2.1 The nuclear atom

Understandings:

- Atoms contain a positively charged dense nucleus composed of protons and neutrons (nucleons).

Guidance
Relative masses and charges of the sub-atomic particles should be known, actual values are given in section 4 of the IB data booklet. The mass of the electron can be considered negligible.

- Negatively charged electrons occupy the space outside the nucleus.
- The mass spectrometer is used to determine the relative atomic mass of an element from its isotopic composition.

Guidance
The operation of the mass spectrometer is not required.

Applications and skills:

- Use of the nuclear symbol notation $^{A}_{Z}X$ to deduce the number of protons, neutrons, and electrons in atoms and ions.
- Calculations involving non-integer relative atomic masses and abundance of isotopes from given data, including mass spectra.

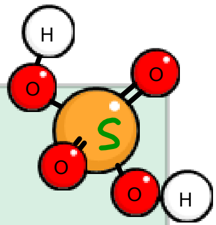
Guidance
Specific examples of isotopes need not be learned.

In the year 1803 scientists came up with the atomic theory

"Atom " comes from the greek: a = without tomos = division

John Dalton proposed...

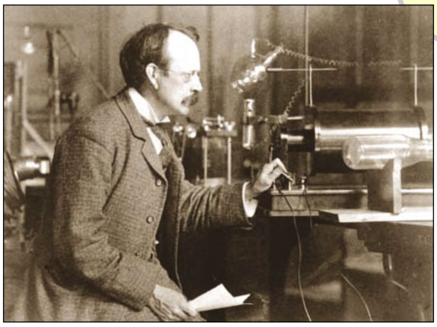
JOHN DALTON, D.C.L., F.R.S.

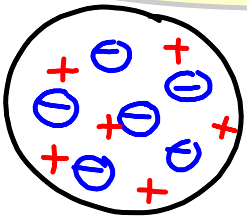


- All matter is made up of indivisible particles - atoms
- All atoms of a certain element are identical
- Atoms of different elements have a different weight and properties
- The atoms of different elements can combine with the atoms of other elements to form compounds
- Atoms cannot be created, divided or destroyed in a chemical process.

Subatomic particles

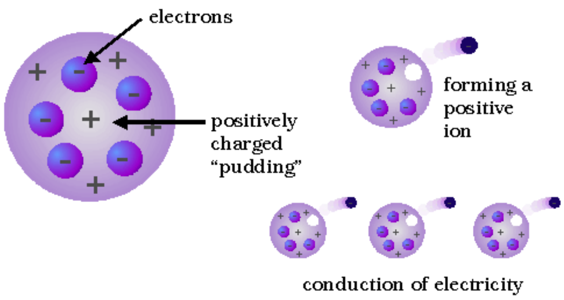
JJ Thompson discovered the <u>electron</u> in 1897
(1856 – 1940)

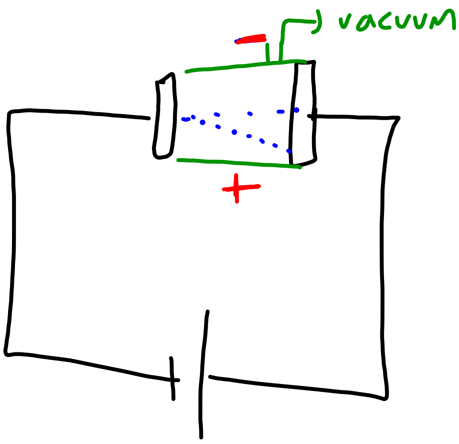


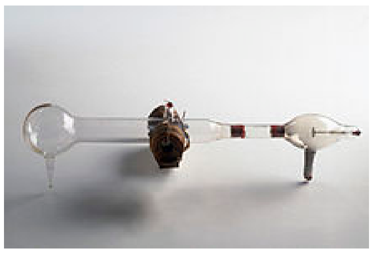


Plum pudding

He stated that there were negatively charged particles called electrons floating around in a positively charged cloud - plum pudding model . This was discovered by firing cathode rays (beams of electrons) through an electromagnetic field and seeing them move towards the positive side.

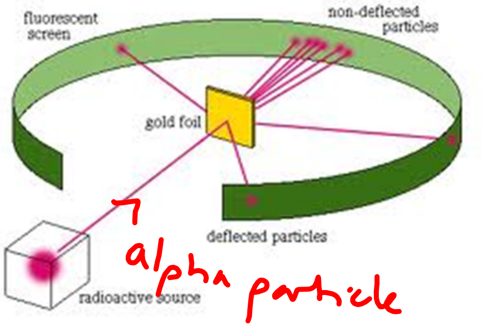




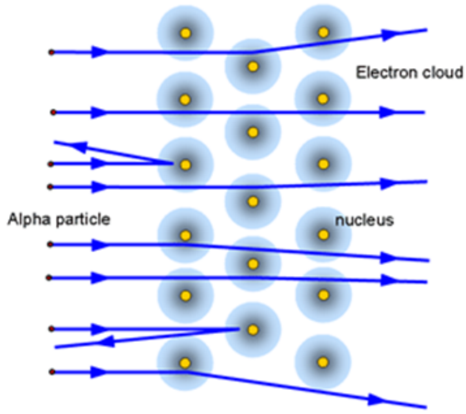


Rutherford

1871 – 1937

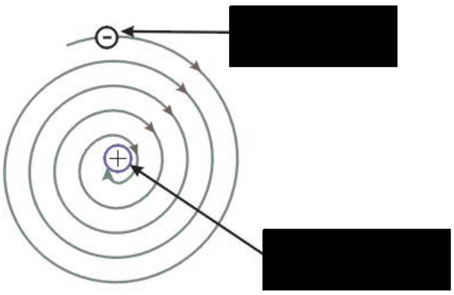


Observations:
1. The majority of the particles passed through the gold foil.
2. A small proportion of particles was strongly deflected.
3. One out of 10.000 particles bounched of the gold foil



Conclusions:
1. The majority of atoms is empty space
2. The atom's positive charge is found in a small dense nucleus
3. The electrons orbit around the nucleus

Problems: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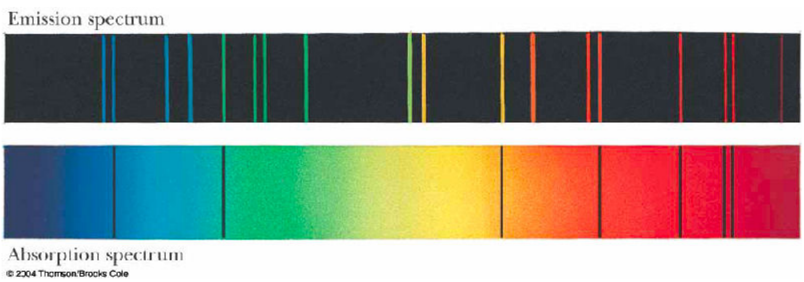
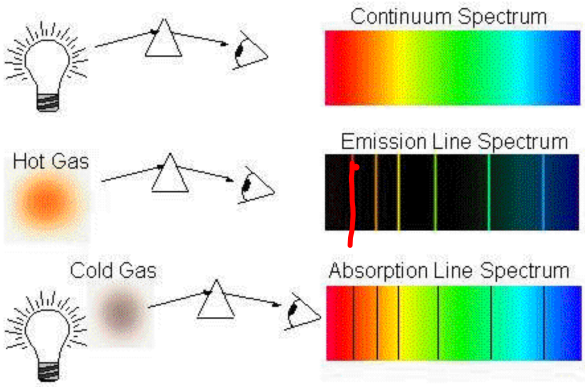
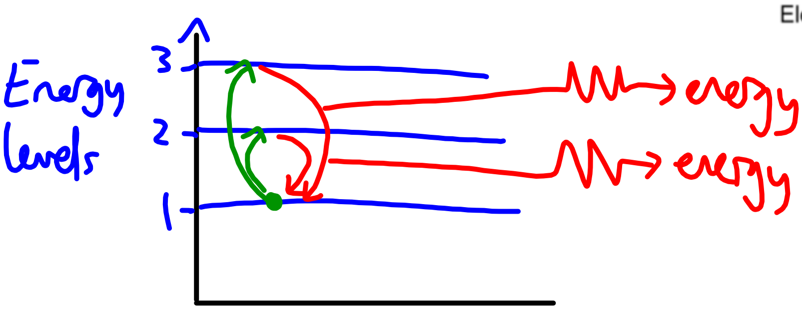


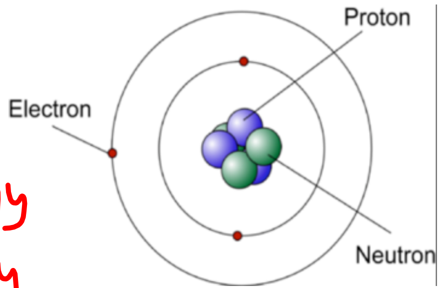
Neils Bohr

7 October 1885 – 18 November 1962



How did these observations lead him to the idea of set energy levels where the electrons exist?

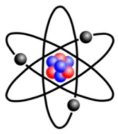




Conclusions:

- Electrons move in orbitals around the nucleus, without gaining or losing energy.
- Only orbitals with specific amounts of energy are permitted.
- An electron can only change its orbit by emission or absorption of specific amounts of energy.

What is in an atom?



Subatomic Particles

Particle	Symbol	Mass (amu)	Charge
Proton	p^+	1	1+
Neutron	n	1	0
Electron	e^-	$\frac{1}{1837}$	1-

amu = atomic mass unit = 1.66×10^{-27} kg

We use the atomic mass unit to avoid having to deal with very small numbers and it allows us to describe the mass of a proton and neutron as 1, and an electron as effectively 0.

Different elements have different numbers of protons, neutrons and electrons. We describe the number of subatomic particles in an atom using the nuclear symbol:

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

Chemical symbol - X

The atomic number - Z

Describes the number of protons in the atom.

Each element has a specific number of protons.

In a neutral atom, we will always have the same number of electrons as protons (so the charges cancel out).

The mass number - A

Describes the total number of protons and neutrons in the atom.

$${}^{12}_{6}\mathrm{C}$$

$${}^{238}_{92}\mathrm{U}$$

$${}^{48}_{22}\mathrm{Ti}$$

Most elements have isotopes which are atoms with the same number of protons but a different number of neutrons.

Why are each of the examples below hydrogen atoms and how many subatomic particles are in each of them?

${}^{1}_{1}\mathrm{H}$	${}^{2}_{1}\mathrm{H}$	${}^{3}_{1}\mathrm{H}$
(a) Hydrogen H – 1	(b) Deuterium H – 2	(c) Tritium H – 3

Isotopes have the same number or electrons so have the same chemical properties but can have different physical properties such as melting point and boiling point.

(*physical properties can be changed without changing the chemical composition of the element)

3 Use the Periodic Table to identify the sub-atomic particles present in the following species.

	Species	No. of protons	No. of neutrons	No. of electrons
(a)	^{7}Li			
(b)	^{1}H			
(c)	^{14}C			
(d)	$^{19}F^-$			
(e)	$^{56}Fe^{3+}$			

4 Isoelectronic species have the same number of electrons. Identify the following isoelectronic species by giving the correct symbol and charge. You will need a Periodic Table.

The first one has been done as an example.

	Species	No. of protons	No. of neutrons	No. of electrons
	$^{40}Ca^{2+}$	20	20	18
(a)		18	22	18
(b)		19	20	18
(c)		17	18	18

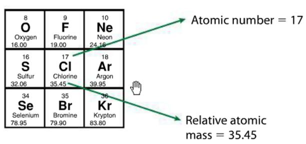
5 Which of the following species contain more electrons than neutrons?

A ${}^{2}_{1}\mathrm{H}$ **B** ${}^{11}_{5}\mathrm{B}$ **C** ${}^{16}_{8}\mathrm{O}^{2-}$ **D** ${}^{19}_{9}\mathrm{F}^{-}$

6 Which of the following gives the correct composition of the $^{71}Ga^{+}$ ion present in the mass spectrometer when gallium is analysed.

	Protons	Neutrons	Electrons
A	31	71	30
B	31	40	30
C	31	40	32
D	32	40	31

So why does the periodic table contain atomic mass numbers that are not integers?



Chlorine exists as two isotopes, ^{35}Cl and ^{37}Cl. The average relative mass of the isotopes is, however, not 36, but 35.45. This value is closer to 35 as there are more ^{35}Cl atoms in nature – it is the more abundant isotope. In a sample of 100 chlorine atoms, there are 77.5 atoms of ^{35}Cl and 22.5 atoms of the heavier isotope, ^{37}Cl.

How do we calculate the average atomic mass?

Try the average mass questions...

4. The relative atomic mass of naturally occurring copper is 63.55. Calculate the abundances of ^{63}Cu and ^{65}Cu in naturally occurring copper.

(Total 2 marks)

5. Silicon has three stable isotopes, ^{28}Si, ^{29}Si and ^{30}Si. The heaviest isotope, ^{30}Si, has a percentage abundance of 3.1 %. Calculate the percentage abundance of the lightest isotope to one decimal place. Average mass of Si = 28.09 amu

4. The relative atomic mass of naturally occurring copper is 63.55. Calculate the abundances of ^{63}Cu and ^{65}Cu in naturally occurring copper.

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5. Silicon has three stable isotopes, ^{28}Si, ^{29}Si and ^{30}Si. The heaviest isotope, ^{30}Si, has a percentage abundance of 3.1 %. Calculate the percentage abundance of the lightest isotope to one decimal place.

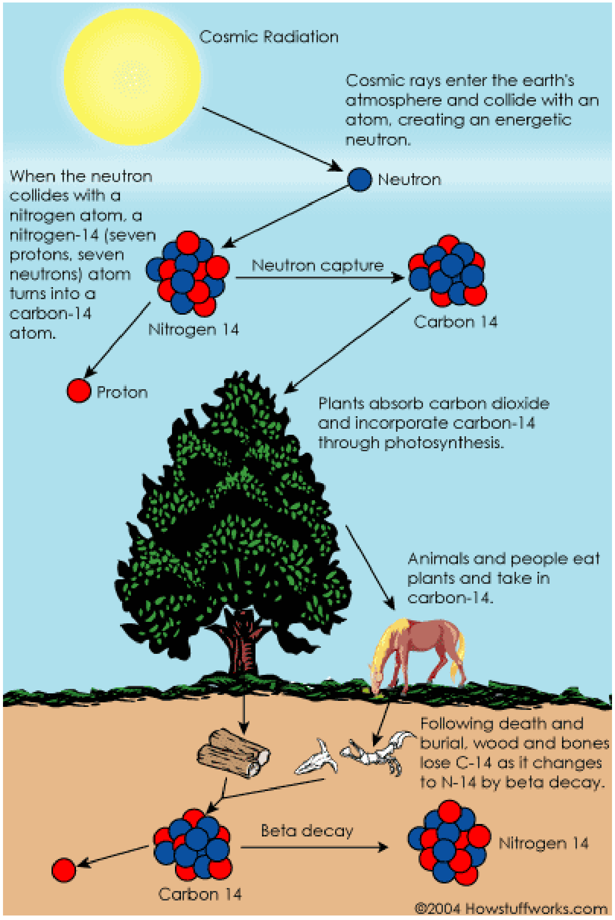
Some isotopes (that can be be made artifically as well as found naturally) are unstable and spontaneously release radiation such as an alpha particle.

We call these radioisotopes.

Common examples:

- Uranium-235 Nuclear energy and weapons
- Iodine-131 Treatmet of thyroid cancer
- Carbon-14 Carbon dating

Don't need to know these



Carbon-14 always decays at the same rate. We measure this decay using its half-life ($t_{1/2}$).

This is the time taken for the amount of a radioisotope to decrease to half of its intitial value.

For ^{14}C this is 5730 years.

We can measure age by looking at the ratio of ^{14}C to ^{12}C in a sample.

The older the sample, the lower the ratio.

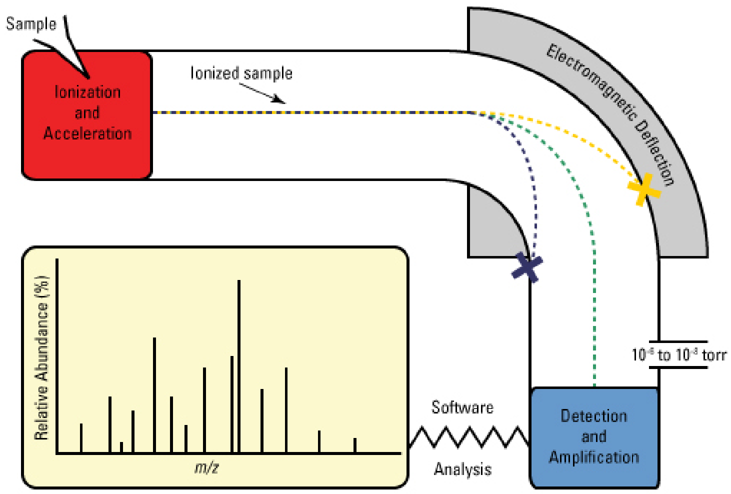
"Atomic mass " and " relative atomic mass "

We have see the atomic mass, A, as simply the mass of an atom. Normally we represent this in atomic mass units. (1 amu = 1.66 x 10^{27} kg).

The relative atomic mass, A_r, is the ratio between the average mass of an atom and the atomic mass unit .

- It is a ratio so contains no units.
- The average includes all its isotopes (as we have calculated).

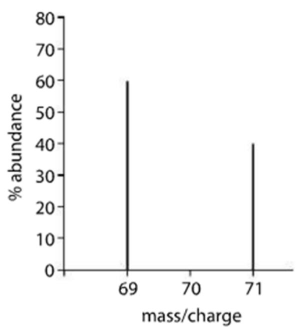
If we have an unknown sample of an element, we use a mass spectrometer to calculate the relative atomic mass. (You do not need to know how it works)

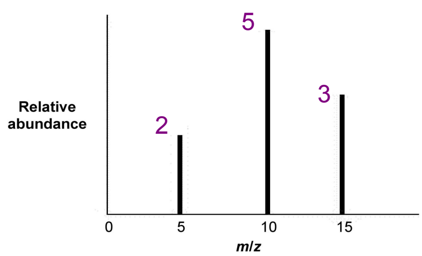


This process produces a mass spectrum .

This is a graph that shows the relative abundance (how much there is) of the different ions against the mass number (or mass-to-charge ratio, m/z).

By looking at the peaks with the largest masses we can identify the isotopes and the A_r.





Exercises

7 What is the same for an atom of phosphorus-26 and an atom of phosphorus-27?

A atomic number and mass number
B number of protons and electrons
C number of neutrons and electrons
D number of protons and neutrons

8 Use the Periodic Table to find the percentage abundance of neon-20, assuming that neon has only one other isotope, neon-22.

9 The relative abundances of the two isotopes of chlorine are shown in this table:

Isotope	Relative abundance
^{35}Cl	75%
^{37}Cl	25%

Use this information to deduce the mass spectrum of chlorine gas, Cl_2.

10 Magnesium has three stable isotopes – ^{24}Mg, ^{25}Mg, and ^{26}Mg. The lightest isotope has an abundance of 78.90%. Calculate the percentage abundance of the other isotopes.

11 The Geiger–Marsden experiment, supervised by Ernest Rutherford, gave important evidence for the structure of the atom. Positively charged alpha particles were fired at a piece of gold foil. Most of the particles passed through with only minor deflections but a small number rebounded from the foil.

How did this experiment change our knowledge of the atom?

A It provided evidence for the existence of discrete atomic energy levels.
B It provided evidence for a positively charged dense nucleus.
C It provided evidence that electrons move in unpredictable paths around the nucleus.
D It provided evidence for the existence of an uncharged particle in the nucleus.

average atomic mass of C = 12.01 amu

relative atomic mass of C = 12.01