The nucleus

Rutherford's nuclear atom (1902-1920)

- Ernest Rutherford was interested in the distribution of electrons in atoms.
- Two of his students, Geiger and Marsden, used radium as a source of alpha particles which they 'fired' at a thin piece of gold foil.
- A mobile fluorescent screen was used to follow the paths of the alpha particles.

Rutherford's nuclear atom (1902-1920)

- Rutherford expected that the alpha particles, which are positively charged, would pass straight through the gold foil or be deflected slightly.
- However, he was astounded by the results of these experiments. Most of the alpha particles passed straight through the foil but a few appeared to **rebound from** the foil.

Rutherford's nuclear atom (1902-1920)

- After further careful measurements Rutherford proposed a model for the atom in which
  - a tiny dense central nucleus contains all the positive charge and most of the mass
  - a much larger outer region is occupied by orbiting electrons and contains all the negative charge but very little of the mass.

Rutherford's nuclear atom (1902-1920)

- Thus if an alpha particle comes close to the minute positively charged nucleus it is strongly repelled and deflected through a large angle.

Rutherford's nuclear atom (1902-1920)

- Rutherford also suggested that the nucleus contained positively charged particles called **protons.**
- He also predicted the existence of the **neutron which** was not discovered, until 1932, by Chadwick.

Moseley and the nucleus (1913)

- Henry Moseley was able to demonstrate the relationship between atomic structure and chemical properties.
- Using X-ray scattering techniques he showed that the amount of positive charge on the nucleus is a fundamental property of each element.

Moseley and the nucleus (1913)

- He assigned an atomic number for each element which corresponded with the numbered position (Z) to each element in the Periodic Table. This helped to explain some of the anomalies in Mendeleev's table which was based on atomic mass.

Atomic Number (Z)

- Is the number of **protons** in the nucleus of an atom.
- It is equal to the number of electrons in the neutral atom.
- All atoms of the same element have the same atomic number.

Frederick Soddy (1877-1956)

- Proposed the existence of isotopes
  - In the early 1900s, scientists discovered dozens of 'new' radioactive elements which could not be fitted into the ten or so gaps in the Periodic Table.
  - However, it was found that some of these elements had identical chemical properties although their radioactive properties, such as half-life and type of emitted radiation, differed.

Frederick Soddy (1877-1956)

- In 1913 Soddy explained these observations by introducing the idea of isotopes (from the Greek, meaning 'same place') as elements with the same chemical properties but containing atoms which differed in mass, physical properties and radioactive behaviour.
- The relative atomic mass of such an element would therefore be an average according to the number and type of each kind of atom present.

Frederick Soddy (1877-1956)

- We know today that **isotopes are different atoms** of the **same element.**
- They are atoms of the same element because they have the same atomic number (same number of protons).
- However, they contain different numbers of neutrons and hence have different mass numbers (number of protons plus neutrons).

Frederick Soddy (1877-1956)

- Soddy predicted that two isotopes of lead, lead-206 and lead-208 would be produced by the radioactive decay of uranium-238 and thorium-232 respectively.
- These two isotopes of lead are stable and therefore not radioactive.
- Careful measurements of their relative atomic masses vindicated Soddy's views in 1914.

Frederick Soddy (1877-1956)

- The existence of isotopes was later shown to be widespread.
- Only a few elements consist of one type of atom (or nuclide) e.g. Be-9, F-19 and Al-27
- The existence of isotopes was confirmed in 1919 when Aston invented the mass spectrometer.

Frederick Soddy (1877-1956)

- This instrument was used to separate isotopes according to the behaviour of their ions in a magnetic field.
- He was able to determine the relative masses and the percentage abundances of naturally occurring isotopes.
- An explanation for the existence of isotopes did not happen until 1932 when Chadwick discovered the **neutron.**

Chadwick's discovery

- Rutherford suggested that hydrogen, the smallest atom, has **one proton** in its nucleus balanced by one orbiting electron.
- An atom of helium should therefore have **two protons** in the nucleus balanced by two orbiting electrons but an atom of helium is **four times** as heavy as an atom of hydrogen, not twice as heavy.
- Chadwick's investigations demonstrated the existence of uncharged particles in the nuclei of atoms.

Chadwick's discovery

- He determined that the mass of a neutron is similar to the mass of a proton.
- Thus a helium atom could contain two protons and two neutrons in its nucleus, surrounded by two electrons.

Chadwick's discovery

- The existence of isotopes could now be explained - isotopes of the same element contain the **same number of protons in the nucleus** but the **number of neutrons may vary.**
- The neutron subsequently proved to be a very useful tool with which to investigate the atom.
- As an uncharged particle it can easily penetrate the nucleus.

Mass Number (N)

- The number of nucleons (protons and neutrons) in the nucleus of an atom.
- Different isotopes of the same element have the same atomic numbers (Z) but different mass numbers.

Identifying an Individual Isotope

- An individual isotope can be identified by its mass number:
  - chlorine- 35, chlorine-37
  - uranium-235, uranium-238
- The symbol for a particular isotope, its atomic number and its mass number is often represented as follows:

Mass number $\rightarrow$ A

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

Atomic number $\rightarrow$ Z

Mass Spectrometry

- Today, mass spectrometry is used to determine:
  - relative masses and abundances of isotopes
  - relative masses and structures of complex molecular substances

Operation of a Mass Spectrometer

- vaporisation
  - Sample must enter as a gas.
- ionisation
  - Atoms of the gaseous sample are bombarded with electrons; mainly singly charged positive Ions are formed.

Operation of a Mass Spectrometer

- acceleration
  - The ions are accelerated by a strong electric field.
- deflection
  - The ions are deflected in circular paths, by a powerful magnetic field, according to their charge and their mass - the greater the deflection, the lower the mass (for ions of the same charge).

Operation of a Mass Spectrometer

- detection
  - The intensities of different ion beams are detected electronically.
- collection
  - The collector records the data as a mass spectrum which is a graph of percentage relative abundance against relative isotopic mass.

Operation of a Mass Spectrometer

- calibration
  - The instrument is calibrated against a standard isotope (carbon-12) which is given a value of 12 units exactly.

Mass Spectrometry
Some useful definitions

- Relative abundance
  - The proportion of each isotope in a sample of an element
- Relative isotopic mass (RIM)
  - The mass of an isotope relative to the mass of the carbon-12 ("C) isotope with a mass of 12 units exactly.

Mass Spectrometry
Some useful definitions

- Relative atomic mass (A, or RAM)
  - the average of the relative isotopic masses of an element weighted according to their relative abundances on a scale where the carbon-12 6 isotope has a mass of 12 units exactly

Mass Spectrometry
Some useful definitions

- Relative molecular mass (M, or RMM)
  - Sum of the relative atomic masses of the atoms that make up a molecule.
- The information obtained from a mass spectrum enables the calculation of relative atomic mass:

$$\textbf{Ar (RAM)} = \frac{\%_1 \text{RIM1} + \%_2 \text{RIM2} + \ldots\ldots}{100}$$