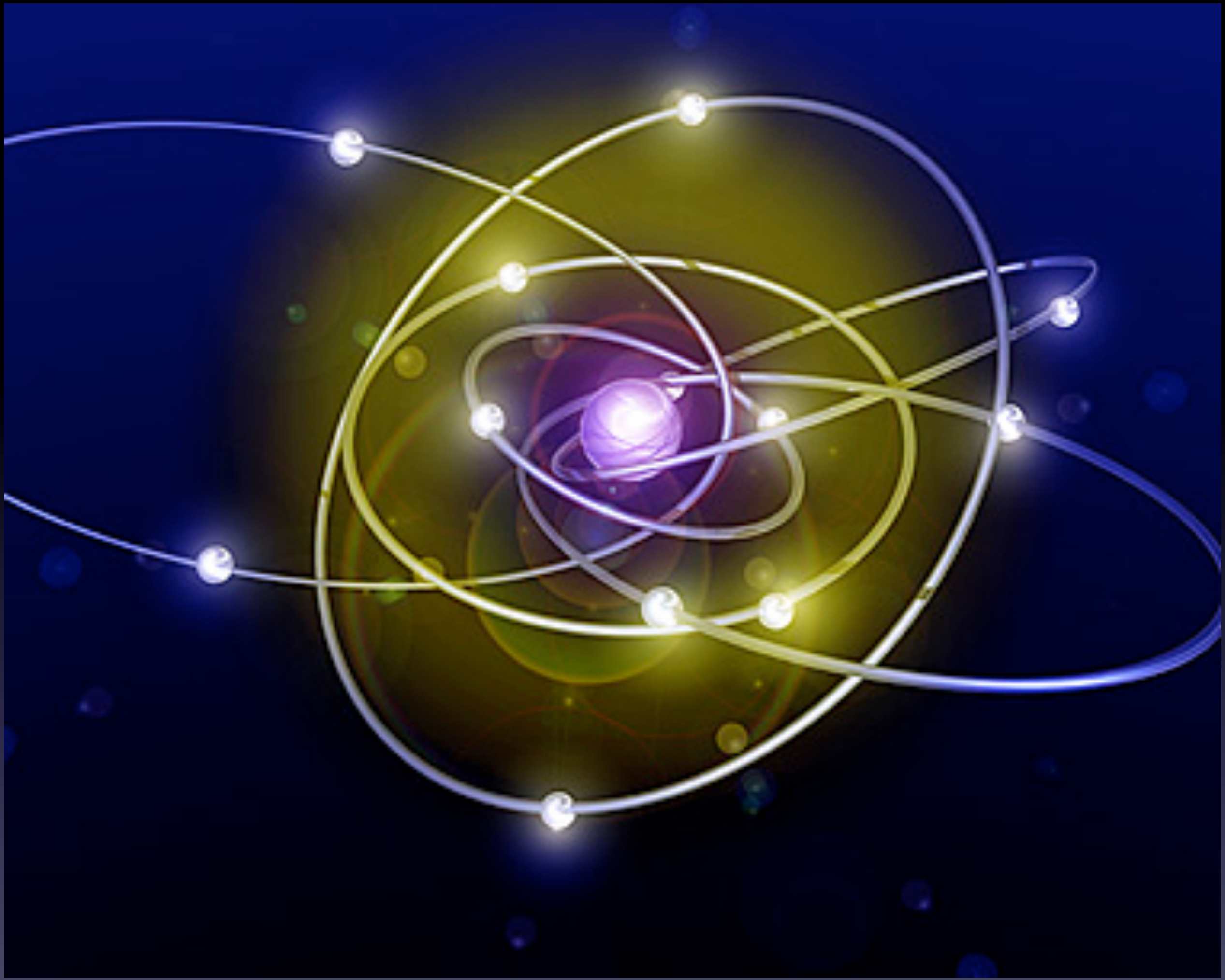
7. Atomic & Nuclear Physics

Topic Outline

Section	Recommended Time	Giancoli Sections
7.1 The Atom	2h	27.8, 27.9
7.2 Radioactive Decay	3h	30.1, 30.3-30.6, 30.8
7.3 Nuclear Reactions, Fission and Fusion	4h	30.2, 30.7, 31.1-31.4, 26.10



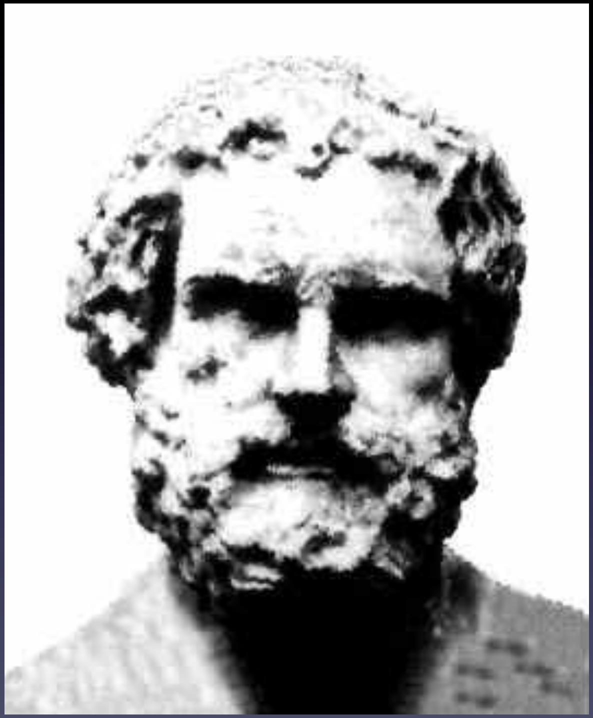
7.1 The Atom

Models of the Atom

Person and Dates	Model of the Atom	Diagram	Evidence
Democritus			
John Dalton			
Sir JJ Thomson			
Sir Ernest Rutherford			
Niels Bohr			

Democritus

(c.460-c.370 BC)



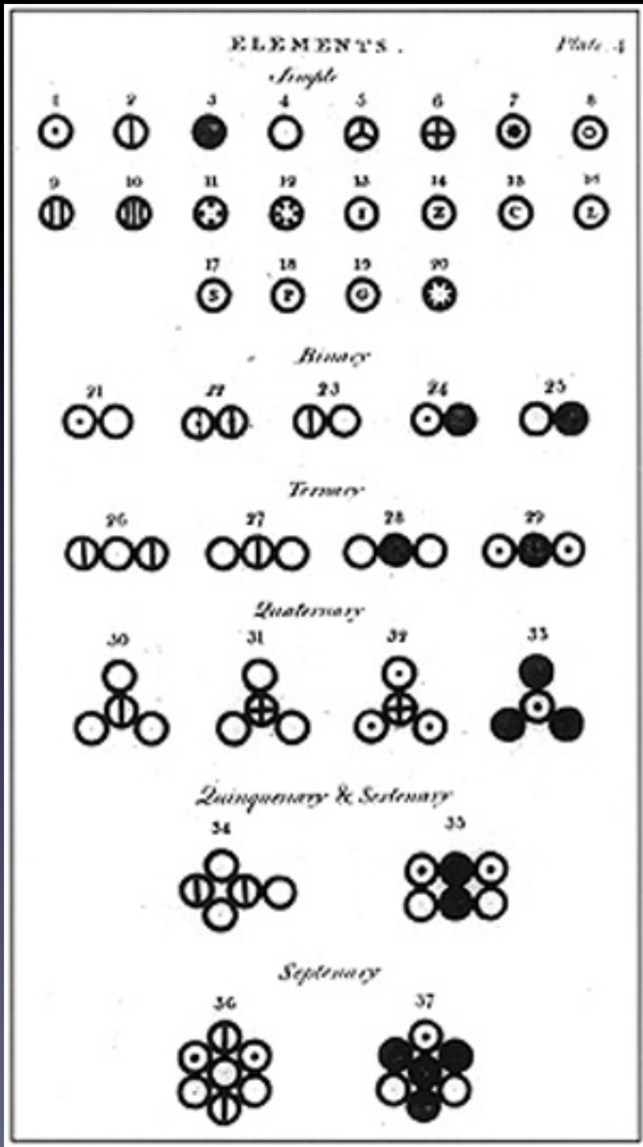
- Democritus was a philosopher in ancient Greece
- He thought that matter was made up of tiny particles too small to be divided
- The Greek word *atomos* means 'indivisible'

John Dalton

(1766-1844)

- Dalton was an English scientist
- He developed modern atomic theory
- His model of the atom is called the 'billiard-ball model'

Dalton’s Billiard-Ball Model



- All matter is made of tiny particles called atoms
- Atoms of the same element are identical
- Atoms can combine to form compounds
- Chemical reactions change the grouping of atoms, but not the atoms themselves

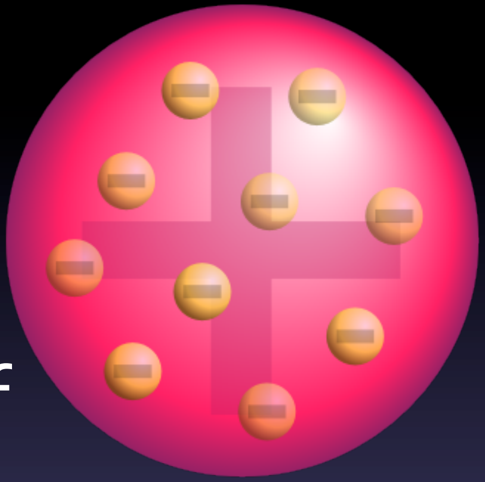
Sir J J Thomson

(1856-1940)

- Thomson was a British physicist
- He did experiments on cathode rays and discovered the electron
- In 1906, he was awarded a Nobel prize for his discovery
- His model of the atom is called the 'plum-pudding model'

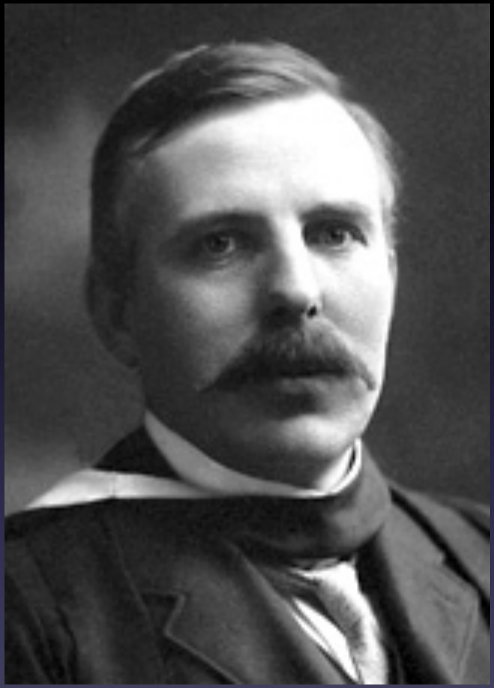
Thomson’s Plum-Pudding Model

- Thomson realised that negatively charged electrons could be removed from an atom
- He proposed that atoms consist of negatively-charged electrons (the ‘plums’) embedded in a positively-charged atom (the ‘pudding’)



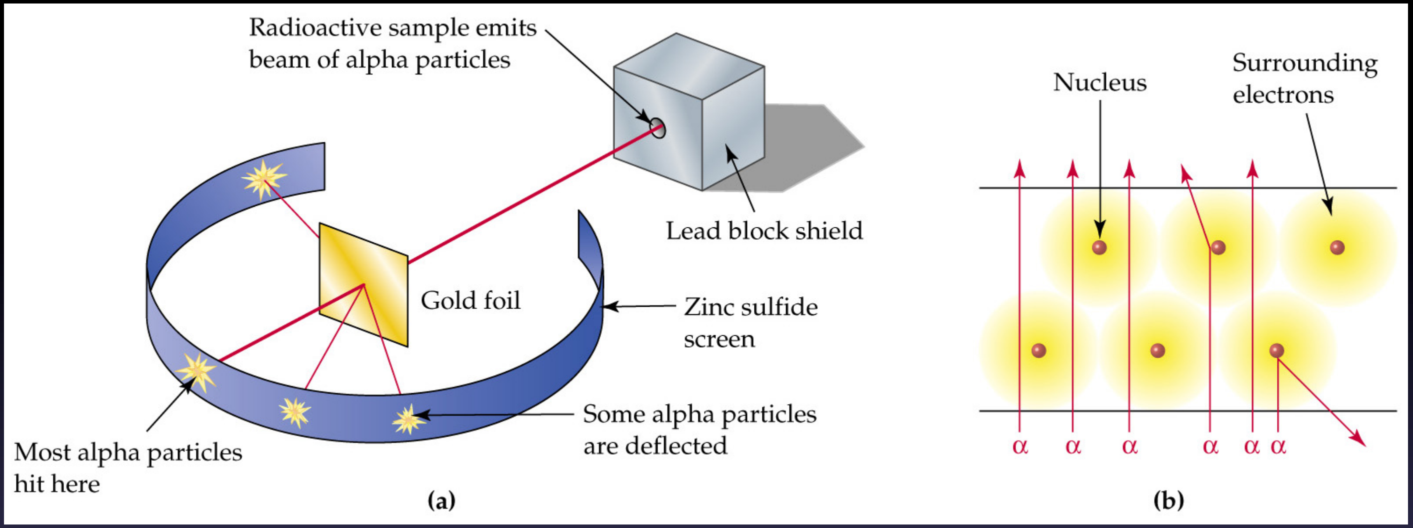
Sir Ernest Rutherford

(1871-1937)



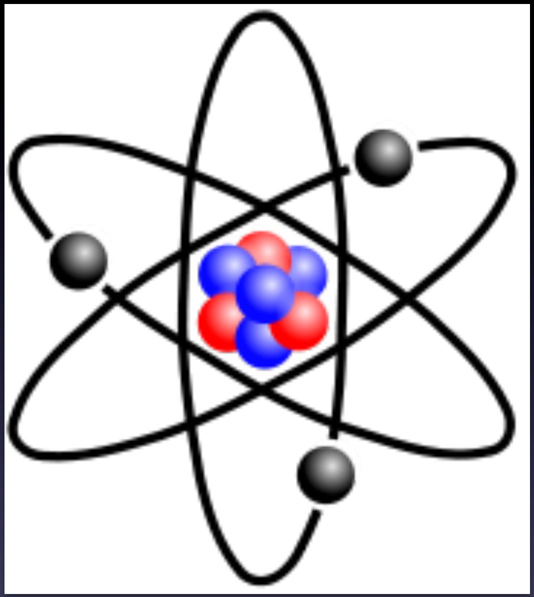
- Rutherford was a chemist from Nelson, New Zealand
- Based on the results of his gold-foil experiment, he proposed that most of the mass of an atom is concentrated in a central *nucleus*
- He won the Nobel Prize for Chemistry in 1908

Rutherford's Gold-Foil Experiment



- In the famous gold-foil experiment, Rutherford's students Geiger and Marsden fired alpha particles at a thin sheet of gold foil
- They found that some particles were deflected through large angles and some even bounced back

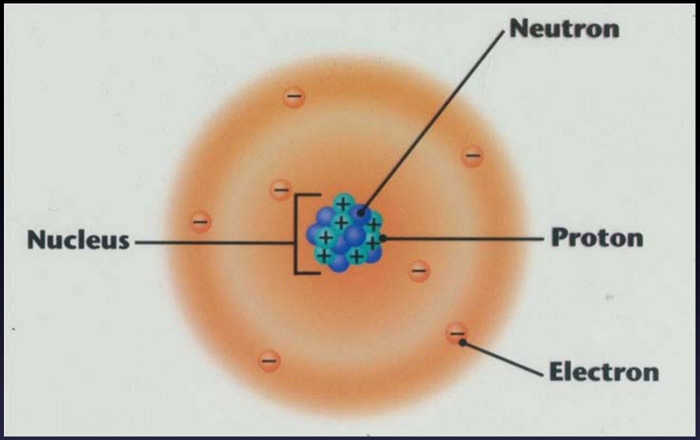
Rutherford's Nuclear Model



- Rutherford concluded that:
  - Most of the mas in an atom must be in a very small, positively-charged nucleus in the centre of the atoms
  - Electrons orbit around this central nucleus
  - There is a basic unit of positive charge in the nucleus, which he called the proton

Further Developments

- **Niels Bohr** (1885-1962) realised that the electrons could only occupy fixed orbits around the nucleus

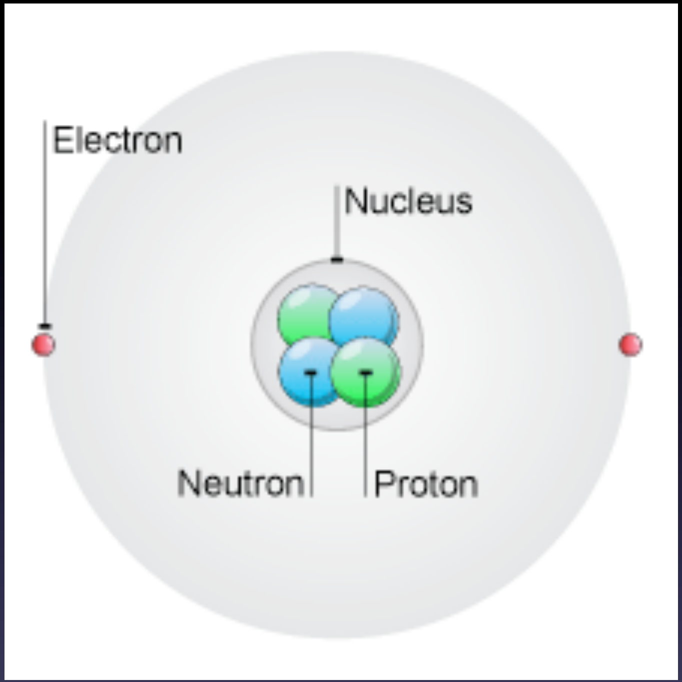


- **Louis de Broglie** (1892-1987) proposed that electrons can be regarded as waves, resulting in an 'electron cloud' around the nucleus
- **Sir James Chadwick** (1891-1974) discovered the **neutron**, a nuclear particle with similar mass to a proton but no electrical charge

Nuclear Model of the Atom

Atoms consist of:

- **Protons**
  - positively charged ($+1.6 \times 10^{-19}$ C)
  - have a mass of about 1.67×10^{-27} kg
  - are found in the nucleus
- **Neutrons**
  - electrically neutral
  - have a mass of about 1.67×10^{-27} kg
  - are found in the nucleus
- **Electrons**
  - negatively charged (-1.6×10^{-19} C)
  - have a mass of about 9.11×10^{-31} kg (about 1/2000 the mass of a proton)
  - orbit around the nucleus in particular energy levels called shells or orbitals
- Protons and neutrons are collectively called **nucleons**



Nuclear Model of the Atom

- The **electrostatic attraction** between the protons and the electrons provides the centripetal force to keep the electrons orbiting around the nucleus
- The **strong nuclear force** binds the protons and neutrons together in the nucleus
  - This force overcomes the electrostatic repulsion (Coulomb force) between the protons
  - This force is only significant over very small distances, and keeps the nucleons in a nucleus about 1.3 fm (i.e. 1.3×10^{-15} m) apart

Nucleon Number

- The standard way of representing an atom shows the chemical symbol (or nuclide), atomic number (or proton number) and mass number (or nucleon number)
- For example:

$$^{23}_{11}Na$$

- This atom of sodium (Na) has an atomic number of 11 (i.e. 11 protons) and a mass number of 23 (i.e. 23 nucleons); since 23 - 11 = 12, the atom has 12 neutrons
- The generalised diagram for this is:

$$^{A}_{Z}X$$

A = mass number (or nucleon number) = number of protons + neutrons

X = element symbol (or nuclide)

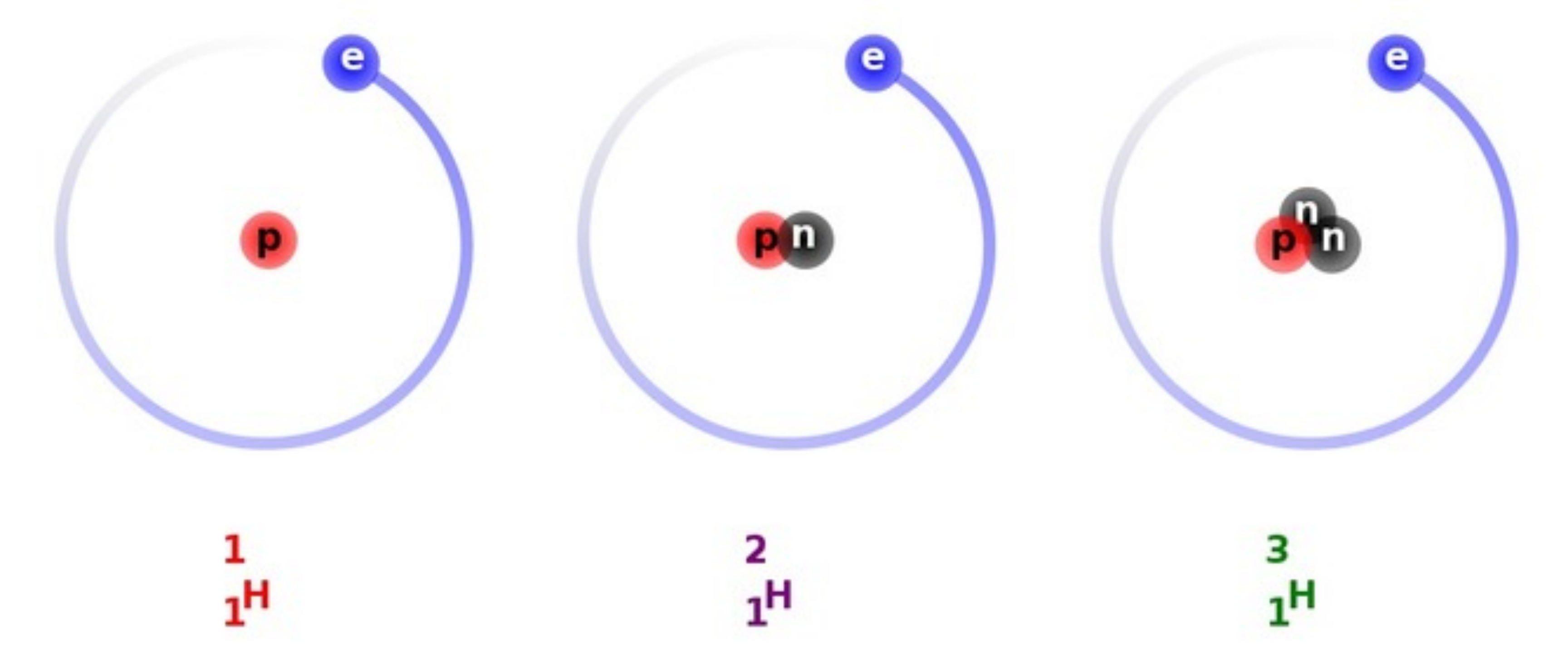
Z = atomic number (or number of protons)

Isotopes

- **Isotopes** of an element have the same number of protons (the number of protons in a nucleus defines which element it is) but a different number of neutrons
- Isotopes of a given element have the same chemical properties, the same atomic number, but a different mass number

Isotopes of Hydrogen

Standard Hydrogen	Deuterium	Tritium
${}^{1}_{1}H$	${}^{2}_{1}H$	${}^{3}_{1}H$
1 proton	1 proton + 1 neutron	1 proton + 2 neutrons



Limitations of the Simple Nuclear Model

Although Rutherford's nuclear model of the atom explained the data from the gold-foil (or Geiger-Marsden) experiment, there were limitations with this model. Firstly, for the electrons to remain in a circular orbit, they must be constantly accelerating towards the centre of the circle, and Maxwell had already shown that an accelerating electron emits electromagnetic radiation (EMR). So the electron should be constantly losing energy as EMR and so spiral in to the centre. Rutherford's model also failed to explain the existence of isotopes and observations of atomic line spectra.

Atomic Energy Levels

- In 1913, Bohr proposed that electrons could only occupy certain stable orbits and that an electron could jump to a higher orbit if it received precisely the right amount of energy
- Bohr hypothesised that the angular moment of orbiting electrons is quantised

The Electromagnetic Spectrum

- The electromagnetic (EM) spectrum is a continuous spectrum of different wavelengths
- Since the velocity of a wave is given by $v = f\,\lambda$, the relationship between frequency and wavelength for EM radiation is:

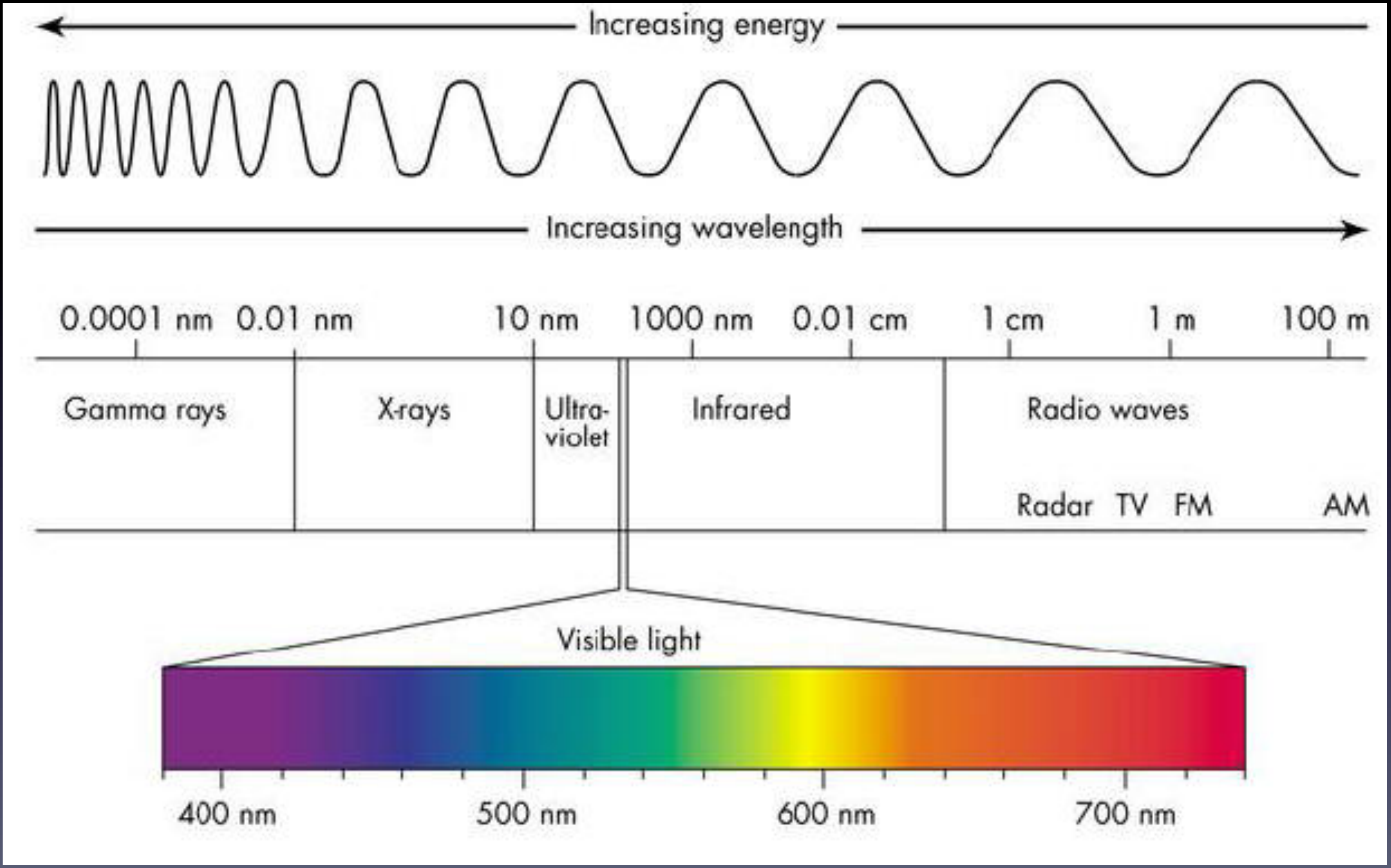
$$c = f\,\lambda$$

c = the speed of light = 2.9979×10^{8} m.s^{-1}

f = frequency, Hz

λ= wavelength, m

- High energy EM radiation has a high frequency (and short wavelength)
- Low energy EM radiation has a low frequency (and long wavelength)



Increasing energy
Increasing wavelength
0.0001 nm
0.01 nm
10 nm
1000 nm
0.01 cm
1 cm
1 m
100 m
Gamma rays
X-rays
Ultra-
violet
Infrared
Radio waves
Radar
TV
FM
AM
Visible light
400 nm
500 nm
600 nm
700 nm

Wave-Particle Duality

- Newton: particle theory to explain reflection and refraction
- Huygens, 1678: a wave theory could also explain reflection and refraction
- Young, 1801: double-slit experiment showing interference → wave nature of light
- Maxwell, 1873: light a high frequency from of EM wave
- Hertz, 1887: confirmed Maxwell by producing and detecting EM waves
- Hertz: also discovered photoelectric effect, E_K of emitted electron dependent on frequency not intensity
- Plank, 1901: proposed light energy is quantised in photons
- ?Einstein, 1905: equation for the quantisation of light energy $E = hf$.

Quantisation of Light

- In 1901, the German scientist Max Plank proposed that the *energy* of electromagnetic radiation comes in discrete **quanta** (amounts), and that the energy of each photon is proportional to its frequency

$$E = hf$$

E = energy of the photon, J

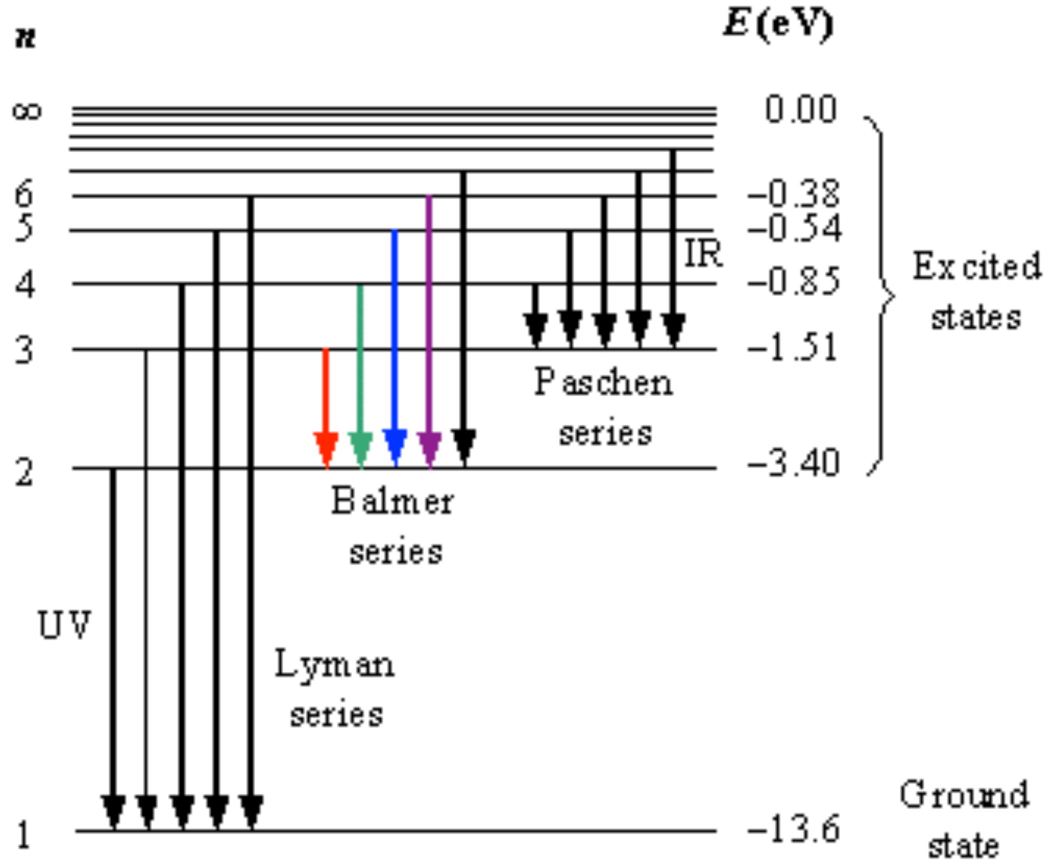
f = frequency of the photon, Hz

h = Plank's constant = 6.626 x 10--$^{-34-}$ J.s

Atomic Line Spectra

- Electrons in gaseous elements can move between higher and lower energy levels of an atom
- Electrons can jump to a higher energy (‘excited’) state by gaining energy from a strong electric field (e.g. electrons in the atoms of fluorescent or neon lights) or by absorbing a photon with energy corresponding exactly to the energy gap between the lower and higher energy states ($E = hf$)
- When electrons fall back to a lower energy state, they emit a photon with energy corresponding exactly to the energy gap between the higher and lower energy states

Atomic Line Spectra



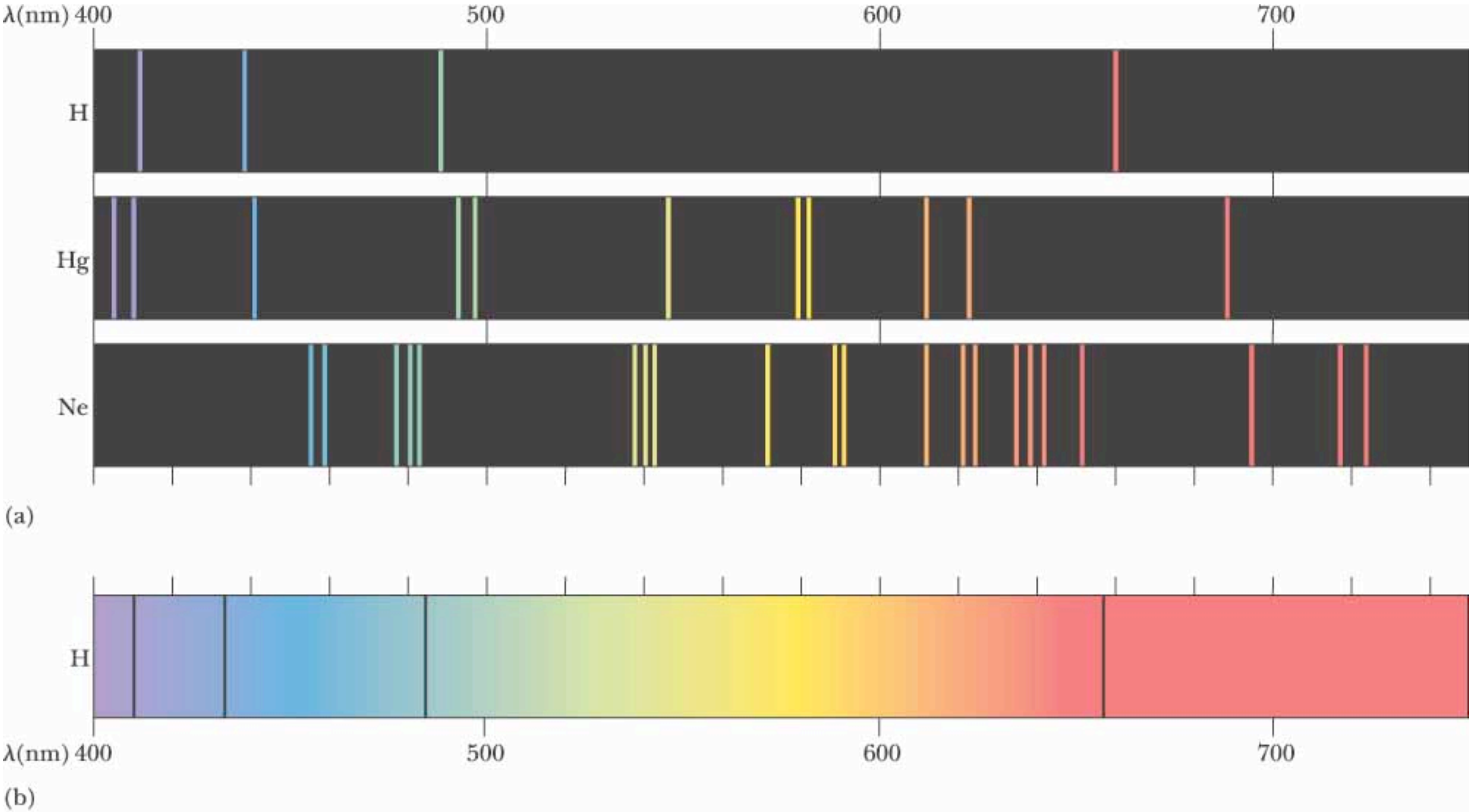
Energy levels of the hydrogen atom with some of the transitions between them that give rise to the spectral lines indicated.

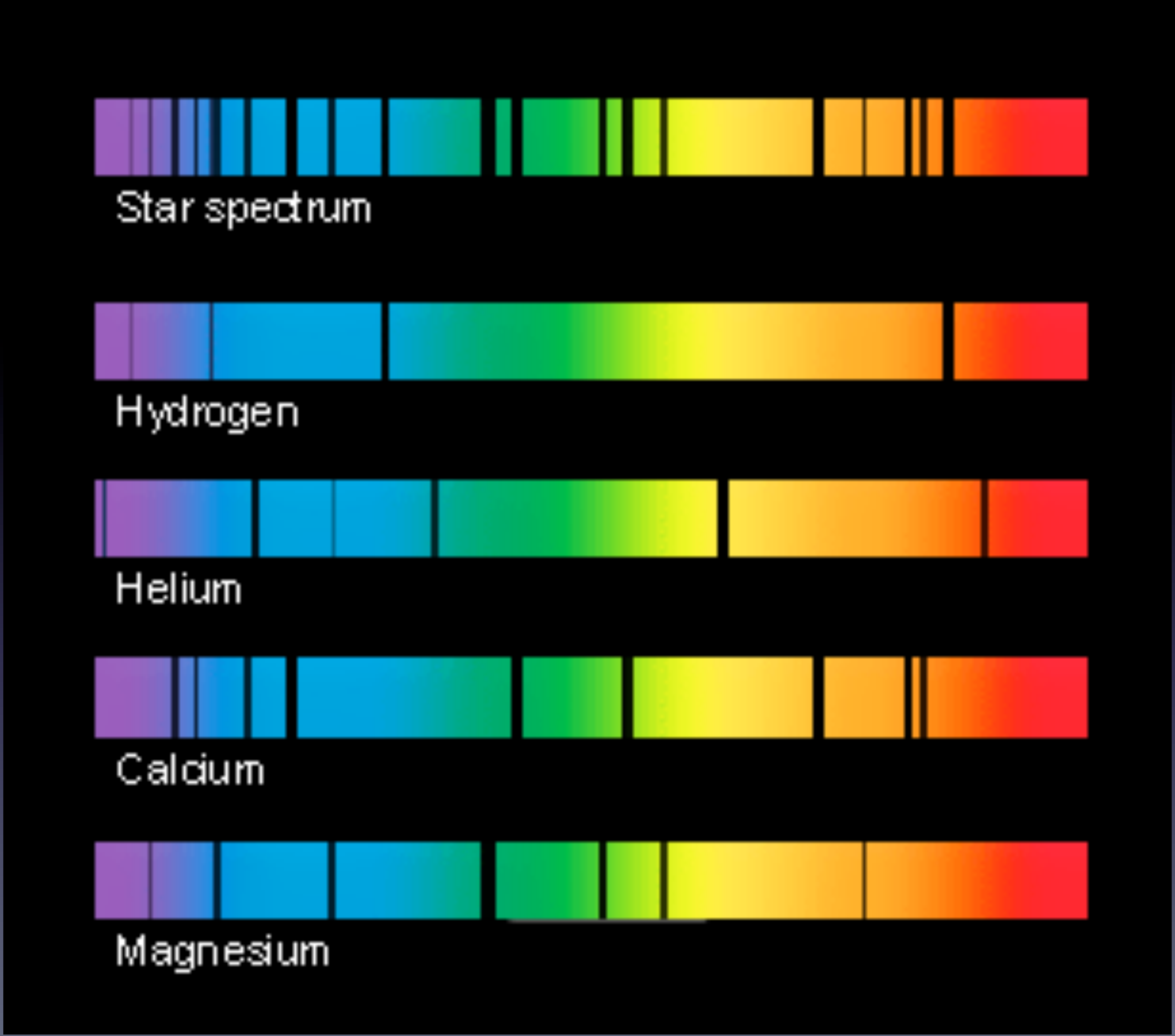
Atomic Line Spectra

- Each atomic element has a specific range of electron energy levels
- When the electrons of a given element are excited, the atom will emit a specific spectrum of wavelengths corresponding to the energy gaps between the electron energy levels
- This is called the **emission spectrum** of that element

Atomic Line Spectra

- Likewise, if an element is exposed to a broad, continuous spectrum of EMR, it will absorb photons with energies corresponding to gaps between the electron energy levels of that element
- This is called the **absorption spectrum** of that element. It is the same as the emission spectrum for that element


λ(nm) 400
500
600
700
H
Hg
Ne
(a)
H
λ(nm) 400
500
600
700
(b)


Star spectrum
Hydrogen
Helium
Calcium
Magnesium

Luminescence

- **Fluorescence** is when an electron absorbs a high energy photon and jumps a number of energy levels. The electron then ‘cascades’ back down through the energy levels, emitting a photon with each transition between energy states. If the emitted photons are in the visible range, we see them as fluorescence, e.g. molecules in certain jellyfish .
- **Phosphorescence** occurs when a material has a ‘metastable’ energy level that can ‘trap’ cascading electrons for up to a number of hours. Such substances can ‘glow’ (emit light) for a number of hours after exposure to high energy EMR, e.g. phosphorescent paint on watch dials.

7.2 Radioactive Decay

Radioactivity

- In 1896, the Austrian physicist Henri Becquerel discovered radioactivity when studying uranium salts
- Shortly after this, Pierre and Marie Curie isolated two other radioactive elements: polonium and radium
- **Radioactive elements** are elements that spontaneously emit high-speed particles or high energy waves from its nucleus
- Radioactivity is a property of the nucleus of the atom and is independent of the chemical state of the atom

Types of Radiation

- An **α-particle** is a helium nucleus (2 protons + 2 neutrons) released at high speed from a nucleus
- An α-particle is represented by ${}^{4}_{2}He$

For example, decay of a uranium nucleus to form a thorium nucleus and an α-particle

$${}^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He$$

Types of Radiation

- A **β-particle** is a high energy electron that is released from a nucleus.
- A β-particle is formed when a neutron splits into a proton (positive) and an electron (negative)
- A b-particle is represented by $^{0}_{-1}\beta$

For example, decay of a thorium nucleus to form a protactinium nucleus and a β-particle

$$^{234}_{90}Th \rightarrow {}^{234}_{91}Pa + {}^{0}_{-1}\beta$$

- In this reaction, a neutron in the thorium nucleus splits into a proton and an electron (β-particle)

Types of Radiation

- A **positron** (${}^{0}_{+1}\beta$) is similar to a β-particle in that it has the same mass and the same magnitude of charge, but it is positive instead of negative
- Positrons are the **antiparticle** of electrons
- When a positron is released from a nucleus it is also a type of β-radiation

Types of Radiation

- **γ-rays** are very high energy (short wavelength) electromagnetic waves that are released from a nuclear reaction

For example, iodine-131 going from an excited state to a lower energy state

$$^{131}_{53}I^* \rightarrow {}^{131}_{53}I + \gamma$$

Neutrinos and Antineutrinos

- **Neutrinos** and **antineutrinos** are tiny particles that are very difficult to detect
- This is because they have zero rest mass, and so have the ability to travel at the speed of light
- Neutrinos were first detected in 1956
- An antineutrino is emitted during β^- decay, and a neutrino is emitted during β^+ decay

Ionising Radiation

- α-particles, β-particles, γ-rays, cosmic rays and X-rays are all forms of **ionising radiation**; they all have the ability to interact with matter to form ions
- Ionising radiation can damage DNA, which can cause uncontrolled cell division and cancer
- Ionising radiation can also affect metabolic pathways and the functioning of cells

Effects of Radiation

Energy Absorbed from Radiation	Effect
0.25 $J.kg^{-1}$	May cause leukaemia
0.8-1 $J.kg^{-1}$	Radiation sickness, chance of recovery after 6 months
5 $J.kg^{-1}$	Usually fatal

Properties of Radiation

α-Radiation

- α-particles are strong ionisers, and they have the ability to remove electrons from other atoms (creating ions)
- Because of this, a-particles can be very damaging to living tissue
- However, α-particles are not very penetrating, and most can be stopped by a layer of thin cardboard

Properties of Radiation

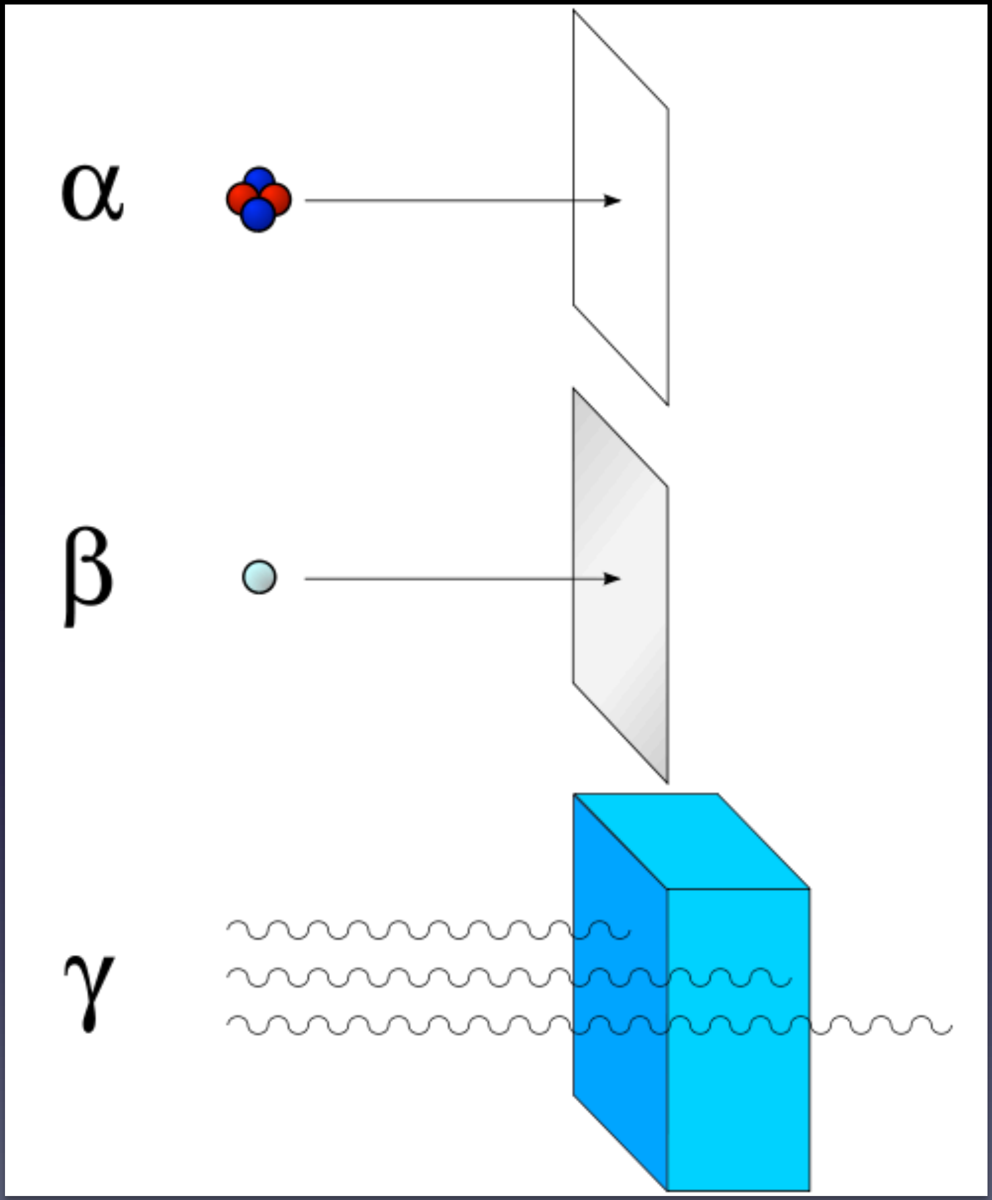
β-Radiation

- β-particles form ions by colliding with the electrons of an atom and removing them from the atom
- β-particles are less ionising than a-particles, but they are lighter, faster and more penetrating
- Most β-particles can be stopped by a thin layer of metal foil, e.g. 18 mm of Aluminium

Properties of Radiation

γ-Radiation

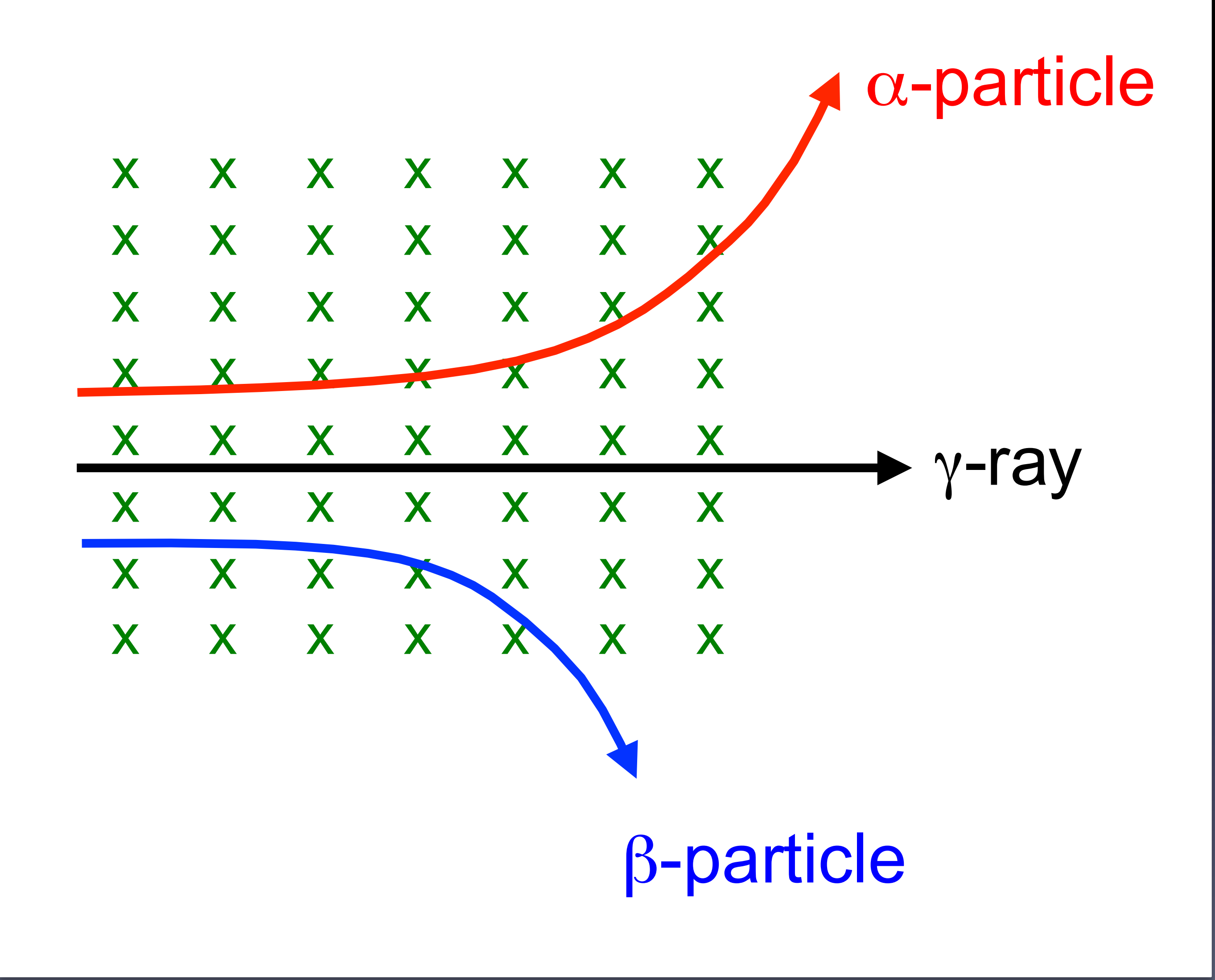
- γ-radiation is the least ionising but most penetrating type of radiation
- γ-radiation travels at the speed of light and has a very high energy
- Most γ-rays can be stopped by a thick block of lead or concrete


α
β
γ

Type of Radiation	Nature of Radiation	Represented by	Ionising ability	Penetrating ability	Stopped by
α-particle	Helium nucleus	${}^{4}_{2}He$	High	Low	Cardboard
β-particle	Electron emitted from a nucleus	${}^{0}_{-1}\beta$	Not as high as α-particles	Higher than α-particles	Sheet of metal
γ-ray	Electromagnetic radiation		Least ionising	Very high	Block of lead

Radioactivity in a Magnetic Field

- The 3 types of radiation behave differently in a magnetic field
- α-particles carry a positive charge, β-particles carry a negative charge, so they are deflected in opposite directions when travelling through a magnetic field (right-hand-slap rule)
- γ-rays are not charged and so are un-deflected in a magnetic field

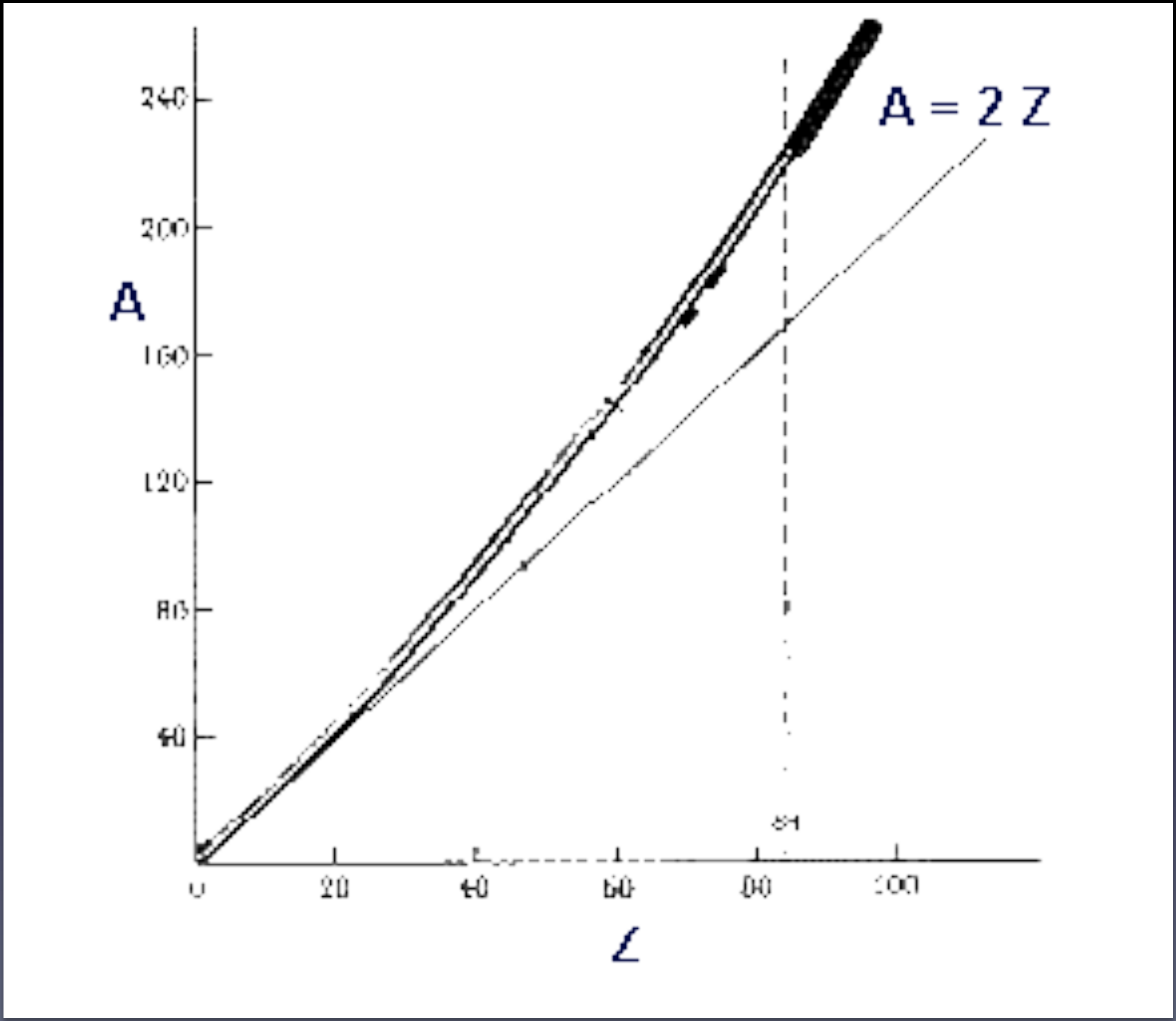


α-particle
γ-ray
β-particle

Stability of Nuclei

- Some nuclei are more stable than others; less stable nuclei are more likely to undergo radioactive decay
- The stability of a nucleus depends on the size of the nucleus and the mixture of protons and neutrons it contains
  - Nuclei with an atomic number greater than 95 are all radioactive (i.e. unstable)
  - Elements with an atomic number greater than 10 (Neon), have increasingly more neutrons than protons
  - Even-even combinations of protons and neutrons form the most stable nuclei; 80% of the Earth's crust comprises of 6 stable even-even nuclei: $^{16}_{8}O$, $^{24}_{12}Mg$, $^{28}_{14}Si$, $^{40}_{20}Ca$, $^{48}_{22}Ti$, $^{56}_{26}Fe$
  - Even-odd and odd-even combinations are the next most stable
  - Only 6 nuclei have odd-odd combinations and are the least stable, with the exception of the stable nucleus $^{14}_{7}N$

Graph of Atomic Mass vs. Atomic Number



Radioactive Half-Life

- Every radioactive substance has a specific **half life** (τ or $t_{1/2}$)
- The half life is the amount of time it takes for half of a sample of a radioactive substance to decay (or the radioactivity to decrease by half)
- Radioactive decay is independent of the temperature or chemical context of the atom
- The timing of when any given radioactive atom will decay is random and is a matter of statistical probability

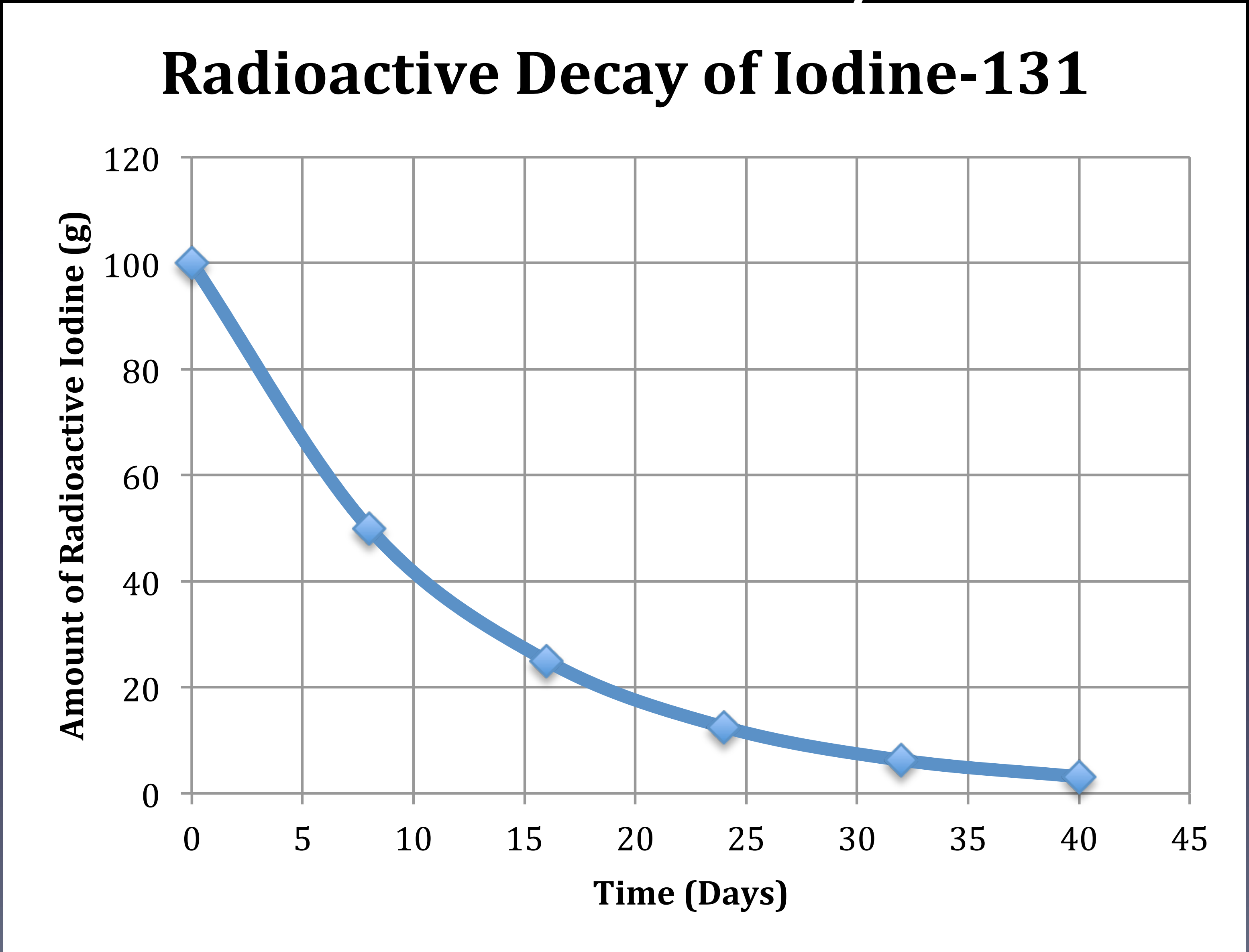
Radioactive Half-Life

- Consider a 100 g sample of iodine-131, which has a half life of 8 days
- After 8 days (1 half life), approximately 50 g of the sample will have decayed, and 50 g will still be radioactive
- After another 8 days (2 half lives), approximately 75 g of the sample will have decayed, and 25 g will still be radioactive
- After another 8 days (3 half lives), approximately 87.5 g of the sample will have decayed, and 12.5 g will still be radioactive

Radioactive-Decay Curves

- A graph of radioactivity vs time is called a **radioactivity-decay curve**
- Radioactivity-decay curves show an exponential decrease in the number of radioactive particles remaining

Radioactive-Decay Curves



7.3 Nuclear Reactions, Fission and Fusion

Artificial Transmutations

- Although the process of radioactive decay is random, we can artificially stimulate a nuclear reaction by bombarding an atom with, for example, an alpha particle, a proton, a neutron or a small nucleus
- The target nucleus generally ‘captures’ the particle bombarding it, making the target nucleus unstable and so stimulating a nuclear reaction
- The first artificial transmutation was carried out by Rutherford in 1919 by bombarding nitrogen with alpha-particles

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

Atomic Mass Units

- An **atomic mass unit** or **unified mass unit** (u) is defined as the rest mass of a carbon-12 atom (which contains 6 protons and 6 neutrons) divided by 12

$$1\ \text{u} = 1.66 \times 10^{-27}\ \text{kg}$$

- The rest mass of a proton is 1.007276 u = 1.6726×10^{-27} kg
- The rest mass of a neutron is 1.008665 u = 1.6749×10^{-27} kg
- The rest mass of an electron is 0.000549 u

Electron Volts

- The **electron volt** (eV) is a measure of energy
- It is often used in atomic physics where the energies concerned are very small
- One electron volt is the amount of energy required to move an electron through a potential difference of one volt

$$1\ \text{eV} = 1.6 \times 10^{-19}\ \text{J}$$

Mass-Energy Equivalence

- Einstein's famous equation describes the relationship between energy and mass

$$E = mc^2$$

m = mass, kg

E = energy stored in that mass, J

c = speed of light = 2.9979×10^8 $m.s^{-1}$

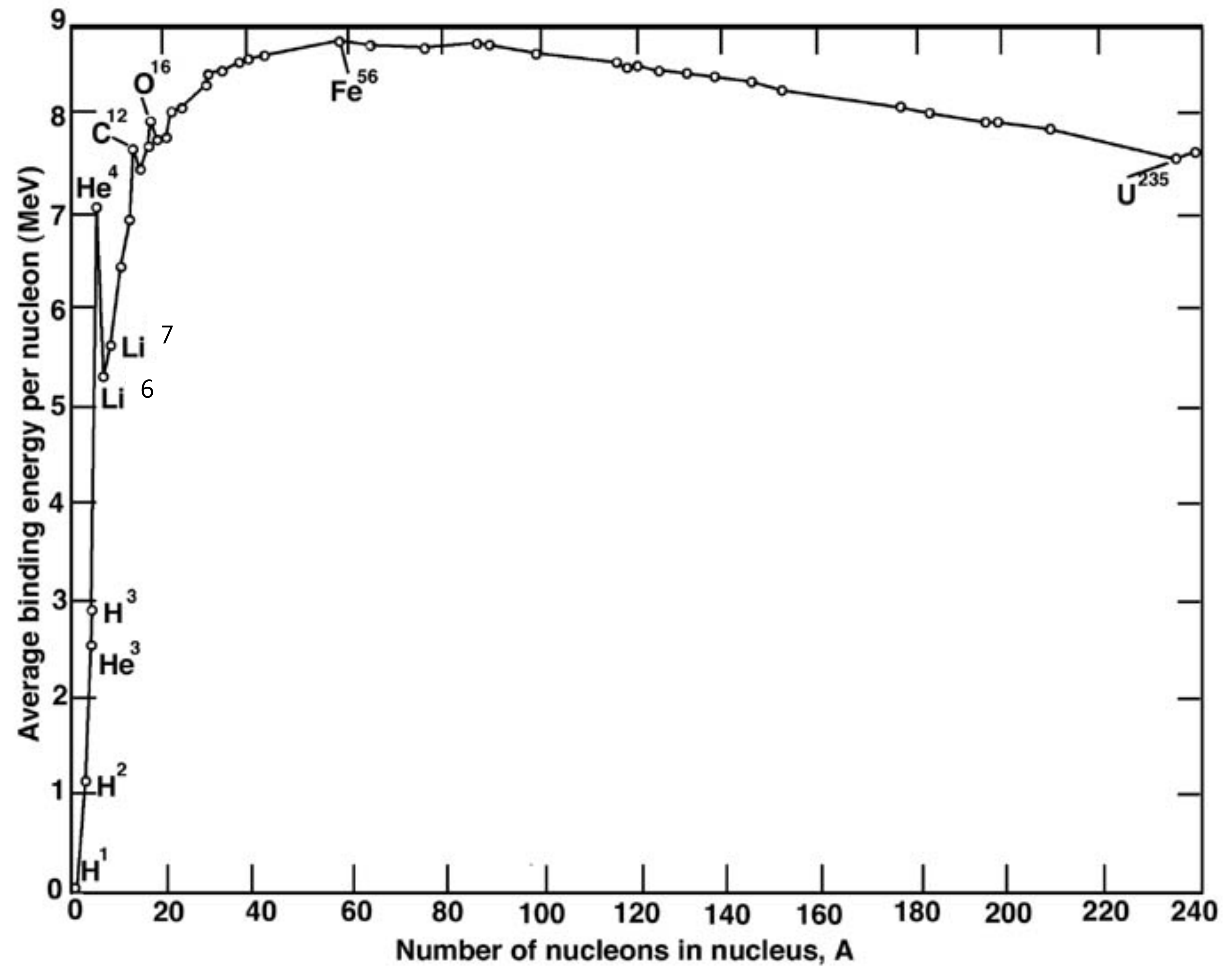
- Essentially, we can consider matter to be a form of stored energy
- Overall, **mass-energy** is conserved, e.g. in a nuclear reaction some mass may be converted to electromagnetic energy, but the total mass and energy in the reaction is conserved

Binding Energy

- A very strong force is required to overcome the electrostatic repulsion between the protons in a nucleus; gravity is not strong enough
- It is found that the mass of a nucleus is slightly less than the total mass of its constituent nucleons
- The 'missing mass' (or **mass deficit**) is converted to energy according to the relationship $E = mc^2$
- The energy that is released when nucleons are assembled into a nucleus is called **nuclear binding energy**
- The binding energy can also be defined as the amount of energy required to split a nucleus into its constituent nucleons

Binding Energy per Nucleon

- The binding energy per nucleon is an important measure of the stability of a nucleus
- More stable nuclei have more binding energy per nucleon (i.e. more mass-energy has been released in the formation of the nucleus)
- The most stable nucleus is ^{56}Fe
- Nuclei with mass numbers greater than 56 can split by nuclear fission to form smaller, more stable nuclei
- Nuclei with mass numbers less than 56 can join together by nuclear fusion to form larger more stable nuclei



Average binding energy per nucleon (MeV)
Number of nucleons in nucleus, A
0
1
2
3
4
5
6
7
8
9
0
20
40
60
80
100
120
140
160
180
200
220
240
H^1
H^2
He^3
H^3
Li 6
Li 7
He^4
C^{12}
O^{16}
Fe^{56}
U^{235}

Nuclear Reactions

- The following rules apply in nuclear reactions:

1. Charge is conserved
2. The total number of nucleons is conserved
3. Mass-energy is conserved
4. Linear momentum is conserved

Nuclear Fission

- **Nuclear fission** is when a heavier nucleus splits into two smaller nuclei and releases energy and, usually, some particles

For example,

$${}^{235}_{92}U + {}^{1}_{0}n \rightarrow {}^{141}_{56}Ba + {}^{92}_{36}Kr + {}^{1}_{0}n$$

$$[{}^{1}_{0}n \rightarrow {}^{1}_{1}p + {}^{0}_{-1}\beta]$$

Critical Mass

- The **critical mass** of a fissionable element is the mass required for a chain reaction to occur
- This happens when neutrons released from a fission reaction are captured by other heavy nuclei, making them unstable, and causing another fission reaction with further release of neutrons

Nuclear Fusion

- **Nuclear fusion** is when two smaller nuclei join (fuse) to form a larger, more stable nucleus

$$^{2}_{1}H + ^{3}_{1}H \rightarrow ^{4}_{2}He + ^{1}_{0}n$$

- The advantage of fusion is that the reactants are readily obtained and it does not produce radioactive products
- The problem with fusion is that very high temperatures and pressures are required for the reaction to occur
- Nuclear fusion is the reaction that powers stars
- The fusion reaction in our Sun is predominantly the fusion of Hydrogen into Helium

Energy Release in Fission and Fusion

- To calculate the energy released in a nuclear reaction:
  - Add up the total mass on the left-hand side of the reaction and add up the total mass on the right-hand side of the reaction
  - Determine the mass deficit by subtraction
  - Calculate the energy released using $E = mc^2$
- Note: if the mass of the products is greater (i.e. there has been a mass increase) the reaction is not energetically favourable and will not proceed without the *input* of energy

Energy Release in Fission and Fusion

- The energy released in a nuclear reaction is in the form of kinetic energy of the particles released and gamma rays
- The energy released by the oxidation of one carbon atom (e.g. in a coal-fuelled power plant) is about 4 eV
- The energy released by the fission of one uranium atom is about 150 MeV
- The energy released by the fusion of a hydrogen-2 nucleus and a hydrogen-3 nucleus is about 18 MeV