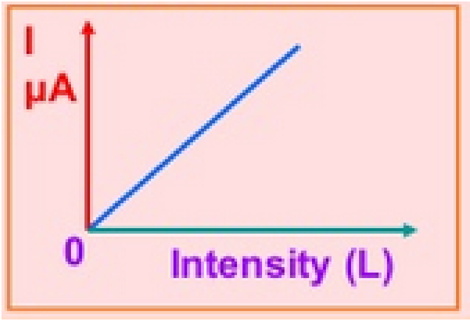
Photoelectric effect – first steps to quantum physics

When light shines on a metal surface, the surface emits electrons (even a current can be measured in a circuit).

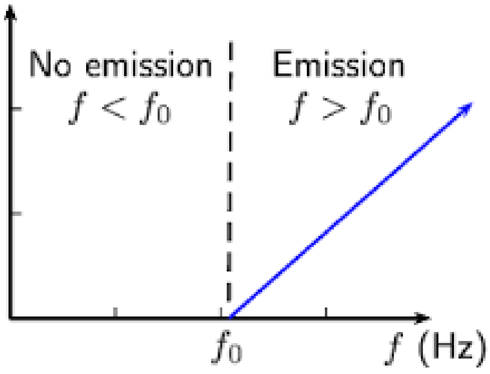




simple videos: https://www.youtube.com/watch?v=WO38qVDGgqw
https://www.youtube.com/watch?v=v-1zjdUTu0o



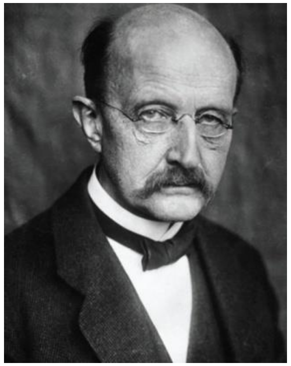
Ideal theoretical situation:
for constant frequency of light
(current is linearly dependent
from the light intensity)



but there exist a boundary (treshold) frequency, below which no electrones are emited – so they are emited in some kind of „packages“

atomic physics – historical developments (12/20)

Photoelectric effect – first steps to quantum physics



Important contribution from Max Planc – he introduced the concept of “quantum of action” (quantized nature of energy).

Plank‘s assumptions:

- The energy of an oscillator can have only certain discrete values E_n.

$$E_n = n\,h\,f$$

n is a positive integer called the quantum number, f is the frequency of oscillation, h is Planck’s constant, $h = 6.62607015\cdot10^{-34}$ [m^2kg/s]

  - this says the energy is quantized,
  - each discrete energy value corresponds to a different quantum state (each quantum state is represented by the quantum number, n).
- The oscillators emit or absorb energy when making a transition from one quantum state to another. Entire energy difference between the initial and final states in the transition is emitted or absorbed as a single quantum of radiation. An oscillator emits or absorbs energy only when it changes quantum states. The energy carried by the quantum of radiation is $E = h\,f$.

Einstein later successfully explained the photoelectric effect within the context of quantum physics – light consists of little particles or quanta, called photons.

comment to energy unit in atomic physics

$$E_n = n\,h\,f$$

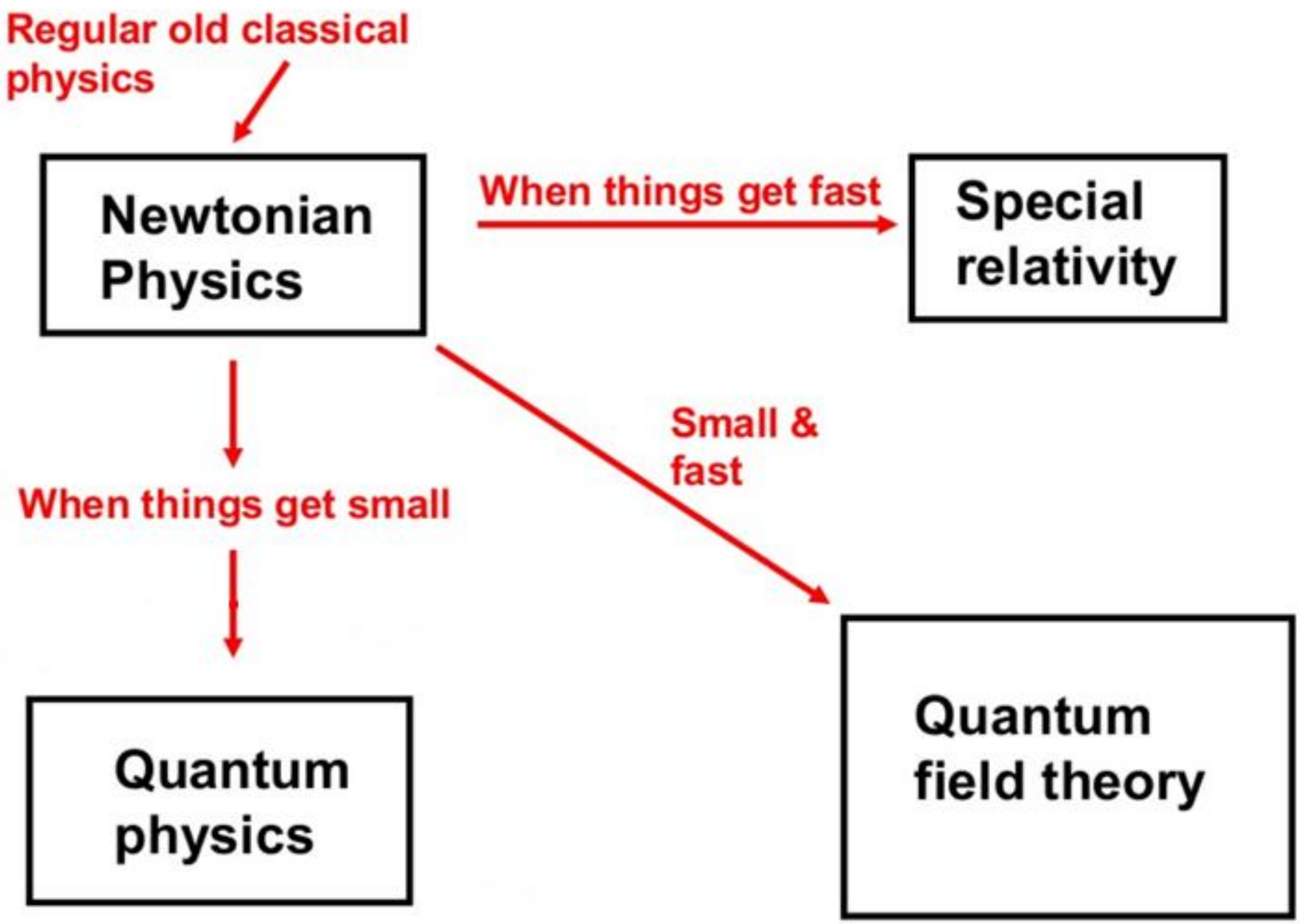
in the case of photoelectric effect: $E = W = h\,f_0$

- in atomic physics we do not use unit [J] for work/energy,
  instead of it, we use the unit electronvolt [eV],
  1 eV = energy gained by an electron when it is accelerated through a potential difference of 1 volt, 1 eV = $1.60217 \cdot 10^{-19}$ J),
- often, we use values MeV, because 1 eV is a very small number.

from lecture nr. 5 in winter term:
unit: volt [V] = [J/C] = [kg·m^2·s^{-2}] / [A·s] = [kg·m^2·s^{-3}·A^{-1}]
and later in this lecture – elementary charge e: 1 e = $1.60217XX \cdot 10^{-19}$ C
(electron has the charge -e, proton has the charge +e).

atomic physics – historical developments (13/20)

These discoveries moved physics from the classical to modern stage



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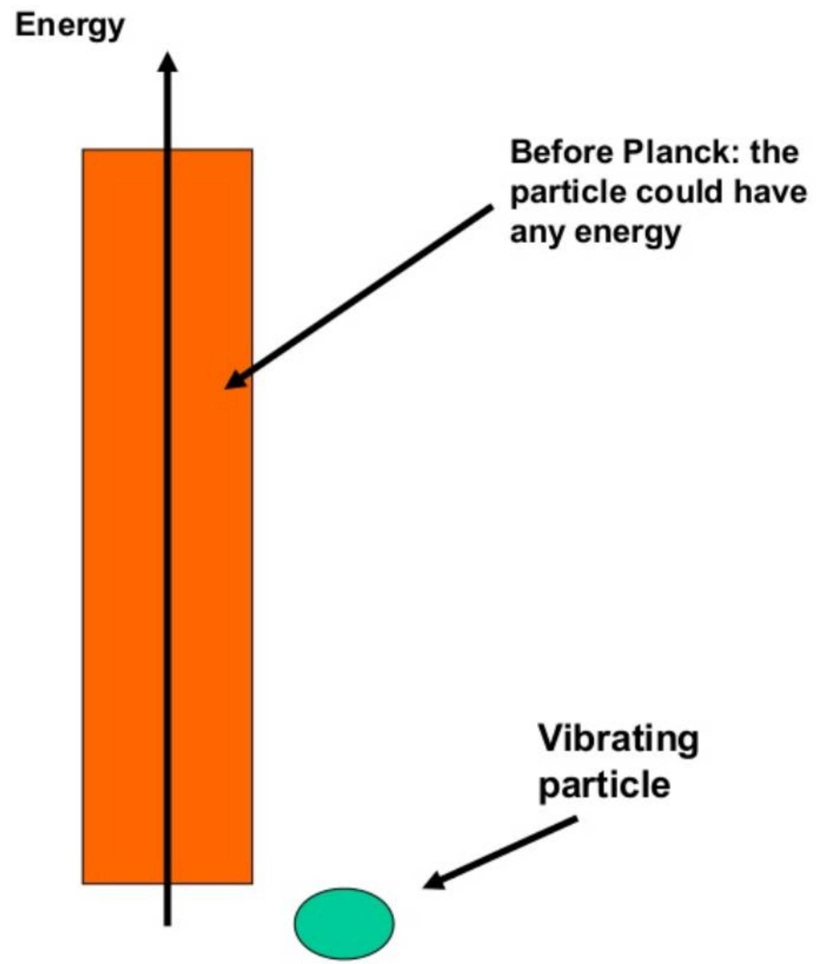
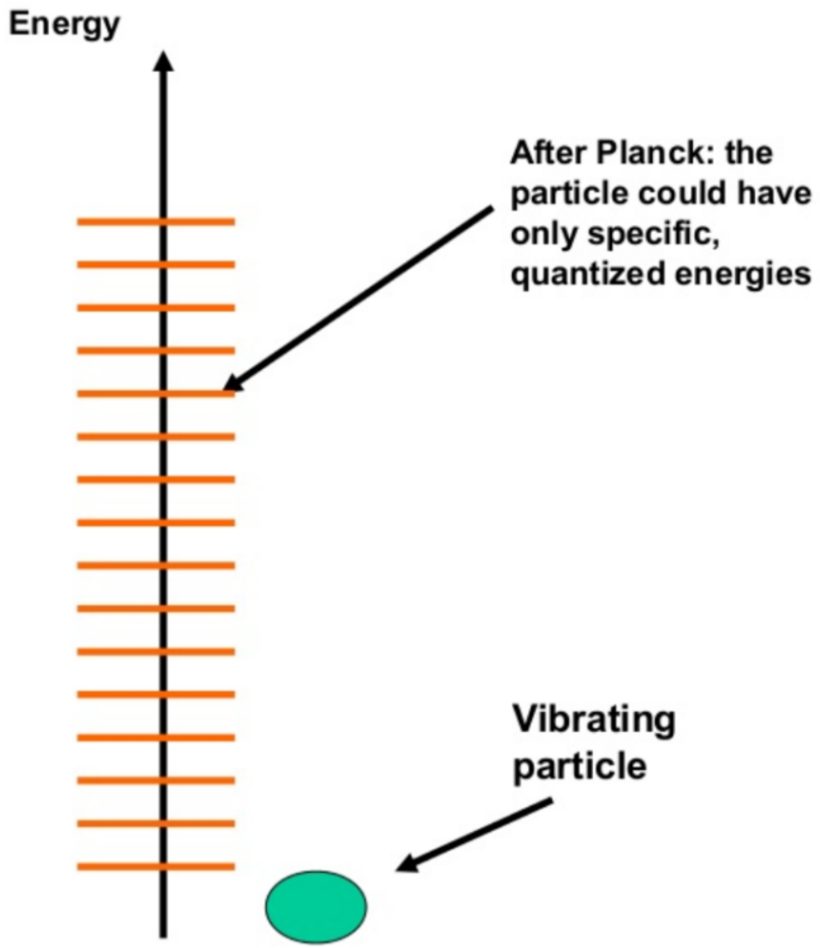


classical physics



quantum physics

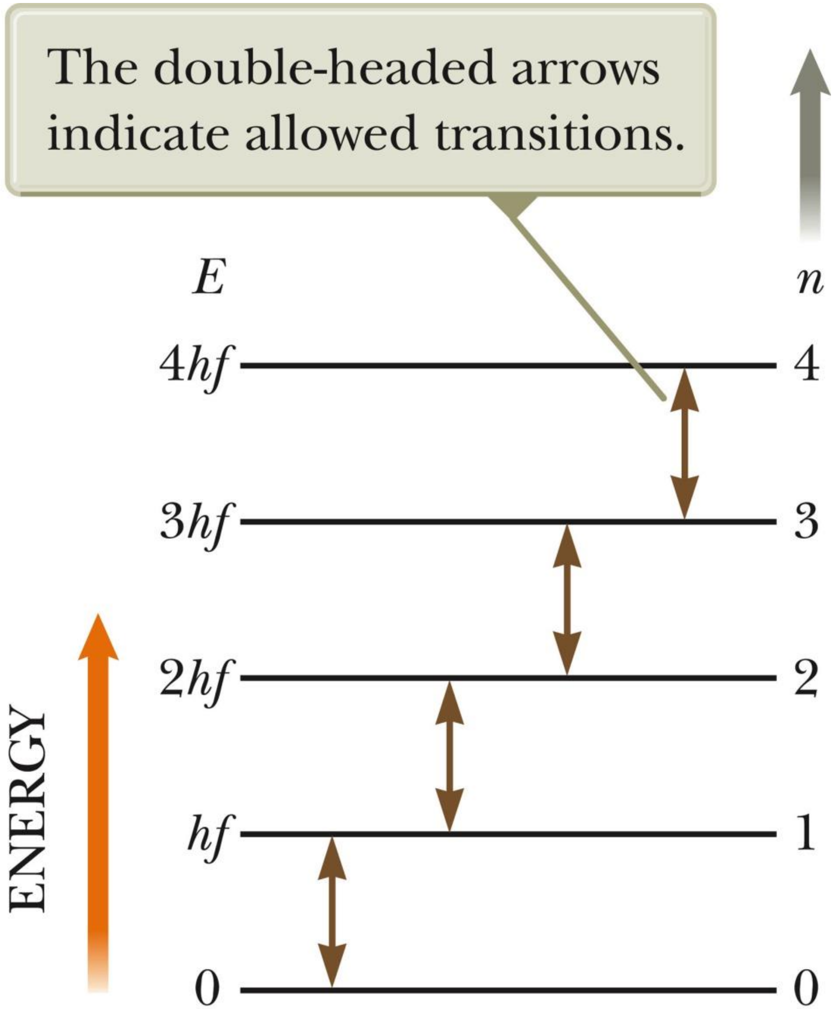
energy-level diagram

• an **energy-level diagram** shows the quantized energy levels and allowed transitions,

• energy is on the vertical axis,

• horizontal lines represent the allowed energy levels,

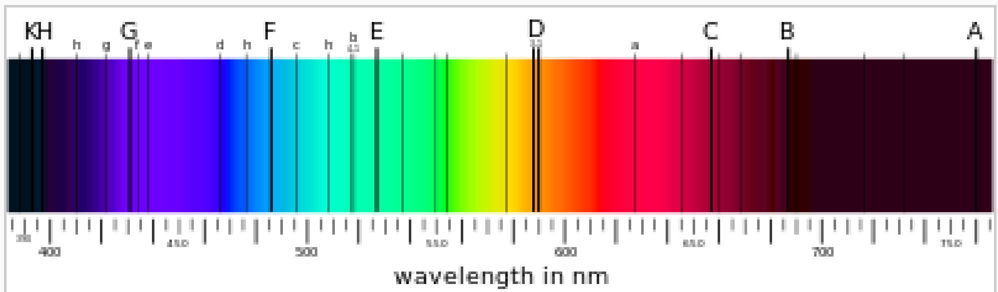
• the double-headed arrows indicate allowed transitions,



atomic physics – historical developments (14/20)

Important comment – absorption lines:

It is very important to mention also the discovery of spectral lines – the dark absorption lines by J. von Frauenhofer (1787-1826) in the optical spectrum of the Sun.
These have been first indication of quantum behaviour of photons - the study of these lines also contributed to the birth of quantum mechanics..



Solar spectrum with Fraunhofer lines as it appears visually.

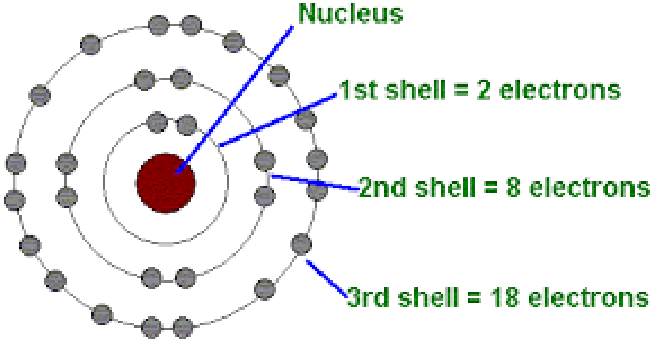
Fraunhofer lines are typical spectral absorption lines. Absorption lines are dark lines, narrow regions of decreased intensity, that are the result of photons being absorbed as light passes from the source to the detector. In the Sun, Fraunhofer lines are a result of gas in the photosphere, the outer region of the sun. The photosphere gas is colder than the inner regions and absorbs light emitted from those regions.

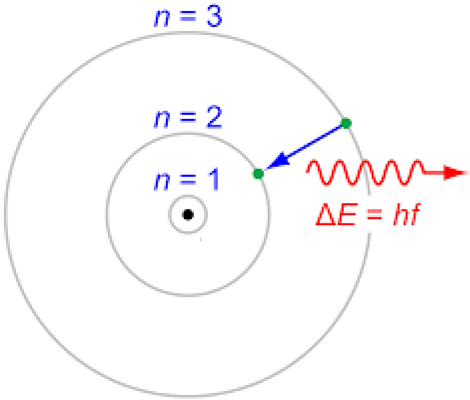
J. Rydberg published in 1888 a formula, which describes wavelengths of spectral lines of many chemical elements. It is a function of frequency and integer numbers (principal quantum numbers).

atomic physics – historical developments (15/20)

Important development – Bohr model of atom:

1913 Niels Bohr - new model of the atom – with quantum physics interpretation (called as Bohr model or Rutherford-Bohr model). It was a simple model of hydrogen atom, but it solved the problems coming from the classical EM theory.





Basic features of Bohr's model:

1. Electrons in atoms orbit the nucleus.
2. The electrons can only orbit stably, without radiating, in certain orbits (called the "stationary orbits") at a certain discrete set of distances from the nucleus. These orbits are associated with definite energies and are also called energy shells or energy levels. In these orbits, the electron's acceleration does not result in radiation and energy loss as required by classical electromagnetics. The Bohr model of an atom was based upon Planck's quantum theory of radiation.
3. Electrons can only gain and lose energy by jumping from one allowed orbit to another, absorbing or emitting electromagnetic radiation with a frequency determined by the energy difference of the levels according to the Planck relation:

$\Delta E = E_2 - E_1 = h f$, where h is Planck's constant and f the frequency of radiation.

atomic physics – historical developments (15/20)

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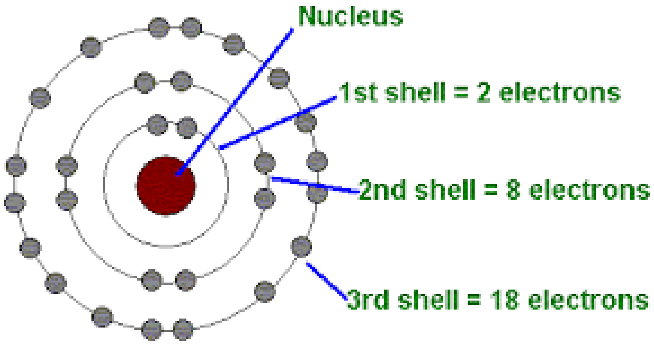
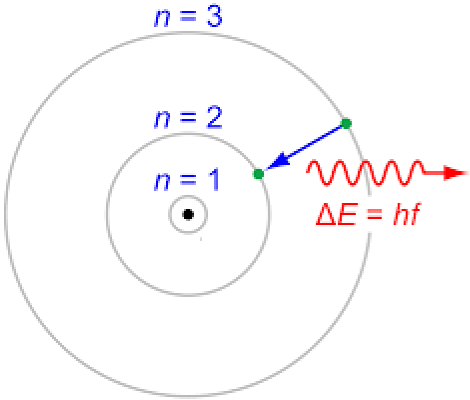


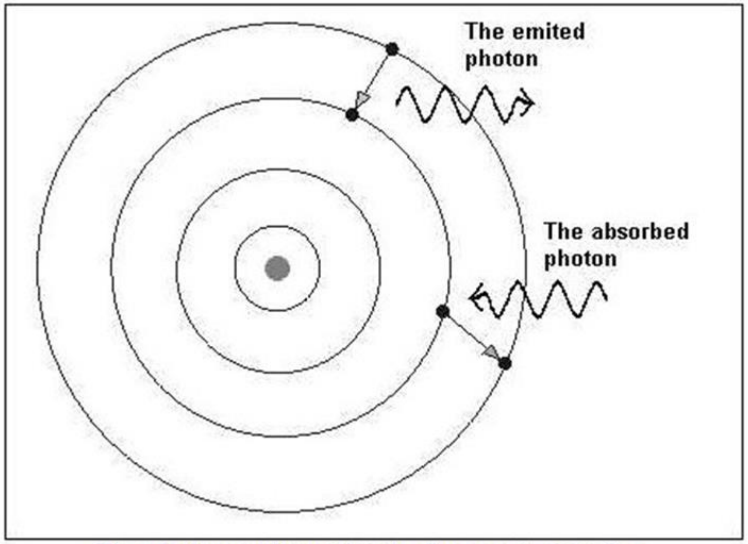


<u>Bohr's formula for the energy of atom</u>:

$$E = (-13.6)Z^2/(n^2) \text{ [eV]}$$

where: Z is atomic (proton) number, n is the order number of orbit.

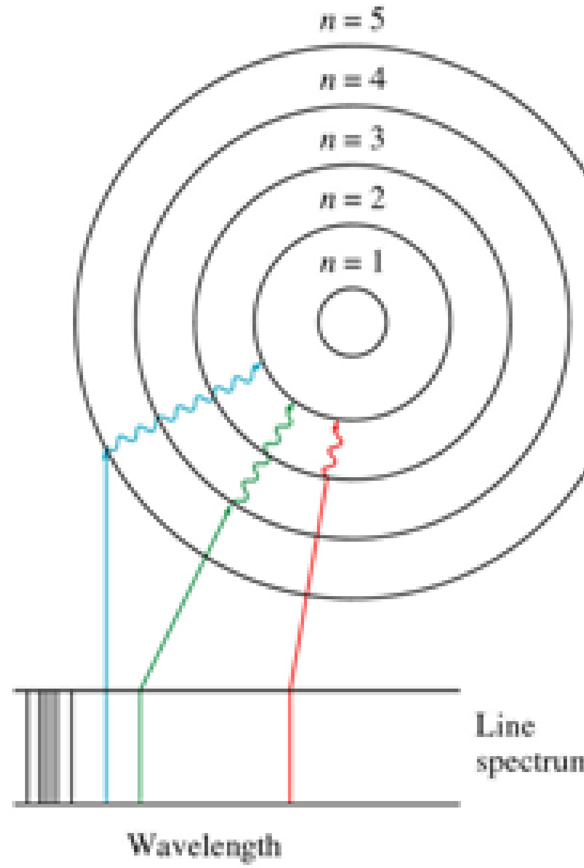
Important development: Bohr model



The electron emits or absorbs the energy changing the orbits.

hydrogen (emission)spectrum

solar (absorption)spectrum



Important development: Bohr model

Bohr's model was a pioneering, “quantized” picture of atomic energy levels, but it has its limits:

- it doesn’t work for multi-electron atoms (because it do not experience electron correlation effects),
- the “electron racetrack” picture is incorrect.

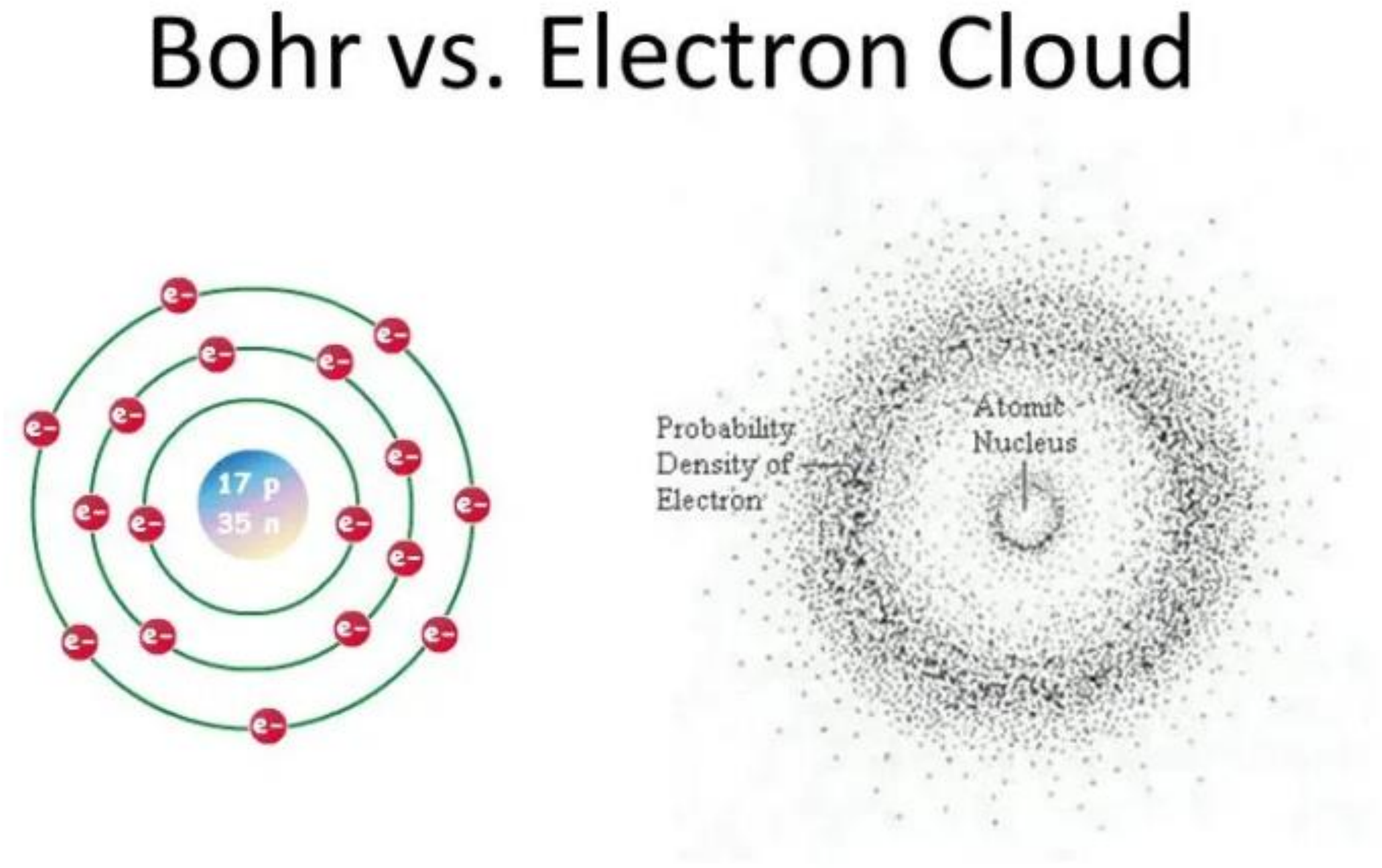
It works for Hydrogen-like atoms: H, He^{+}, Li^{2+} , Be^{3+}, …

Next step: the so called cloud model

Bohr's model treats electron energy levels as clearly defined orbital paths around the nucleus (like planets orbit the Sun). From this aspect it is close to the Rutherford model.

The cloud model treats the energy levels as probability clouds,
i.e. regions in which electrons are likely to be found.
The shapes of the clouds are based on the shapes formed by electrons that are trapped like standing waves.

Important development: cloud model



atomic physics – historical developments (16/20)

Important development – de Broglie waves:

In 1924 Louis de Broglie formulated a fundamental idea:
„If radiation which is basically a wave can exhibit particle nature under certain circumstances, and since nature likes symmetry, then entities which exhibit particle nature ordinarily, should also exhibit wave nature under suitable circumstances".

A wave is described by frequency, wavelength, phase velocity and intensity. It is spread out and occupies a relatively large region of space

A particle is specified by mass, velocity, momentum and energy. A particle occupies a definite position in space. In order for that it must be small.

wavelength

$$\lambda = \frac{h}{p}$$

momentum

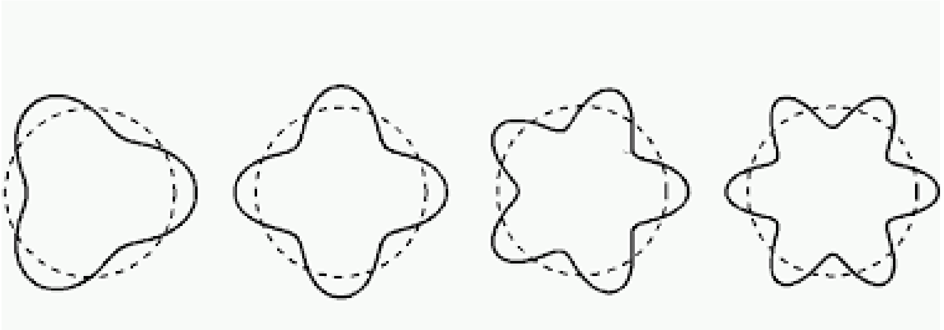
λ [m], p [kg·m·s^{-1}]
h [kg·m^2·s^{-1}]

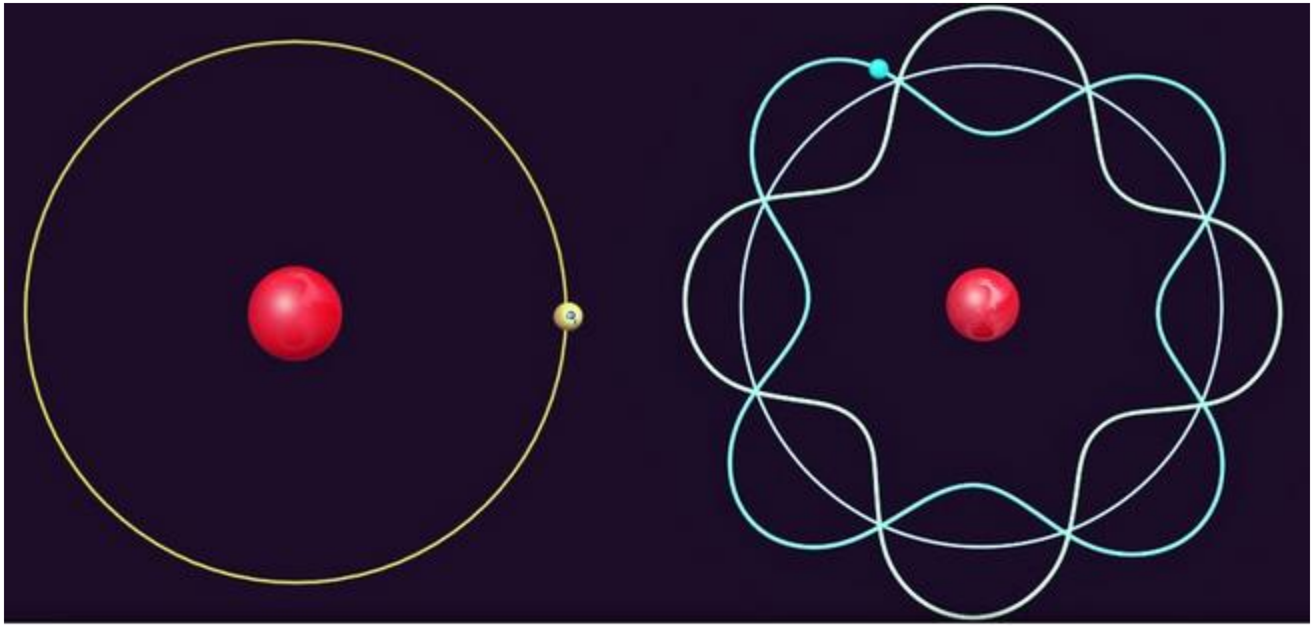
The wave associated with the matter particle is called Matter Wave.
The wavelength associated is called de Broglie Wavelength.

atomic physics – historical developments (17/20)

Important development – de Broglie waves:

In the case of electrons, these are so called standing waves:



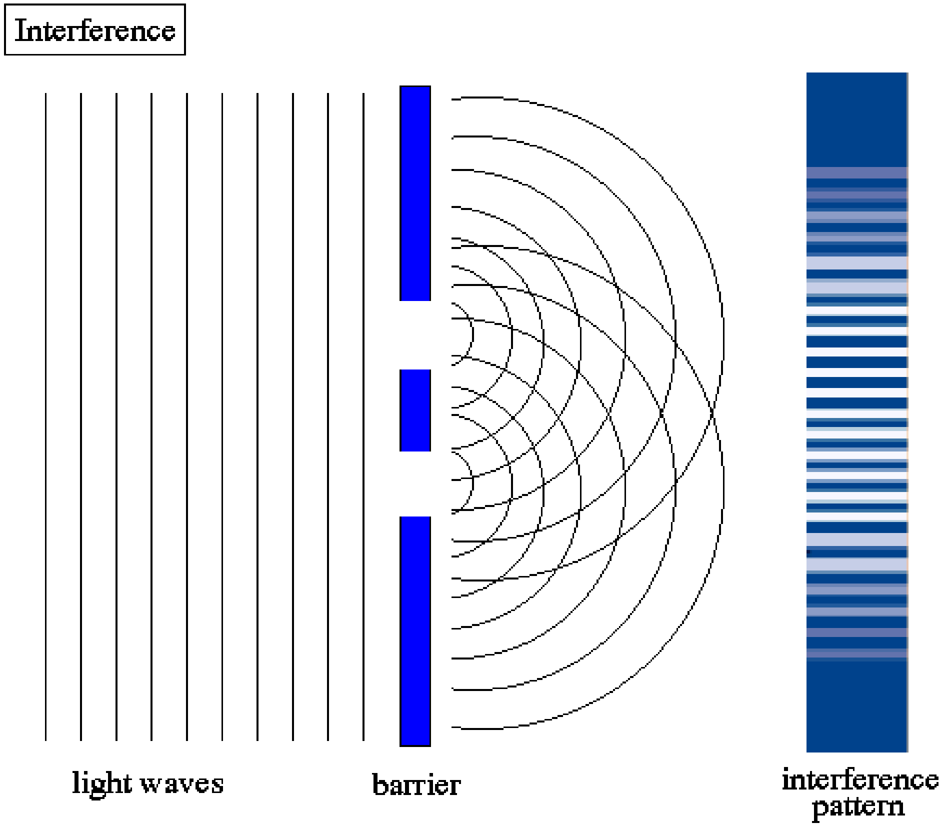


atomic physics – historical developments (18/20)

Important development – de Broglie waves:

One famous experiment (discussed over decades) is the so called **two-slit experiment** is showing the wave character of the light – the well known experiment of T. Young in 1801.

It can be applied also on a ray (beam) of electrons.



good visualization: https://en.wikipedia.org/wiki/Double-slit_experiment

atomic physics – historical developments (19/20)

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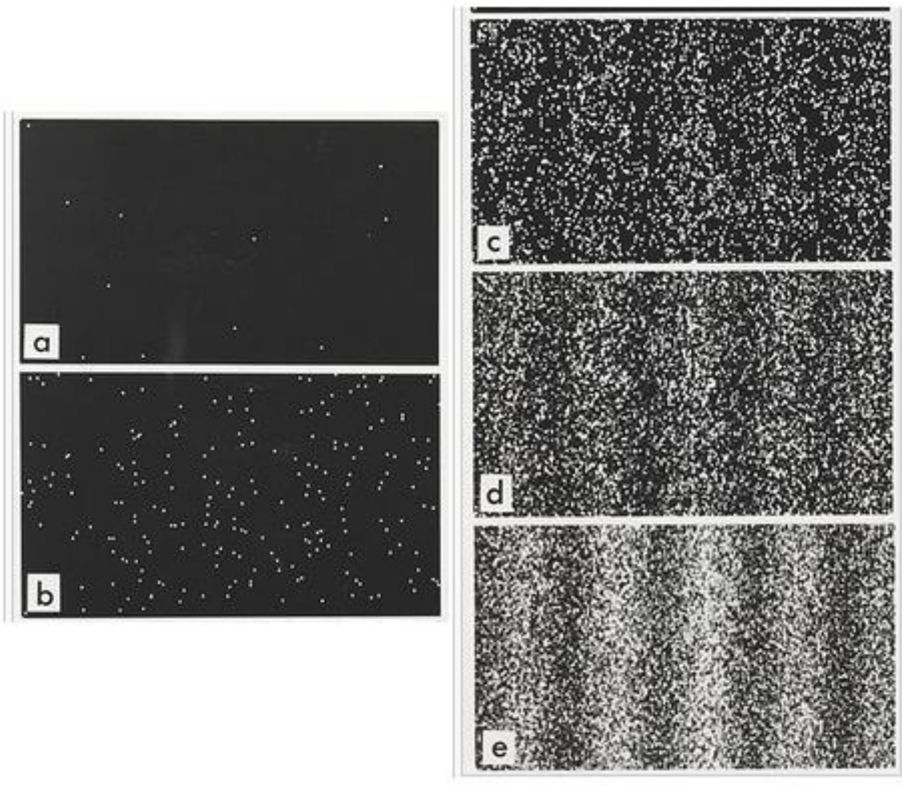
One famous experiment (discussed over decades) is the so called **two-slit experiment** is showing the wave character of the light – the well known experiment of T. Young in 1801.

It can be applied also on a ray (beam) of electrons.



visualization: https://www.youtube.com/watch?v=ZJ-0PBRuthc

atomic physics – historical developments (20/20)

Important development – duality:

Wave–particle duality is the concept that every elementary particle or quantic entity may be partly described in terms not only of particles, but also of waves.

Electron – duality properties:

Wave-like properties:

1. The electrons do not orbit the nucleus in the sense of a planet orbiting the sun, but instead exist as standing waves. The lowest possible energy an electron can take is therefore analogous to the fundamental frequency of a wave on a string. Higher energy states are then similar to harmonics of the fundamental frequency.
2. The electrons are never in a single point location, although the probability of interacting with the electron at a single point can be found from the wave function of the electron.

Particle-like properties:

1. There is always an integer number of electrons orbiting the nucleus.
2. Electrons jump between orbitals in a particle-like fashion. For example, if a single photon strikes the electrons, only a single electron changes states in response to the photon.
3. The electrons retain particle like-properties such as: each wave state has the same electrical charge as the electron particle. Each wave state has a single discrete spin (spin up or spin down).

Lecture 9: atomic physics

Content:

- atomic physics – basic historical developments
- models of atom (Thomson, Rutheford, Bohr)
- electrons, protons, neutrons
- first steps to quantum mechanics (Planck, Einstein)
- De Broglie waves, duality
- nuclear physics
- natural radioactivity, nuclear reactions

basics of nuclear physics

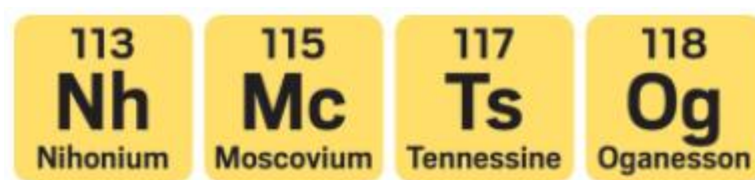
- nucleus contains nucleons: protons and neutrons
- atomic (proton) number Z = number of protons
- neutron number N = number of neutrons
- mass number A = number of nucleons = $Z + N$
- atomic mass = mass of nucleons + electrons = mass of nucleons

  (unit: Dalton, Da = 1⁄12 of the mass of an unbound neutral atom of carbon-12)
- each element has unique Z value
- isotopes of element have same Z, but different N and A values

notation in chemical reactions: $^{A}_{Z}\mathrm{X}$

notation in periodic table:

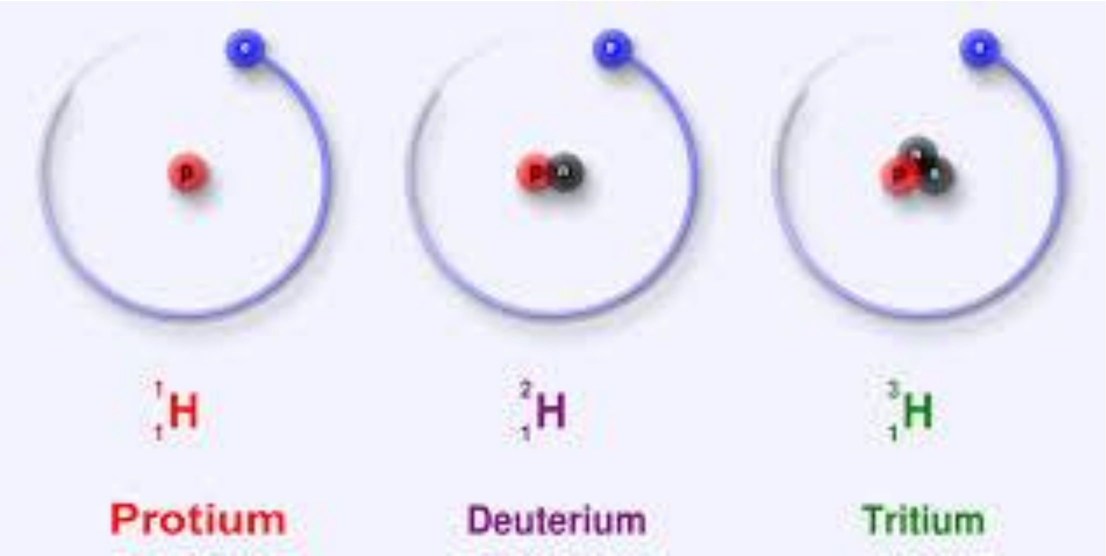
Comment:
… elements from the end
of the periodic table (actual situation):



basics of nuclear physics

Designation	Characteristics	Example
Isotopes	equal proton number	${}^{12}_{6}C$, ${}^{13}_{6}C$
Isotones	equal neutron number	${}^{13}_{6}C$, ${}^{14}_{7}N$
Isobars	equal mass number	${}^{17}_{7}N$, ${}^{17}_{8}O$, ${}^{17}_{9}F$

example:
isotopes of H:



nucleus charge and mass

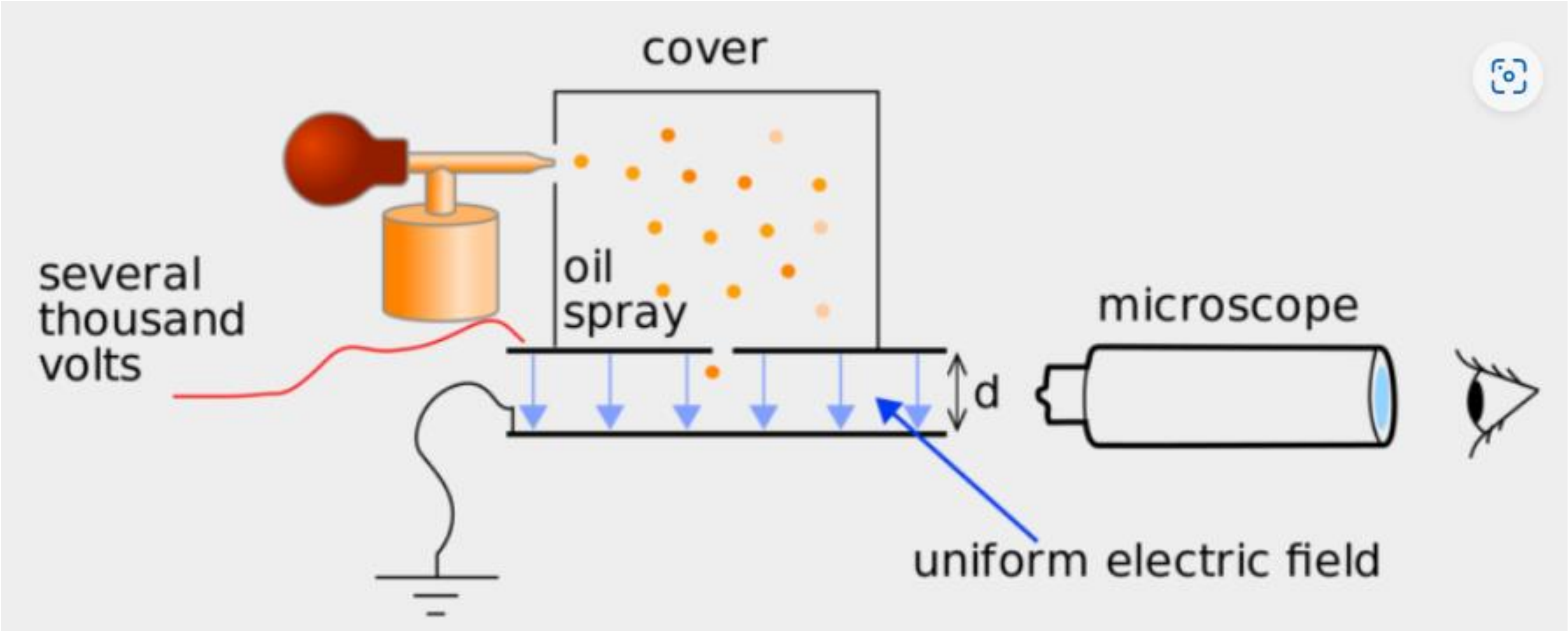
particle	charge	mass (kg)	mass (u)	mass (MeV/c^2)
proton	$+e$	$1.6726 \cdot 10^{-27}$	1.007 276	938.28
neutron	0	$1.6750 \cdot 10^{-27}$	1.008 665	939.57
electron	$-e$	$9.109 \cdot 10^{-31}$	$5.486 \cdot 10^{-4}$	0.511

- unified mass unit, u (or Da), defined using Carbon 12
- mass of 1 atom of $^{12}C \equiv 12$ u
- energy is not usually given in [J], but in [eV]

  (1 eV = energy gained by an electron when it is accelerated through a potential difference of 1 volt, 1 eV = $1.60217 \cdot 10^{-19}$ J)
- electrical charge of electron/proton is the elementary charge –e/+e (e=$1.60217 \cdot 10^{-19}$ C)

elementary charge

The first determination of the charge of an electron – so called **Millikan oil-drop experiment** (in 1909).
It is based on a fundamental theoretical law stating that the force acting on an electric charge in an electric field is equal to the product of the charge and the strength of that field.

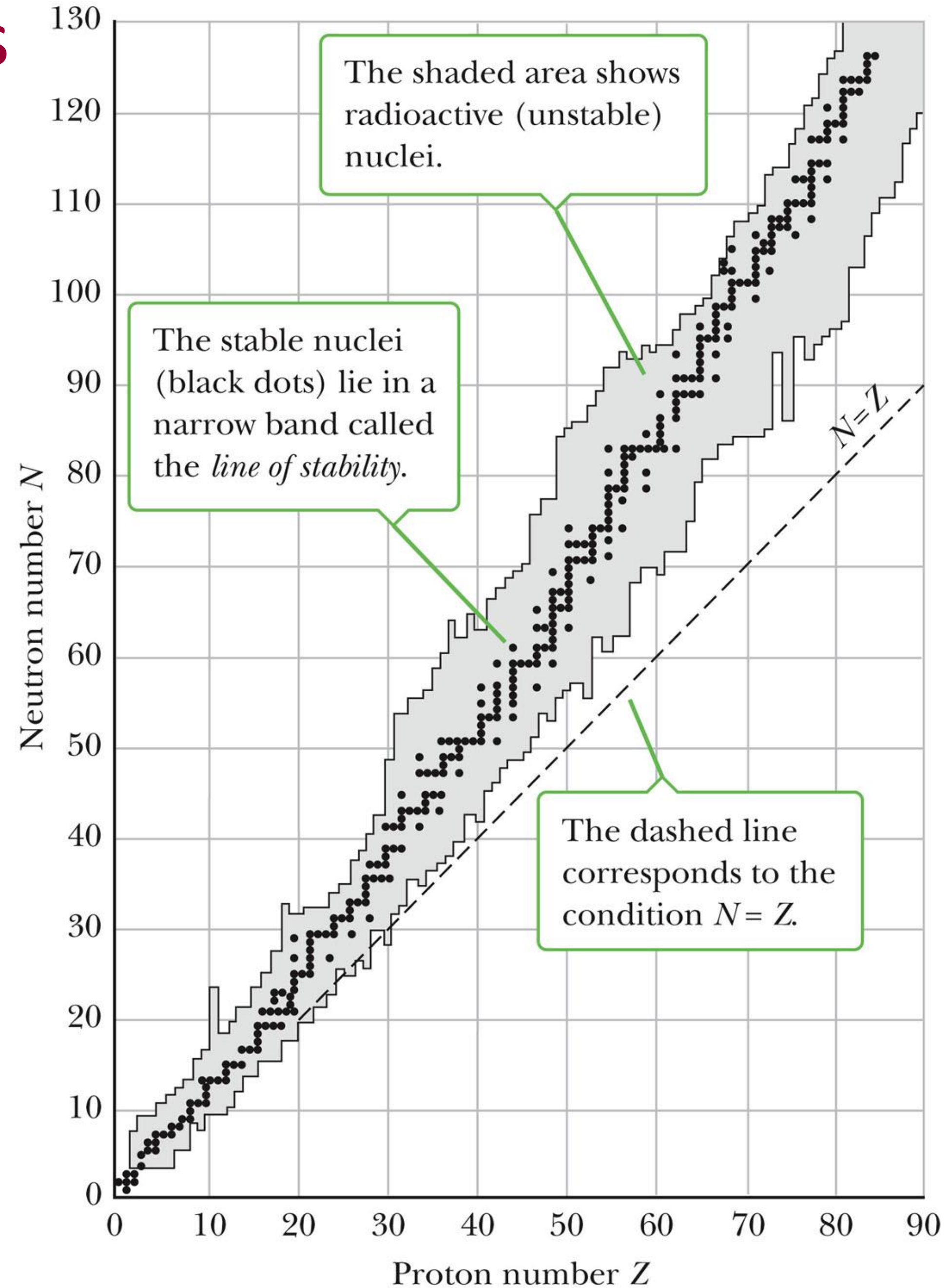


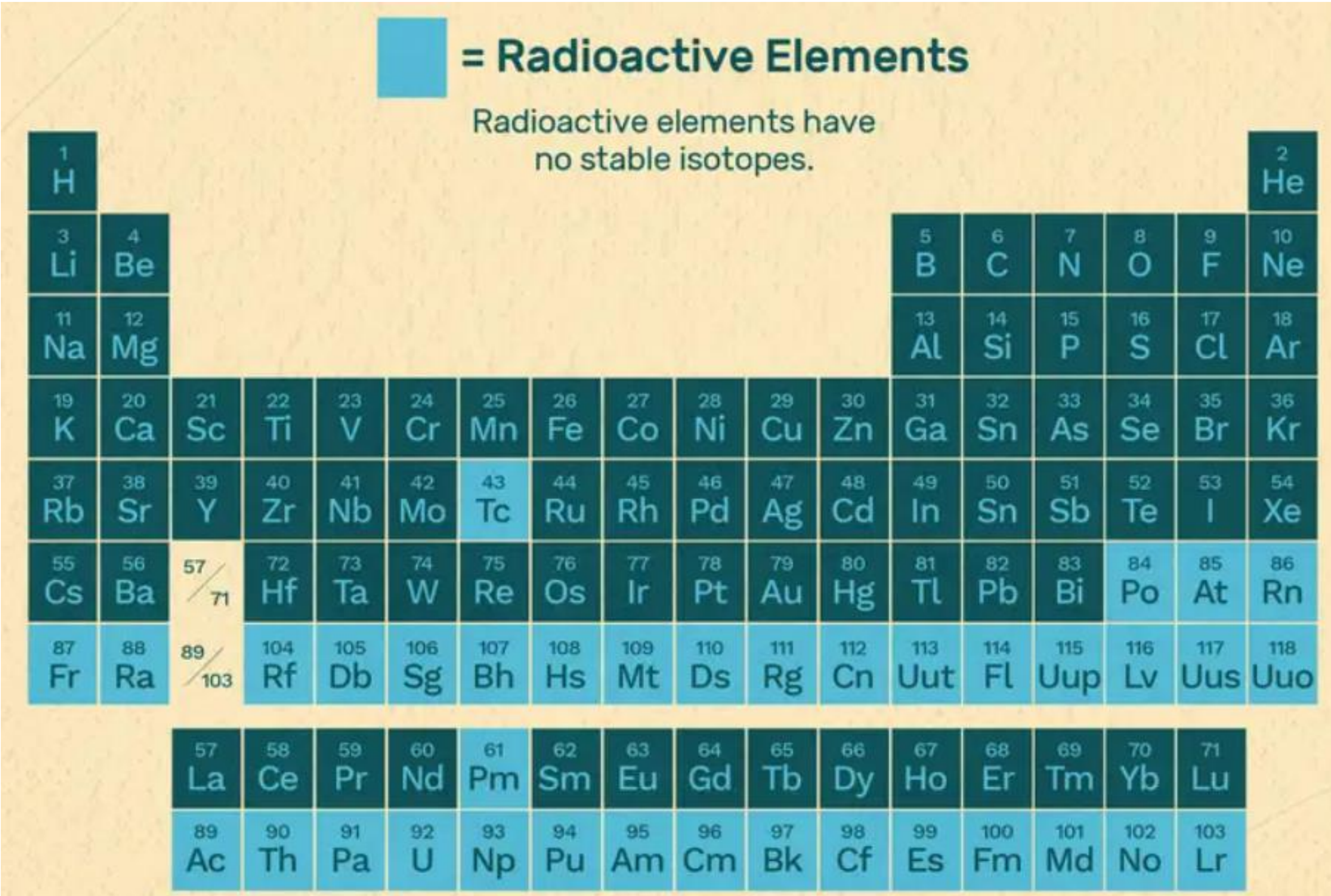
nice short explanation: https://www.youtube.com/watch?v=UFiPWv03f6g

stability of the nucleus

- an attractive nuclear force must balance the repulsive electric force
- this is called as the strong nuclear force
- neutrons and protons are affected by the strong nuclear force
- 260 stable nuclei
- If Z > 83 nuclei not stable

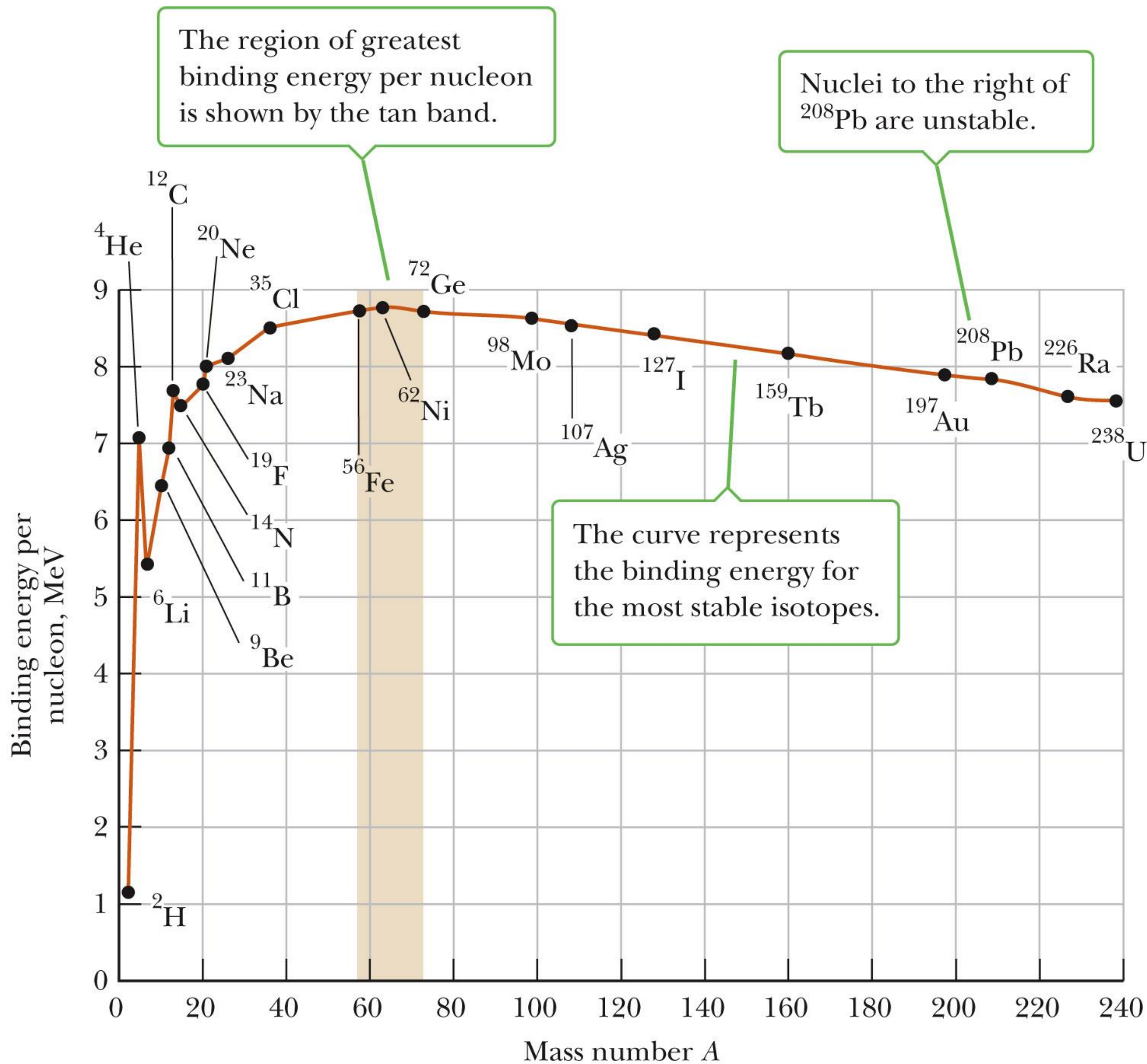
(one exception: Tc)




= Radioactive Elements
Radioactive elements have no stable isotopes.
1 H
2 He
3 Li
4 Be
5 B
6 C
7 N
8 O
9 F
10 Ne
11 Na
12 Mg
13 Al
14 Si
15 P
16 S
17 Cl
18 Ar
19 K
20 Ca
21 Sc
22 Ti
23 V
24 Cr
25 Mn
26 Fe
27 Co
28 Ni
29 Cu
30 Zn
31 Ga
32 Sn
33 As
34 Se
35 Br
36 Kr
37 Rb
38 Sr
39 Y
40 Zr
41 Nb
42 Mo
43 Tc
44 Ru
45 Rh
46 Pd
47 Ag
48 Cd
49 In
50 Sn
51 Sb
52 Te
53 I
54 Xe
55 Cs
56 Ba
57/71
72 Hf
73 Ta
74 W
75 Re
76 Os
77 Ir
78 Pt
79 Au
80 Hg
81 Tl
82 Pb
83 Bi
84 Po
85 At
86 Rn
87 Fr
88 Ra
89/103
104 Rf
105 Db
106 Sg
107 Bh
108 Hs
109 Mt
110 Ds
111 Rg
112 Cn
113 Uut
114 Fl
115 Uup
116 Lv
117 Uus
118 Uuo
57 La
58 Ce
59 Pr
60 Nd
61 Pm
62 Sm
63 Eu
64 Gd
65 Tb
66 Dy
67 Ho
68 Er
69 Tm
70 Yb
71 Lu
89 Ac
90 Th
91 Pa
92 U
93 Np
94 Pu
95 Am
96 Cm
97 Bk
98 Cf
99 Es
100 Fm
101 Md
102 No
103 Lr

binding energy

- total energy of nucleus is less than combined energy of individual nucleons
- difference is called the binding energy
- it is the energy required to separate nucleus into its constituents

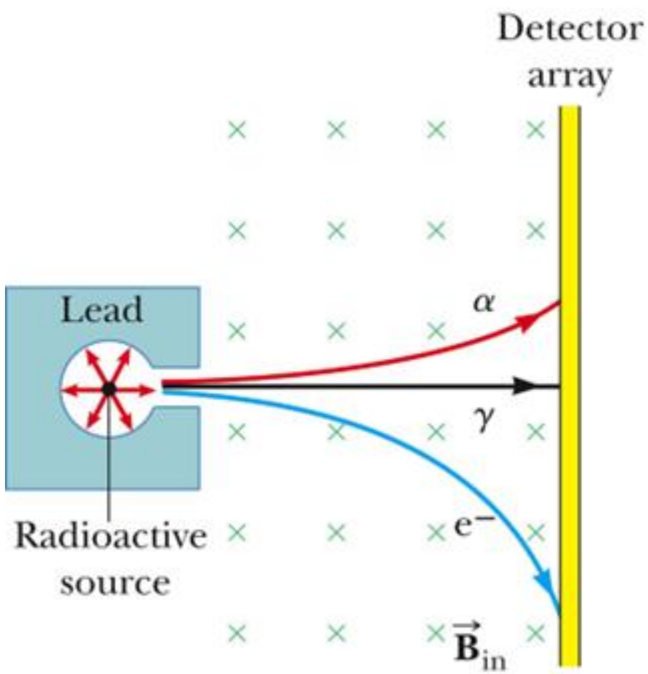


binding energy vs. mass number

radioactivity

Radioactivity = property exhibited by certain types of matter of emitting energy and subatomic particles spontaneously.

- unstable nuclei decay to more stable nuclei
- can emit 3 basic types of radiation (alpha, beta and gamma)

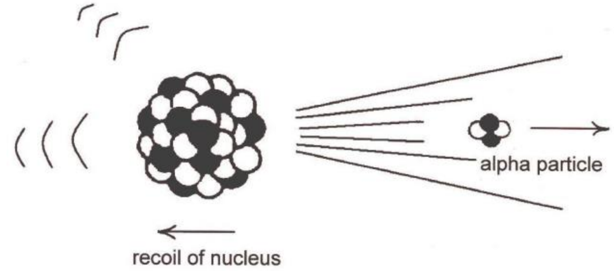


Alpha - these are fast moving helium atom nuclei. They have high energy, typically in the MeV range, but due to their large mass, they are stopped by just a few inches of air, or a piece of paper.

Beta - fast moving electrons or positrons. They typically have energies in the range of a few hundred keV to several MeV. They are able to penetrate further, through several feet of air, or several millimeters of plastic or less of very light metals.

Gamma - these are photons, just like light, except of much higher energy, typically from several keV to several MeV. X-Rays and gamma rays are the same kind of radiation, the difference is how they were produced. Depending on their energy, they can be stopped by a thin piece of aluminum foil, or they can penetrate several inches of lead.

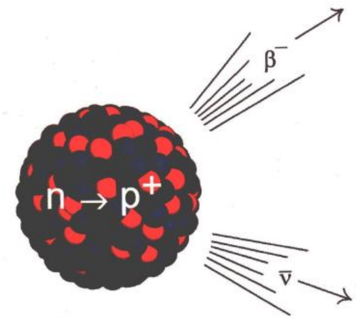
Alpha – α, Rutherford (1899) $${}_{Z}^{A}X \rightarrow {}_{Z-2}^{A-4}Y + {}_{2}^{4}He$$



Beta – β-, β+, Rutherford (1899)

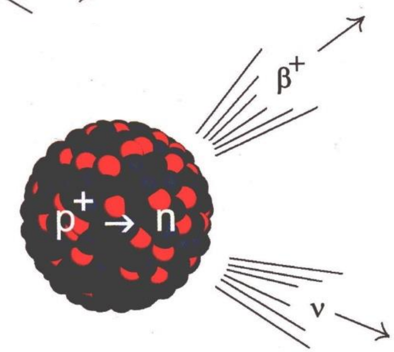
$${}_{Z}^{A}X \rightarrow {}_{Z+1}^{A}Y + {}_{-1}^{0}e + \tilde{\nu}, \qquad n \rightarrow p + {}_{-1}^{0}e + \tilde{\nu},$$

where ${}_{-1}^{0}e$ *is electron and* $\tilde{\nu}$ *is antineutrino*



$${}_{Z}^{A}X \rightarrow {}_{Z-1}^{A}Y + {}_{1}^{0}e + \nu, \qquad p \rightarrow n + {}_{1}^{0}e + \nu,$$

where ${}_{1}^{0}e$ *is positron (anti – electron) and* ν *is neutrino*



Gamma – γ, Villard (1900) $${}_{Z}^{A}X^{*} \rightarrow {}_{Z}^{A}X + \gamma$$

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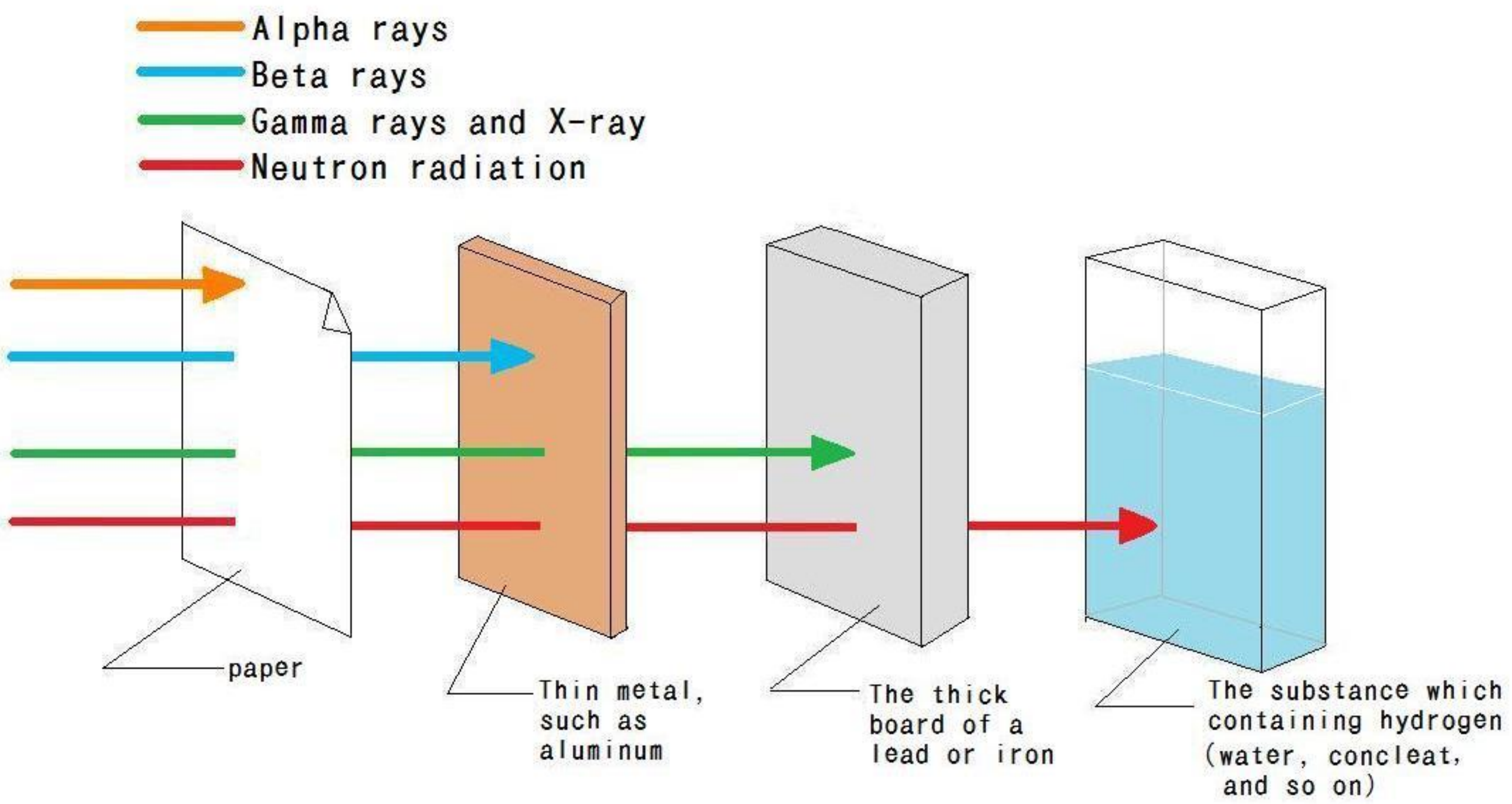
neutron radiation – n, Chadwick (1932)

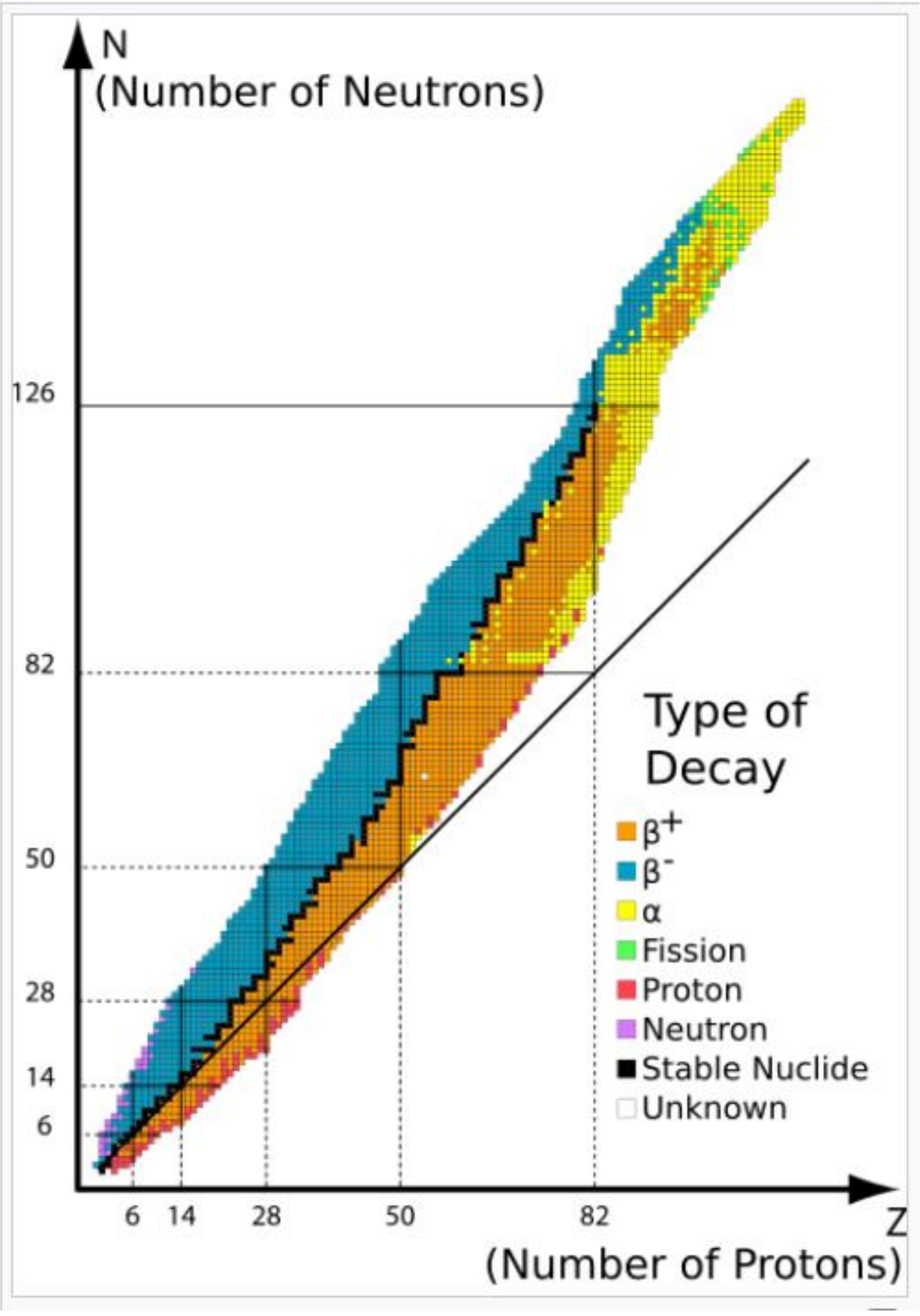
fission decay – certain unstable nuclei of heavier elements split into two nearly equal fragments (nuclei of lighter elements)

radioactivity

penetration of different kind of radiations



radioactivity



decay constant and half-life

- decay rate (aka activity) is number of decays per second
- λ is the decay constant
- unit is Curie (Ci) or Becquerel (Bq)
- decay is exponential decrease:

$$R = \left|\frac{\Delta N}{\Delta t}\right| = \lambda N$$

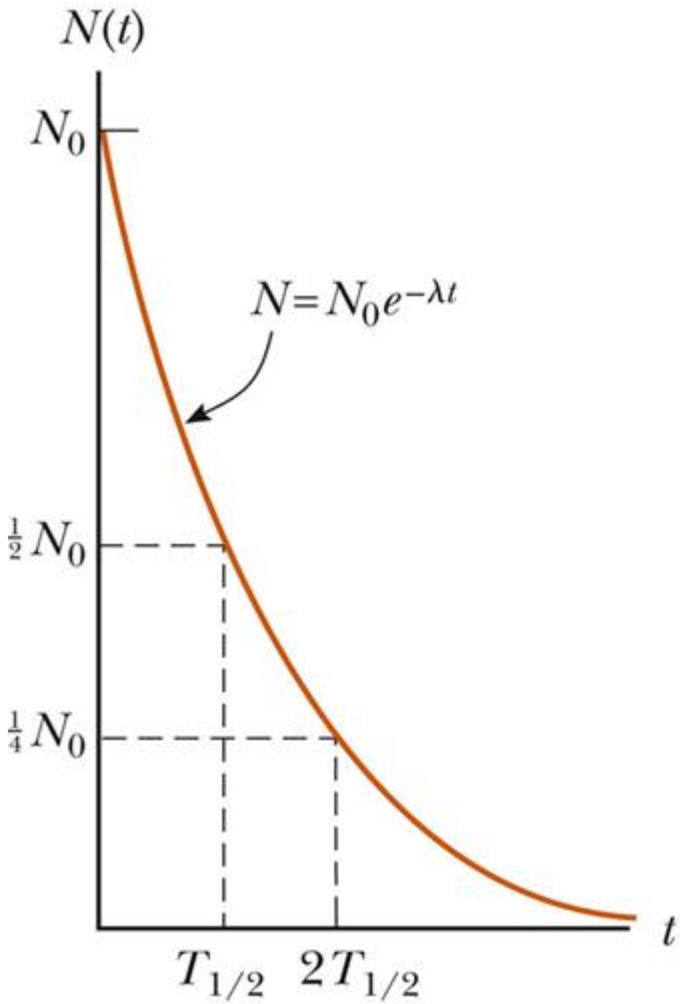
$$\boxed{N = N_0 e^{-\lambda t}}$$

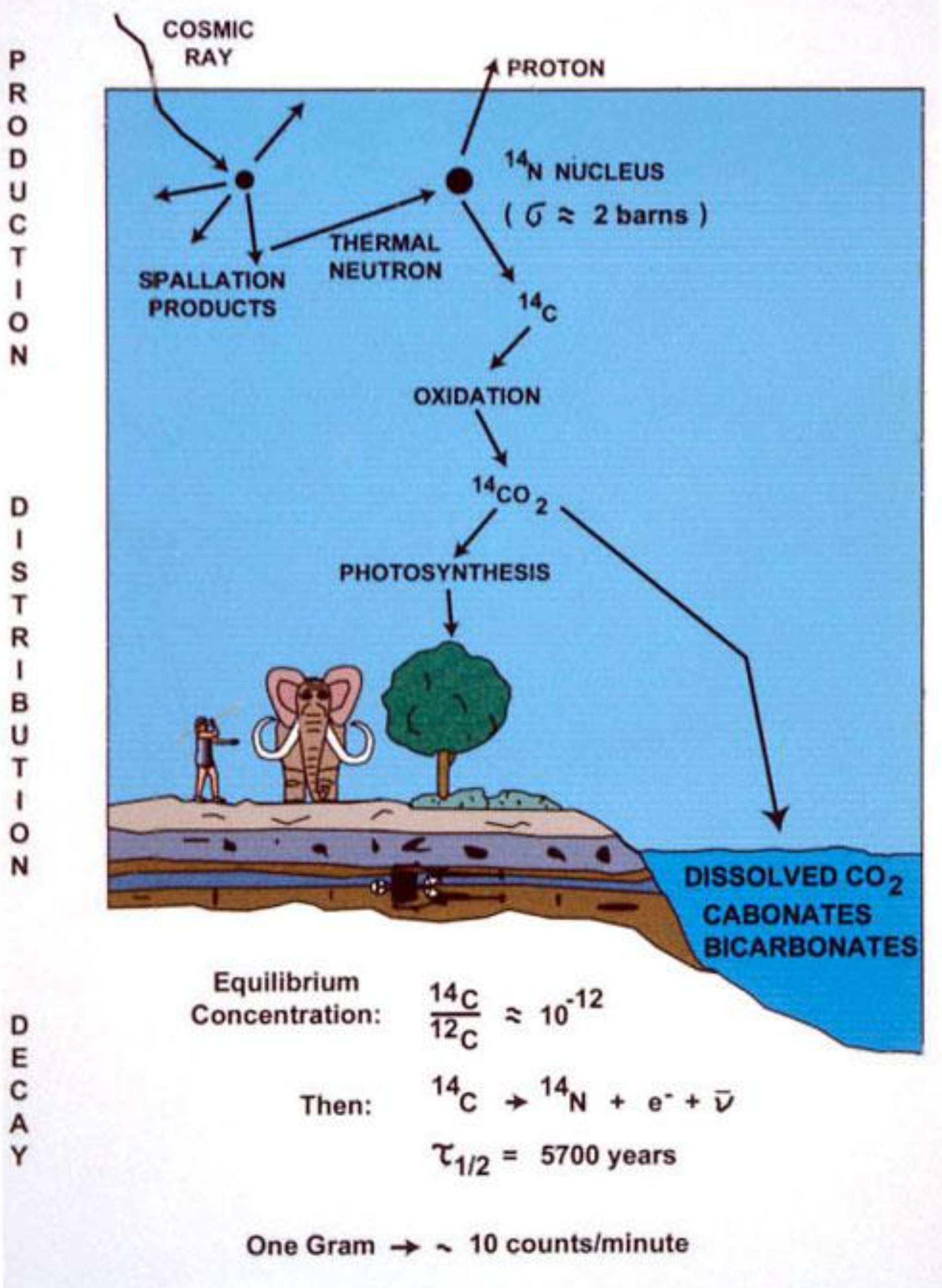
- half-life $T_{1/2}$ is time it takes for half of the sample to decay

$$1\,\text{Ci} \equiv 3.7\times 10^{10}\ \text{decays/s}$$

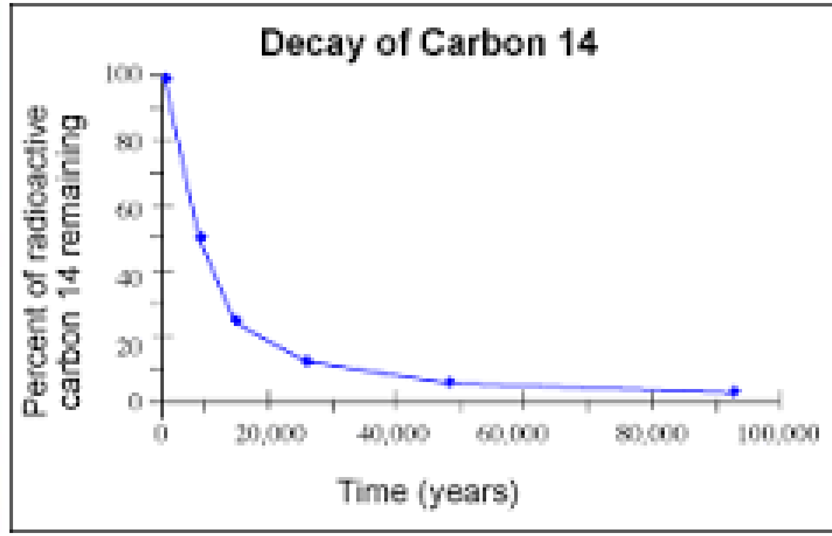
$$1\,\text{Bq} = 1\,\text{decay/s}$$

$$T_{1/2} = \frac{\ln 2}{\lambda} \doteq \frac{0.693}{\lambda} \quad \Rightarrow \quad \lambda = \frac{\ln 2}{T_{1/2}}$$





radioactive carbon dating

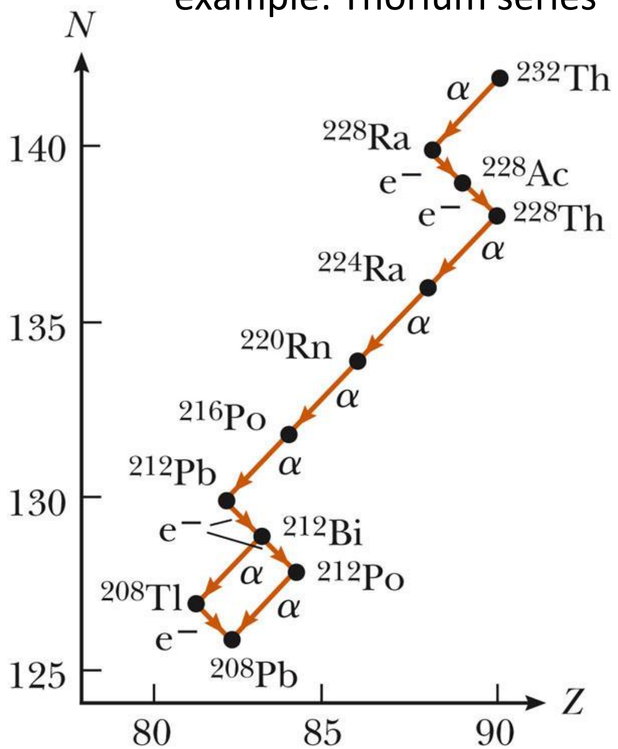


Radiocarbon dating is generally limited to dating samples no more than 50,000 years old, as samples older than that have insufficient ^{14}C to be measurable.

natural radioactivity

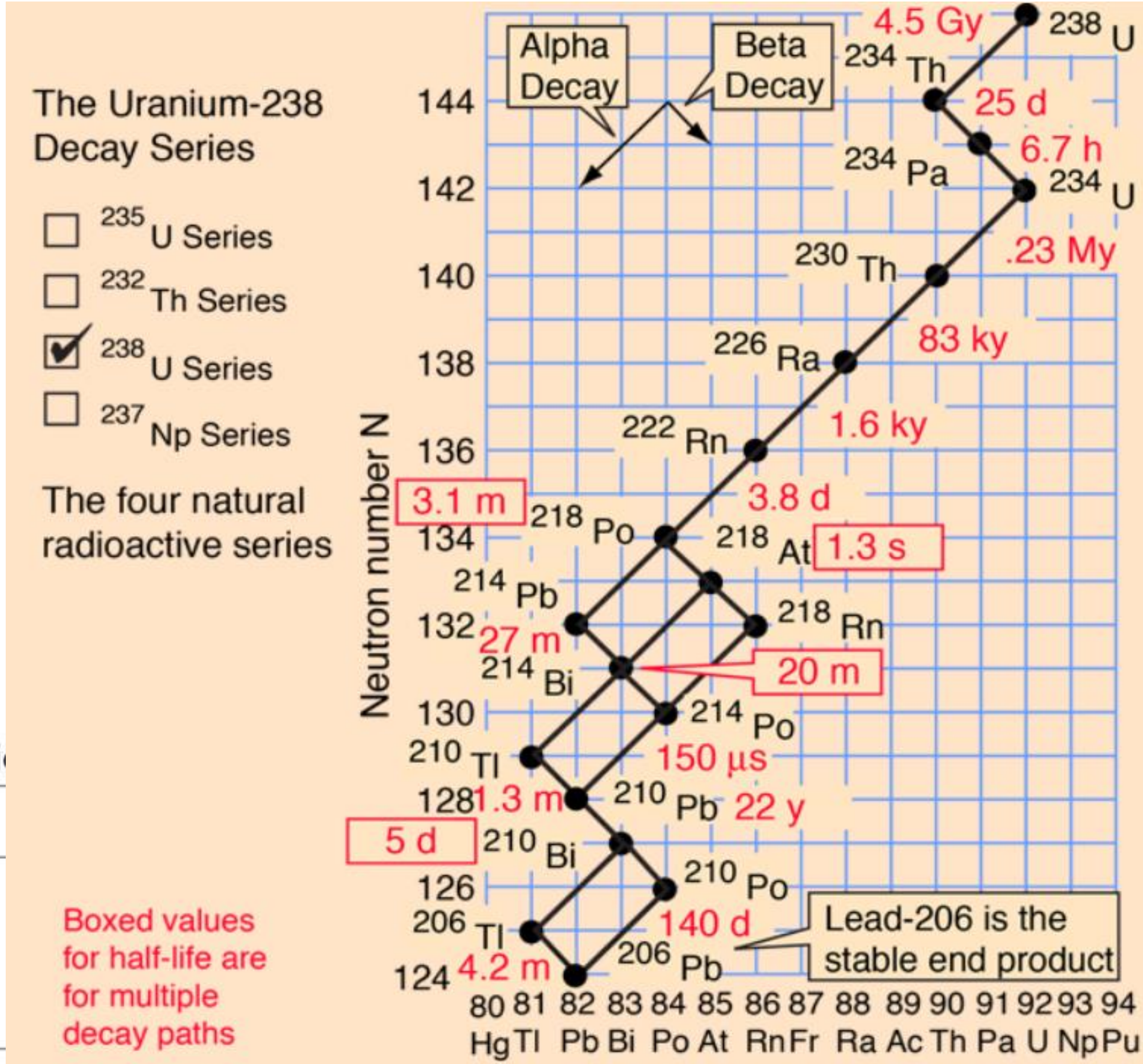
- three series of naturally occurring radioactivity,
- ^{232}Th more plentiful than ^{238}U or ^{235}U,
- other series are artificially produced.

example: Thorium series



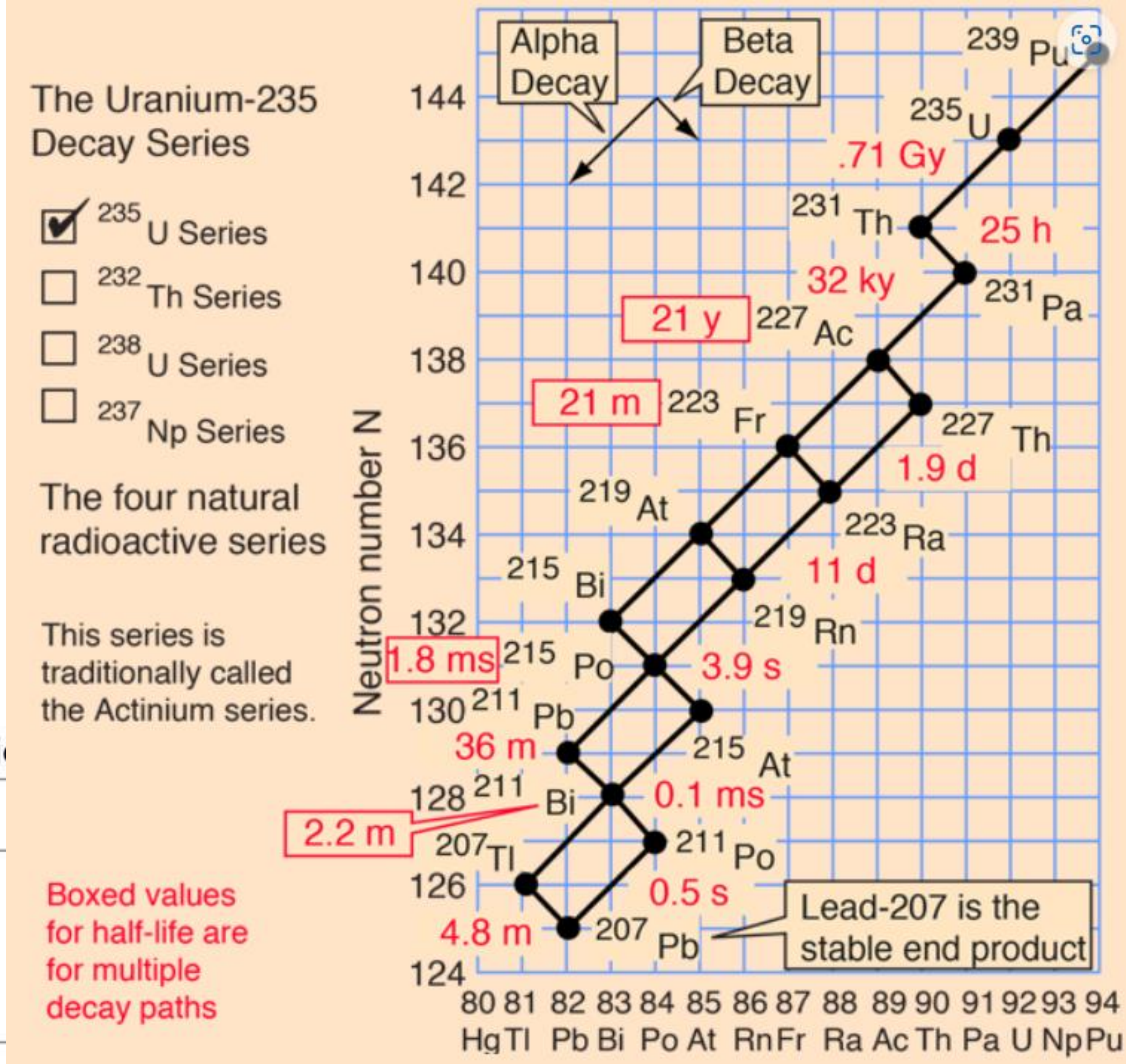
The Four Radioactive Series

Series	Starting Isotope	Half-life (years)	Stable End Product
Uranium	$^{238}_{92}U$	4.47×10^9	$^{206}_{82}Pb$
Actinium	$^{235}_{92}U$	7.04×10^8	$^{207}_{82}Pb$
Thorium	$^{232}_{90}Th$	1.41×10^{10}	$^{208}_{82}Pb$
Neptunium	$^{237}_{93}Np$	2.14×10^6	$^{209}_{83}Bi$



The Four Radi

Series
Uranium
Actinium
Thorium
Neptunium



The Four Radi

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The Four Radi

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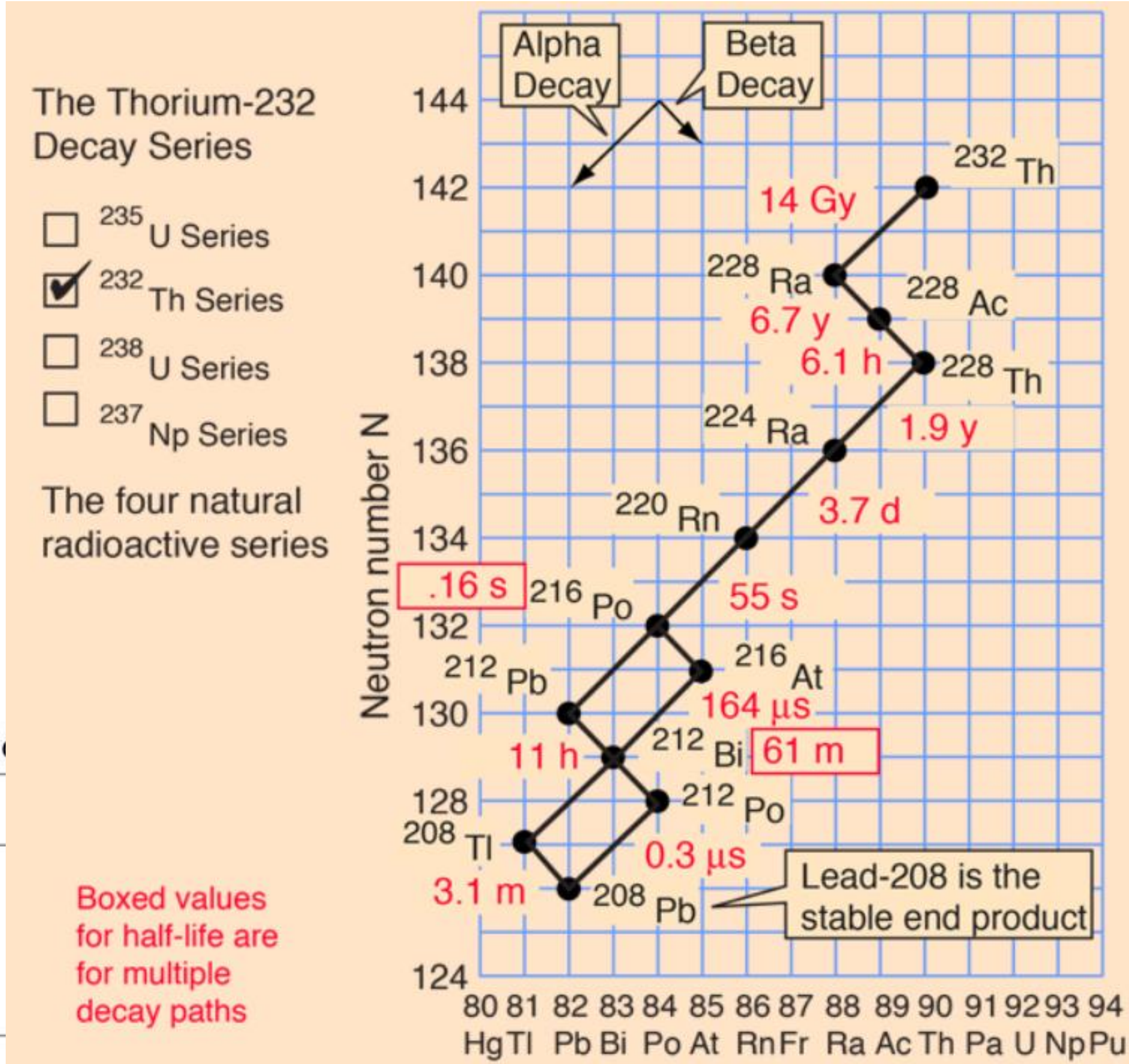
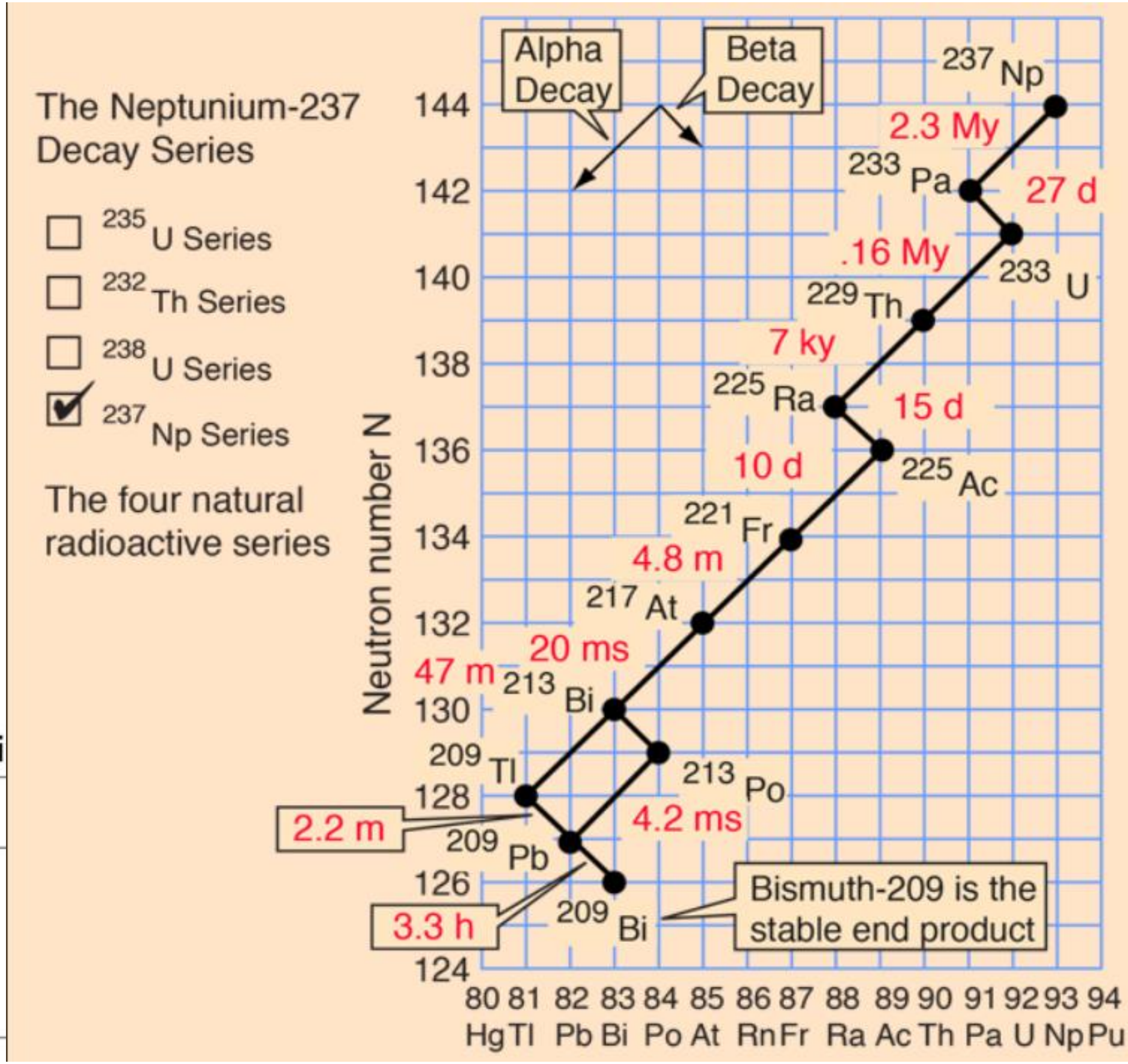




The Four Radi

Series
Uranium
Actinium
Thorium
Neptunium

nuclear reactions

In nuclear physics and nuclear chemistry, a nuclear reaction is the process in which two nuclei, or else a nucleus of an atom and a subatomic particle (such as a proton, neutron, or high energy electron) from outside the atom, collide to produce one or more nuclides that are different from the nuclide(s) that began the process.

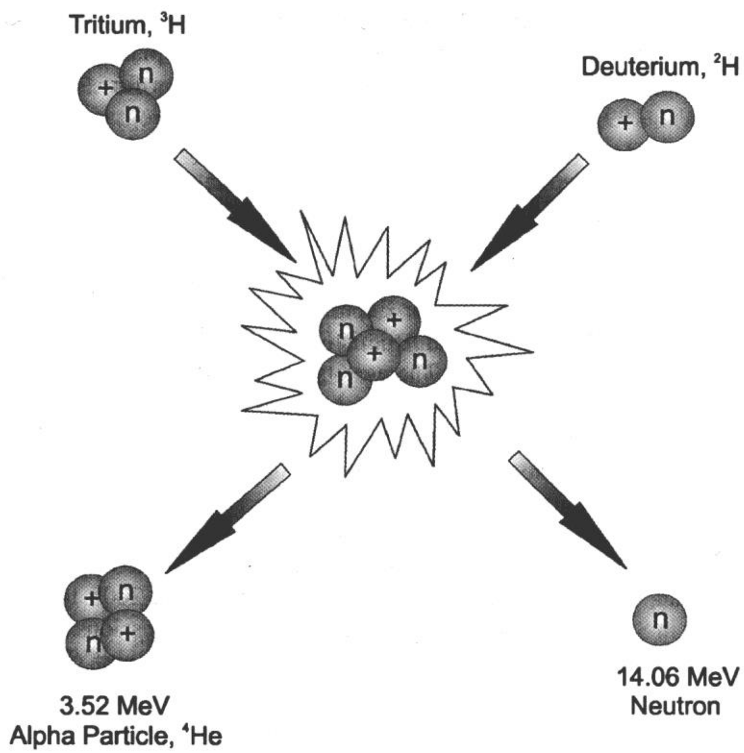
- bombard a nucleus with energetic particles

  (accelerators can generate particle energies up to 1 TeV),
- nucleus captures the particle,
- result is fission or fusion.

History: In 1932 at Cambridge University, a fully artificial nuclear reaction and nuclear transmutation was achieved by Rutherford's colleagues John Cockcroft and Ernest Walton, who used artificially accelerated protons against lithium-7, to split the nucleus into two alpha particles.

fusion

Fusion occurs when **two atoms slam together to form a heavier atom**, like when two hydrogen atoms fuse to form one helium atom.
This is the same process that powers the sun and creates huge amounts of energy - several times greater than fission.
It also doesn't produce highly radioactive fission products.



$$ {}_1^2H + {}_1^3H \rightarrow {}_2^4He + {}_0^1n $$

The larger nucleus has a greater binding energy and less mass per nucleon than the two that are combined. Thus mass is "destroyed" in the fusion reaction, and energy is released

fusion

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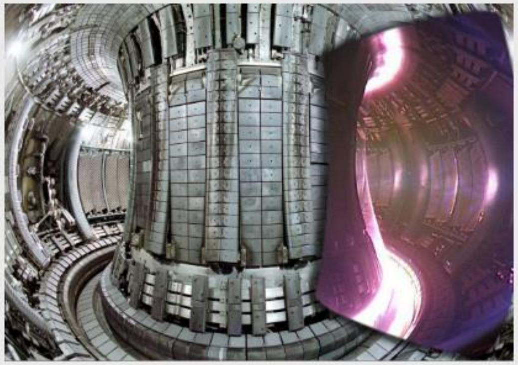


A small Sun on Earth

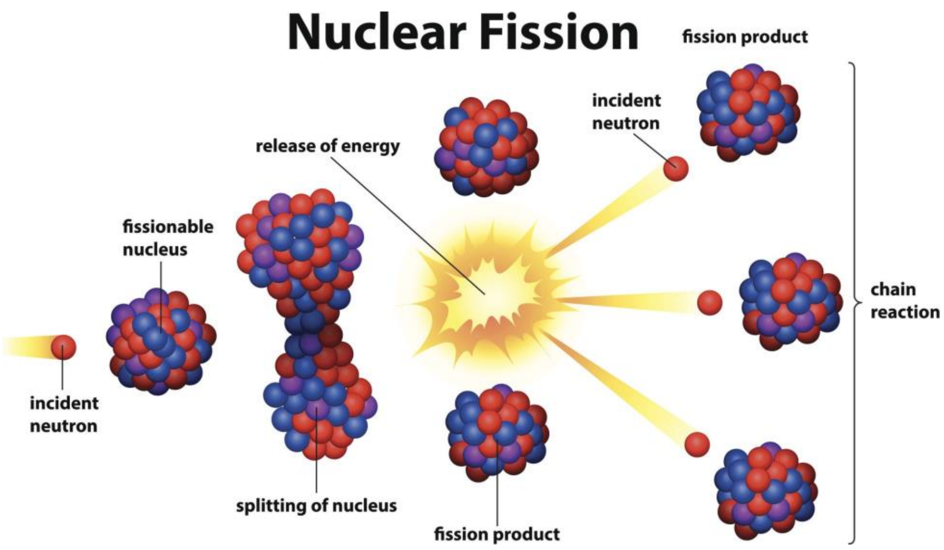
The quest to produce power from fusion is back on track after the test-bed reactor JET restarted successfully. But some serious challenges, which require heavy computing, still lie on the scientists' path between here and realising their dream. Scientists are installing extremely sensitive infra-red cameras to find tiny hot-spots in the walls of the container, which might show why and where energy is lost. And they must also tackle one of the most important unsolved problems of classical physics: turbulance.

The fusion record was achieved at the National Ignition Facility at California's Lawrence Livermore National Laboratory, which ignites fusion fuel with an array of 192 lasers.
These lasers reach high energies thanks in part to devices called preamplifiers.

fission

Fission occurs when a **neutron slams into a larger atom**, forcing it to excite and **split into two smaller atoms** – also known as fission products. Additional neutrons are also released that can initiate a chain reaction. When each atom splits, a tremendous amount of energy is released.

Uranium and plutonium are most commonly used for fission reactions in nuclear power reactors because they are easy to initiate and control.

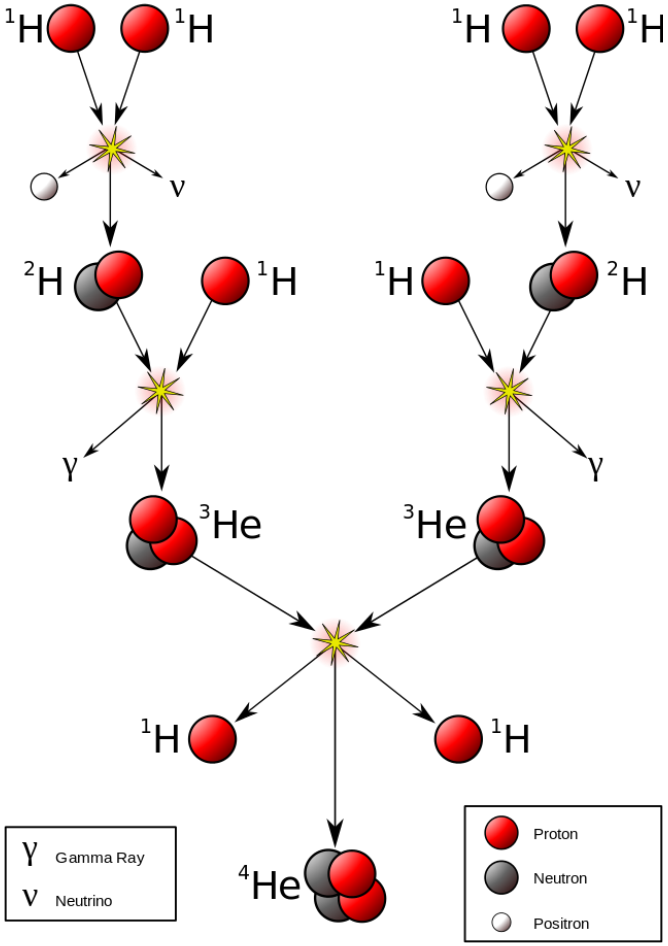


$$^{235}_{92}U + ^{1}_{0}n \rightarrow {}_{56}Ba + {}_{36}Kr + k\,^{1}_{0}n$$

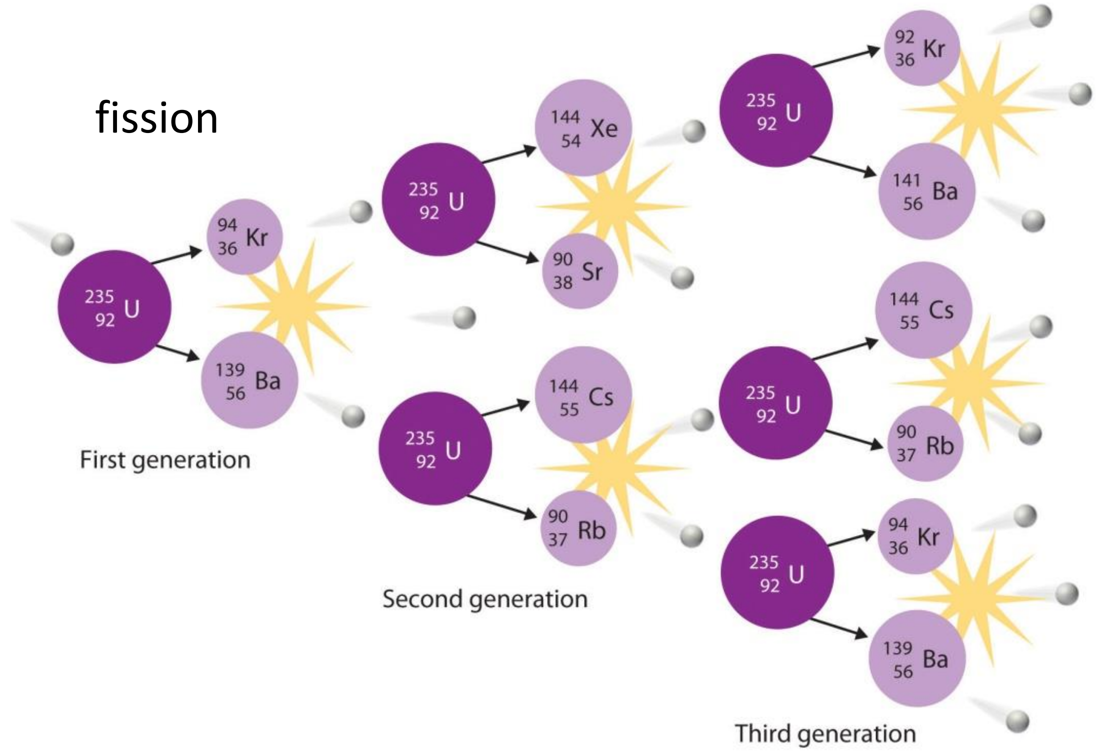
$$^{235}_{92}U + ^{1}_{0}n \rightarrow {}_{54}Xe + {}_{38}Sr + k\,^{1}_{0}n$$

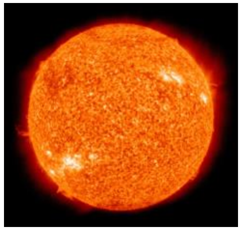
fusion and fission

fusion



fission





interaction of radiation with matter

- radioactive emissions can ionize atoms
- problems occur when these ions (e.g., OH^-, H^+) react chemically with other ions
- *genetic* damage affects reproductive cells
- *somatic* damage affects other cells (lesions, cataracts, cancer, fibrosis, etc.)

Quantifying Radioactivity

Quantity	*Definition*	*SI unit*	*Common Unit*
Activity	# nuclei that decay per sec	1 Bq ≡ 1 decay/s	$1\ \mathrm{Ci} = 3.70\times10^{10}$ Bq
Exposure (defined for X and γ rays **only**)	Ionization per kg	1 R ≡ amount of radiation that produces 2.58×10^{-4} C/kg	Roentgen (R)
Absorbed Dose (D)	Energy absorbed per kg	1 Gray (Gy) ≡ 1 J/kg	$1\ \mathrm{rad} = 10^{-2}$ Gy
Relative Biological Effectiveness (RBE)	How much more damage is done compared to X or γ rays of equivalent energy (unitless).		
Dose Equivalent (H)	Damage expected	1 Sv ≡ 1 RBE×Gy	$1\ \mathrm{rem} = 10^{-2}$ Sv

RBE Factors

Radiation Type	***RBE Factor***
X and γ rays	1.0
β particles	1.0–1.7
α particles	10–20
Slow n	4–5
Fast n and p	10
Heavy ions	20

In radiobiology, the **relative biological effectiveness** (often abbreviated as RBE) is the ratio of biological effectiveness of one type of ionizing radiation relative to another, given the same amount of absorbed energy.

sources of ionizing radiation

Sources to which average person in United States is exposed[a]	
(100% = an annual effective dose of about 3.6 mSv per person)	
From natural sources	
Radon	55%
Radioactive elements within body	11%
Rock, soil, and ground water	8%
Cosmic rays	8%
From artificial sources	
Medical X rays	11%
Nuclear medicine	4%
Consumer products	3%
Miscellaneous	$<1\%$

From Touger, *Introductory Physics*, Table 28-4, p. 817