Atomic Structure, Isotopes, Periodic Table and Electronic Structure of Atoms

KEY WORDS ➡ allotropes ➡Alpha particle scattering ➡Atomic (proton) number ➡ Atomic structure ➡ Electron ➡ Electron arrangement (examples) ➡ Electron shell rules ➡ Isotopes ➡ Mass (nucleon) number ➡ Neutron ➡ Neutron number ➡ Nuclide symbol notation ➡ Periodic Table (and electron structure) ➡ Periodic Table (its general structure) ➡ Proton ➡stable/unstable electron arrangements

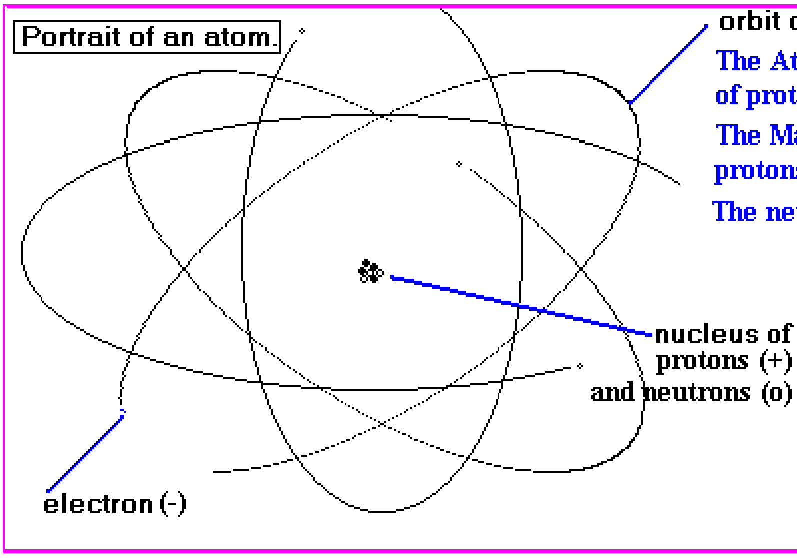
The Structure of Atoms - fundamental particles

Atoms are the smallest particles of matter whose properties we study in Chemistry. However from experiments done in the late 19th and early 20th century it was deduced that atoms were made up of three fundamental sub-atomic particles (listed below)

Particle	Relative mass	Electric charge	Comments
Proton	1	+1 (positive)	In the nucleus (a nucleon)
Neutron	1	0 (zero)	In the nucleus (a nucleon)
Electron	$^1/_{1850}$	-1 (negative)	Arranged in energy levels or shells around the nucleus (see later)

A Portrait of an Atom - an image of what you can't see!

- The diagram below gives some idea on the structure of an atom, it also includes some important definitions and notation used to describe atomic structure.
- Protons and neutrons are the ' nucleons' or 'sub-atomic' particles present in the minute positive nucleus and the negative electrons are held by the positive protons in 'orbits' called energy levels or shells.
- Some important evidence for this 'picture' is obtained from alpha particle scattering experiments (see Appendix 1).
- The atomic number (Z) is the number of protons in the nucleus and is also known as the proton number of the particular element.
- It is the proton number that determines the specific identity of a particular element and its electron structure.
- The mass number (A) is also known as the nucleon number, that is the number of particles in the nucleus of a particular isotope. These
- The neutron number (N) = mass number - proton/atomic number
- In a neutral atom the number of protons (+) equals the number of electrons (-), that is the number of positive charges is equal to the number of negative charges. If not, the atom has an overall charge and is then called an ion e.g. Na^+ (11p, 10e) or Cl^- (17p, 18e).
- Other more 'practical' diagrams than the one below are shown in Appendix 2.



The Atomic Number (Z) = the number of protons in the nucleus

The Mass number (A) = the sum of the protons and neutrons in the nucleus

The neutron number (N) = A - Z

Nuclide notation for isotopes

top left number = Mass N.(A)

bottom number = Atomic N.(Z)

$^{7}_{3}Li$ • this means an atom of lithium, atomic number of 3 (3 protons) and a mass number of 7 (3 protons plus 4 neutrons)

ISOTOPES

- Isotopes are atoms of the same element with different numbers of neutrons. This gives each isotope of the element a different mass or nucleon number but being the same element they have the same atomic or proton number.
- There are small physical differences between the isotopes eg the heavier isotope has a greater density or boiling point.
- However, because they have the same number of protons they have the same electronic structure and identical chemically. Examples are illustrated below.
- Do NOT assume the word isotope means it is radioactive, this depends on the stability of the nucleus i.e. unstable atoms (radioactive) might be referred to as radioisotopes. Many isotopes are stable and NOT radioactive i.e. most of the atoms that make up you and the world around you!
- $^{1}_{1}H$, $^{2}_{1}H$ and $^{3}_{1}H$ are the three isotopes of hydrogen with mass numbers of 1, 2 and 3, with 0, 1 and 2 neutrons respectively, but all have 1 proton. Hydrogen-1 is the most common, there is a trace of hydrogen-2 naturally but hydrogen-3 is very unstable and is used in atomic fusion weapons.
- $^{3}_{2}He$ and $^{4}_{2}He$ are the two isotopes of helium with mass numbers of 3 and 4, with 1 and 2 neutrons respectively but both have 2 protons. Helium-3 is formed in the Sun by the initial nuclear fusion process. Helium-4 is also formed in the Sun and as a product of radioactive alpha decay of an unstable nucleus. An alpha particle is a helium nucleus, it picks up two electrons and becomes the atoms of the gas helium.
- $^{23}_{11}Na$ and $^{24}_{11}Na$ are the two isotopes of sodium with mass numbers of 23 and 24, with 12 and 13 neutrons respectively but both have 11 protons. Sodium-23 is quite stable e.g. in common salt (NaCl, sodium chloride) but sodium-24 is a radio-isotope and is a gamma emitter used in medicine as a radioactive tracer e.g. to examine organs and the blood system.

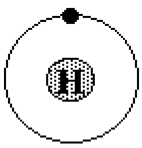
- **The relative atomic mass of an element is the average mass of all the isotopes present compared to 1/12th of the mass of carbon-12 atom (^{12}C = 12.00000 ie the standard).**
- **DO NOT CONFUSE ISOTOPES and ALLOTROPES - see Appendix 3.**

The Electronic Structure of Atoms - rules to be learned

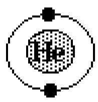
(electron configuration, arrangement in shells or energy levels)

- The electrons are arranged in **energy levels** or **shells** around the nucleus and with increasing distance from the nucleus.
- **Each electron in an atom is in a particular energy level** (or shell) and the electrons must occupy the lowest available energy level (or shell) available nearest the nucleus.
- When the level is full, the next electron goes into the next highest level (shell) available.
- There are rules about the maximum number of electrons allowed in each shell and you have to be able to work out the arrangements for the first 20 elements (see the Periodic Table diagrams further down).
    - **The 1st shell has a maximum of 2 electrons**
    - **The 2nd shell has a maximum of 8 electrons**
    - **The 3rd shell has a maximum of 8 electrons**
        - (only up to atomic number 20, 18 after that, but that's for advanced level work!))
    - **The 19th and 20th electrons go into the 4th shell (limit of GCSE knowledge)**
- **If you know the atomic (proton) number, you know it equals the number of electrons in a neutral atom, you then apply the rules to work out the electron arrangement (configuration).**

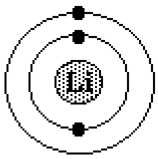
Examples: diagram, symbol or name of element (Atomic Number = number of electrons in a neutral atom), shorthand electron arrangement

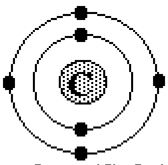


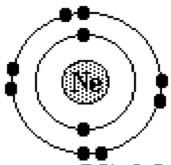
On Period 1 hydrogen (1) 1
• = electron
(atomic no.)
or
(electrons)



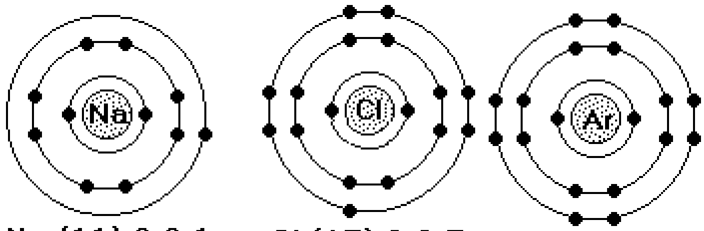
helium (2) 2 (2 elements only)



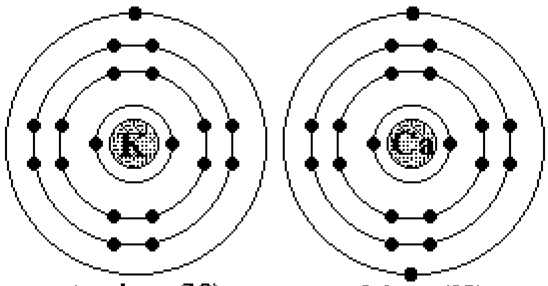




On Period 2 lithium (3) 2.1 to carbon (6) 2.4 to neon (10) 2.8 (3 of the 8 elements)



On Period 3 Na (11) 2,8,1 to Cl (17) 2,8,7 Ar (18) 2,8,8 (3 of the 8 elements)



On Period 4 potassium (19) 2.8.8.1 calcium (20) 2.8.8.2 to Kr [2.8.18.8],

Just the first 2 of the 18 elements of Period 4 are shown, the rule for 3rd shell changes from element 21 Sc onwards (studied at Advanced level)

So which electron arrangements are stable and which are not?

- When an atom has its **outer level full** to the maximum number of electrons allowed, the atom is **particularly stable electronically** and **very unreactive**.
    - This is the situation with the **Noble Gases**: He is [**2**], neon is [**2,8**] and argon is [**2,8,8**] etc.
    - There atoms are the most reluctant to lose, share or gain electrons in any sort of chemical interaction because they are so electronically stable - **most of chemistry is about what outer electrons do or don't**!
    - [**2**],[**2,8**] and [**2,8,8**] etc. are known as the '**stable Noble Gas arrangements**', and the **atoms of other elements try to attain this sort of electron structure when reacting** to become more stable.
- The **most reactive metals have just one outer electron**.
    - These are the **Group 1 Alkali Metals**, lithium [**2,1**], sodium [**2,8,1**], potassium [**2,8,8,1**]
    - With one outer shell electron, they have one more electron than a stable Noble Gas electron structure.
    - So, they **readily lose the outer electron** when they chemically react to try to form (if possible) one of the stable Noble Gas electron arrangements - which is why atoms react in the first place!
- The **most reactive non-metals** are just **one electron short of a full outer shell**.
    - These are the **Group 7 Halogens**, namely fluorine [**2,7**], chlorine [**2,8,7**] etc.
    - These atoms are one electron short of a stable full outer shell and **seek an 8th outer electron to become electronically stable** - yet again, this is why atoms react!
    - They readily gain an outer electron, when they chemically react, to form one of the stable Noble Gas electron arrangements either by sharing electrons (in a

covalent bond) or by electron transfer forming a singly charged negative ion (ionic bonding).

The Periodic Table and Electronic Structure - more patterns!

selected Elements of the Periodic Table are shown below

Revision of its structure - atomic number, symbol

Group Number	1	2				G	R	O	U	P			3	4	5	6	7	8 or 0
Group Name	Alkali Metals	parts of the PERIODIC TABLE atomic numbers 1-38, 49-54															The Halogens	Noble Gases
Period 1							H 1											Helium 2, He
Period 2	Lithium 3, Li	Be 4											B 5	C 6	N 7	O 8	Fluorine 9, F	Neon 10, Ne
Period 3	Sodium 11, Na	Mg 12											Al 13	Si 14	P 15	S 16	Chlorine 17, Cl	Argon 18, Ar
Period 4	Potassium 19, K	Ca 20	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30	Ga 31	Ge 32	As 33	Se 34	Bromine 35, Br	Krypton 36, Kr
Period 5	Rubidium 37, Rb	Sr 38	T	R	A M	N E	S T	I A	T L	I S	O	N	In 49	Sn 50	Sb 51	Te 52	Iodine 53, I	Xenon 54, Xe

- The elements are laid out in order of Atomic Number.
- Originally they were laid out in order of ' relative atomic mass' (the old term was 'atomic weight'). This is not correct for some elements now that we know their detailed atomic structure in terms of protons, neutrons and electrons, and of course, their chemical and physical properties.
- For example: Argon (at. no. 18, electrons 2,8,8) has a relative atomic mass of 40. Potassium (at. no. 19, electrons 2,8,8,1) has a relative atomic mass of 39. Argon, in terms of its physical, chemical and electronic properties is clearly a Noble Gas in Group 8 (0). Likewise, potassium is clearly an Alkali Metal in Group 1.
- Hydrogen, 1, H, **does not readily fit into any group**
- A Group is a vertical column of chemically and physically similar elements e.g. Group 1 The Alkali Metals (Li, Na, K etc.), Group 7 The Halogens (F, Cl, Br, I etc.) and Group 8 or 0 The Noble Gases (He, Ne, Ar etc.). The group number equals the number of electrons in the outer shell (e.g. chlorine's electron arrangement is 2.8.7, the second element down Group 7 on period 3).
- A Period is a horizontal row of elements with a variety of properties (left to right goes from metallic to non-metallic elements. All the elements use the same number of electron shells which equals the period number (e.g. sodium's electron arrangement 2.8.1, the first element in Period 3).
- The ten elements Sc to Zn are called the Transition Metals Series and form part of a period between Group 2 and Group 3 from Period 4 onwards.
- Below are the electron arrangements for elements 1 to 20 set out in Periodic Table format (Hydrogen and The Transition metals etc. have been omitted). When you move down to the next period you start to fill in the next shell according to the maximum electrons in a shell rule (see previous section)

Group number and Name	Group 1 The Alkali Metals	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7 The Halogens	Group 0 Noble Gases
Period 1	$_{1}$**H** hydrogen doesn't really fit in any group The electron arrangements of the first twenty elements							$_{2}$**He** **2**
Period 2	$_{3}$**Li** **2.1**	$_{4}$**Be** **2.2**	$_{5}$**B** **2.3**	$_{6}$**C** **2.4**	$_{7}$**N** **2.5**	$_{8}$**O** **2.6**	$_{9}$**F** **2.7**	$_{10}$**Ne** **2.8**
Period 3	$_{11}$**Na** **2.8.1**	$_{12}$**Mg** **2.8.2**	$_{13}$**Al** **2.8.3**	$_{14}$**Si** **2.8.4**	$_{15}$**P** **2.8.5**	$_{16}$**S** **2.8.6**	$_{17}$**Cl** **2.8.7**	$_{18}$**Ar** **2.8.8**
Period 4	$_{19}$**K** **2.8.8.1**	$_{20}$**Ca** **2.8.8.2**						

- The first element in a period has one outer electron (eg sodium Na 2.8.1), and the last element has a full outer shell (e.g. argon Ar 2.8.8)
- Apart from hydrogen (H, 1) and helium (He, 2) the last electron number is the group number.
- and the number of shells used is equal to the Period Number.

APPENDIX 1. The Alpha Particle Scattering Experiment

- The Rutherford and Marsden scattering experiment.
- When alpha particle beams are fired on very thin layers of metals (eg very fine gold leaf) the results were found to be rather surprising to scientists of the early 20th century
- By using a 360° charged particle detection system it was found that
  - most particles passed through un-deflected
  - a small proportion were deflected slightly
  - and about 1 in 20,000 were 'bounced' back through an angle of over 90°
- From a detailed mathematical analysis of the scattering results, the only 'model' which could account for the pattern was an atom of ...
  - mainly empty space (why most alpha particles passed through)
  - a positive centre causing deflection (like charges repel, alpha particles are positively charged and so are the 'later to be discovered' protons in the nucleus)
  - a tiny dense centre of similar or greater charge or mass to an alpha particle (which we now call the nucleus)
- Earlier theories of atomic structure, eg the 'plum pudding' model in which 'protons' and 'electrons' were scattered or arranged evenly across the atom
- It was the only model that could explain the scattering of the high speed alpha particles by a small dense and positive atomic centre.

- **Later experiments showed that the out bits could be knocked off atoms and these had a very tiny mass and a negative charge, in other words the electron!**

Appendix 2. Atomic structure diagrams - some variations!

Different styles of atomic structure and diagram notation

eg lithium: atomic number = 3 protons in the nucleus

mass number = 7 = 3 protons + 4 neutrons in the nucleus

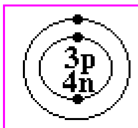
Li element symbol (no sub-atomic particle information)

$_{3}Li$ symbol with atomic number (often seen in periodic table)

^{7}Li symbol with mass number (often used to indicate particular isotope)

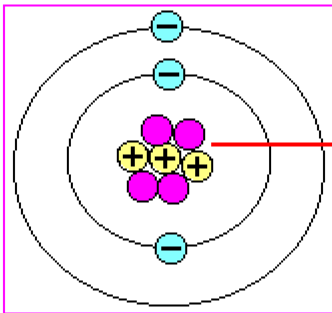
$^{7}_{3}Li$ full sub-atomic particle composition (often seen in periodic table)

Two representations of this Lithium atom



• = electron
p = proton
n = neutron

Abbreviated version



⊖ = electron (negative)
⊕ = proton (positive)
● = neutron (neutral)
nucleus

A more detailed version (but a bit cumbersome for bigger atoms!)

Appendix 3. Allotropes - don't confuse with isotopes!

- **Isotopes are atoms of the same element with different masses due to different numbers of neutrons in the nucleus**. Same protons and electrons. eg atomic number 6 = 6 protons = carbon, but there can be 6, 7 or 8 neutrons giving isotopes of carbon-12, 13 or 14.
- Oxygen atoms usually form **'stable' O_2 oxygen molecules** (also called dioxygen), BUT they can form an **unstable molecule O_3 ozone** (also called trioxygen). The mass of the oxygen atoms in each of the molecules is mainly 16 (99.8%), and about 0.2% of two other stable isotopes of masses 17 and 18. **Whatever isotope or isotopes make up the molecule, it doesn't affect the molecular structure or the respective chemistry of the O_2 or O_3 molecules.**
- However, what sometimes confuses the issue is the fact that **oxygen O_2 and ozone O_3 are examples of allotropes**. Allotropes are defined as different forms of the same element in the same physical state. They are usually chemically very similar but always physically different in some way. eg (i) O_2 and O_3 are both gases but have different densities, boiling points etc. Other examples are (ii) **graphite, diamond and buckminsterfullerene are all solid allotropes of the element carbon** and have significantly different physical and in some ways chemical properties and (iii) **rhombic and monoclinic sulphur have different geometrical crystal structures**, that is different ways of packing the sulphur atoms (which are actually both made up of different packing arrangements of S_8 molecules). They have different solubilities and melting points.
    - Again it doesn't matter which isotopes make up the structure of any of an elements allotropes described above.
- So to summarise by example!,

- eg **oxygen-16, 17 or 18 are isotopes of oxygen with different nuclear structures due to different numbers of neutrons**,
- and **O_2 and O_3 are different molecular structures of the same element in the same physical state and are called allotropes** irrespective of the isotopes that make up the molecules.