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Part of this site was last updated on 2nd January 2007.

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Introduction

A simple model of the structure of the atom is that atoms are thought to consist of three sub-atomic particles - protons, neutrons and electrons.

Protons and neutrons have a mass almost equal to that of a hydrogen atom. Virtually all the mass of the atoms is contained in the nucleus which occupies only a small fraction of the total volume of the atom. The nucleus has an incredibly high density which suggests that the particles are held together by very powerful forces. Protons and neutrons are contained in the nucleus, electrons move around the nucleus.

Particle	Symbol	Charge (Rel. to proton)	Mass (Rel. to proton)
Proton	p	+1	1
Neutron	n	0	1
Electron	e	-1	0.00055

An atom is electrically neutral and contains equal numbers of protons and electrons. The atomic number, Z, equals the number of protons in the nucleus and also equals the number of electrons. The mass number, A, is the number of protons plus the number of neutrons in the nucleus.

$$A = Z + n$$

Isotopes

J. J. Thomson discovered that many elements could have atoms with different masses. These were called isotopes. All the isotopes of one particular element have the same number of protons and electrons but different number of neutrons, hence different mass numbers.

For example chlorine has two isotopes, chlorine-35 and chlorine-37. Naturally occurring chlorine consists of 75% chlorine-35 and 25% chlorine-37. The relative atomic mass of chlorine is therefore (35x75 + 37x25)/100 = 35.5.

Radioactivity

The nuclei of some atoms are unstable. They can break down forming new elements and giving out radiation. Such atoms are said to be radioactive. The main radioactive processes are:

1. Alpha decay - the nucleus emits an alpha particle which is a helium nucleus - containing two protons and two neutrons.

2. Beta decay - the nucleus emits a beta particle which is an electron.

3. Gamma emission - this is high energy electromagnetic radiation from the nucleus. This occurs along with alpha or beta decay.

4. Fission - the nucleus splits into two to give two new nuclei, each of smaller mass.

In the first two cases, the radiations can knock out electrons from atoms to produce ions. Hence they are called ionising radiations.

Alpha particles are the least penetrating in that they can be stopped by paper. Beta particles are stopped by aluminium sheet. It takes a thick piece of lead to stop gamma rays.

Nuclear Equations

Alpha Decay involves the emission of an alpha particle which consists of two protons and two neutrons. Hence the atomic number decreases by two and the mass number by four. This can be represented by a nuclear equation:

$$^{229}_{90}\mathrm{Th} \longrightarrow {}^{4}_{2}\mathrm{He} + {}^{225}_{88}\mathrm{Ra}$$

Beta Decay involves the loss of an electron from the nucleus not from the electron shells. For this to happen a neutron turns into a proton and an electron. The proton remains in the nucleus hence the atomic number increases by one and the mass number remains unchanged.

$$^{216}_{84}\mathrm{Po} \longrightarrow {}^{216}_{85}\mathrm{At} + {}^{0}_{-1}\mathrm{e}$$

Radioactive Decay - Half-life

Different radioactive elements decay at widely different rates. It is always found that the rate of decay is proportional to the amount of radioactive substance remaining. This means that the time taken for the radioactivity to decay to half the original amount is always the same for a particular element. This time is called the half-life. It can be found from a graph of count rate against time.

The half-lives of radioisotopes present in radioactive waste from nuclear power stations have an important bearing on the pollution problems they present. Radioactive elements which have a short half-life, e.g. 10 years will have largely decayed in 50 years, whereas ones with a half-life of 1000 years will have hardly dropped in activity at all.

Arrangement of Electrons Around the Nucleus

The evidence for the way electrons are arranged in atoms comes largely from the study of light and other types of electromagnetic radiation given out by atoms.

Light is one form of electromagnetic radiation. Like all electromagnetic radiation, it behaves like a wave, with a characteristic wavelength and frequency. These are related by the Wave Equation:

$$c = \nu \times \lambda$$

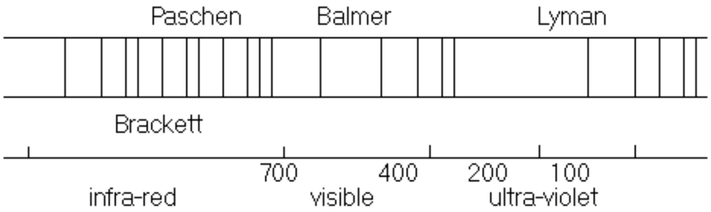
Where c = velocity, ν = frequency, λ = wavelength.

When we use the term light we normally mean the visible light to which our eyes respond. Visible light is only a part of the electromagnetic spectrum.

(Note: ν and λ are the lowercase Greek letters Nu and Lambda.)

Line Spectra

When sodium chloride is heated strongly in a bunsen flame, it gives off a brilliant yellow light. Some gaseous materials emit light when they are subjected to high voltages in electric discharge tubes. If the light emitted by these substances is examined using a spectroscope, it is found not to consist of a continuous range of colours like the spectrum of white light or the colours of a rainbow. Instead the light is composed of separate lines of different colours. This kind of spectrum is called line emission spectra. Each element has its own characteristic set of lines different from those of any other element. Consequently elements can be identified by a study of their line emission spectra. Each line corresponds to light of a particular frequency. The emission spectrum of hydrogen is shewn below and consists of a number of separate sets of lines or series of lines. These series of lines are named after their discoverers.

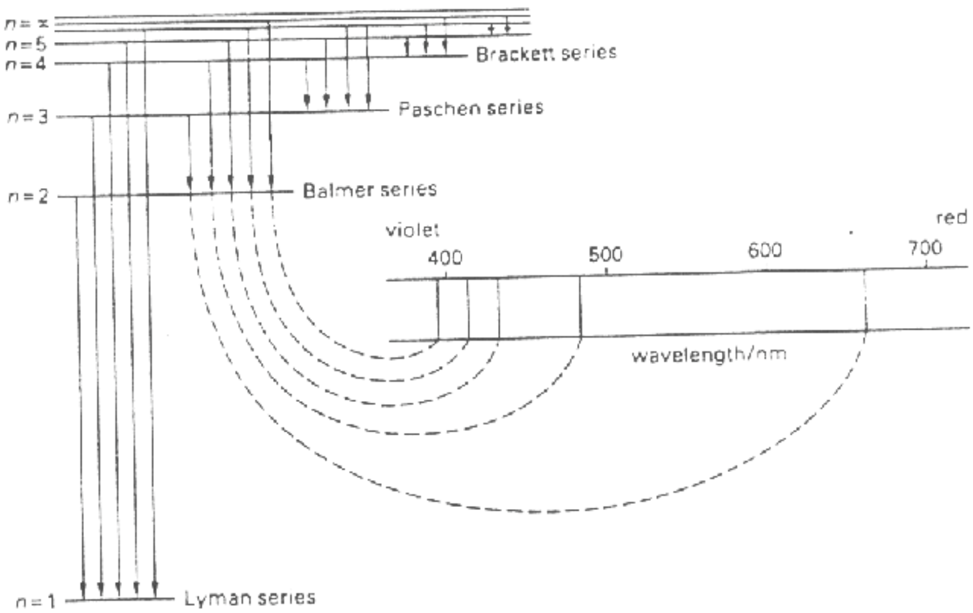


In each series, the intervals between the frequencies of the lines become smaller towards the high frequency end of the spectrum until the lines run together or converge.

Bohr explained line spectra as follows. Electrons in an atom can exist only at certain energy levels. Under normal conditions the electrons in an atom fill up the lowest energy level first. When sufficient energy is supplied to an atom, it may be possible to promote an electron from a lower energy level to a higher one. Since the electron is unstable in the higher energy level, it will drop back to the lower energy level emitting the excess energy as radiation. The energy difference between the two levels can only have certain fixed values. Since the energy of any radiation is determined by its frequency, it means that the radiation emitted when the electron falls from a higher to a lower energy level can have only certain fixed frequencies (certain specific colours). This small amount of energy emitted is called a quanta given by the relationship:

$$E = h \times \nu$$

The diagram below shews the relationship between the main series in the hydrogen spectrum and the energy levels with which they are associated.



Electron transitions to the n = 2 level occur in the visible region and correspond to the Balmer Series. Transitions to the n = 1 level release more energy; lines appear at higher frequencies in the UV region and correspond to the Lyman Series.

If sufficient energy is given to an atom to excite an electron beyond the highest energy level, the electron will escape from the atom: ionisation has occurred.

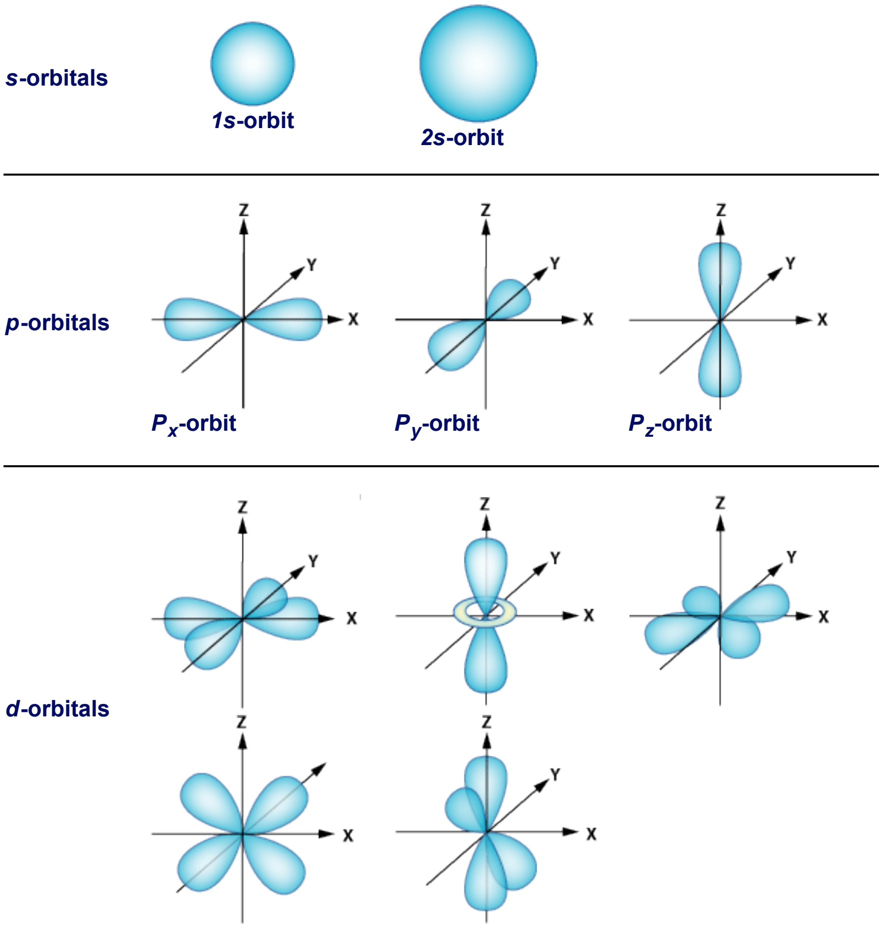
Wave Properties of the Electron

In the 1920s doubt was cast upon the validity of treating the electron as a revolving particle. This new evidence laid the foundation for what is called wave-mechanics.

In order to explain some properties, we regard electrons as particles; to explain others we regard them as waves, i.e. they appear to have a dual nature. By applying wave-mechanics, scientists came up with a picture of electron distribution. The main distinction between the planetary and orbital models of the atom is that while the planetary model assumes that electrons keep to fixed "shells" around the nucleus, the orbital model is based on the probability of finding an electron in a certain volume of space. The volume of space in which there is a 95% chance of finding the electron is called the atomic orbital. There are four types of orbitals whose shapes have been worked out using wave-mechanics. They are referred by the letters ***s***, ***p***, ***d*** and ***f***. The shapes of ***s***, ***p*** and ***d*** orbitals are shewn below. It can be seen that the shape of an ***s*** orbital is spherically

symmetrical about the nucleus.

Electronic Structure: Sub-shells and Orbitals



Bohr's theory for the arrangement of electrons placed electrons in shells. The shells are labelled by giving each one a principal quantum number, n. For the first shell, n = 1, for the second shell n = 2, etc. The maximum number of electrons which can be held in the first three shells are:

first shell	**n = 1**	**2 electrons**
second shell	**n = 2**	**8 electrons**
third shell	**n = 3**	**18 electrons**

Electrons are arranged so that the lowest energy shells are filled first. Sodium, Z = 11, has an arrangement 2.8.1.

It has been established that the shells are themselves split into sub-shells. The sub-shells are designated *s*, *p*, *d*, and *f*. The different types of sub-shells can hold different numbers of electrons.

For n = 1	**2 electrons in *s* sub-shell**	
For n = 2	**2 electrons in *s* sub-shell**	
	6 electrons in *p* sub-shell	**a total of 8 electrons**
For n = 3	**2 electrons in *s* sub-shell**	
	6 electrons in *p* sub-shell	
	10 electrons in *d* sub-shell	**a total of 18 electrons**
For n = 4	**2 electrons in *s* sub-shell**	
	6 electrons in *p* sub-shell	
	10 electrons in *d* sub-shell	
	14 electrons in *f* sub-shell	**a total of 32 electrons**

The *s*, *p*, *d*, and *f* sub-shells are themselves divided further into atomic orbitals. An electron in a given orbital can be found in a particular space around the nucleus.

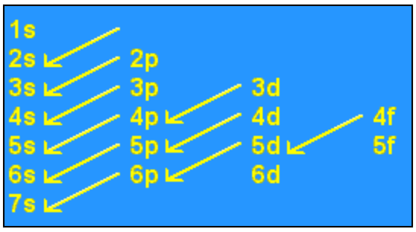
An *s* sub-shell contains one *s* orbital
A *p* sub-shell contains three *p* orbitals
A *d* sub-shell contains five *d* orbitals
An *f* sub-shell contains seven *f* orbitals.

Note that there are two electrons in each orbital.

The arrangement of electrons in shells and orbitals is called the electronic arrangement. The orbitals are filled in order of increasing energy.

E.g.	**Hydrogen**	$1s^1$
	Helium	$1s^2$
	Lithium	$1s^2 2s^1$
	Beryllium	$1s^2 2s^2$
	Boron	$1s^2 2s^2 2p^1$
	Carbon	$1s^2 2s^2 2p^2$
	Nitrogen	$1s^2 2s^2 2p^3$
	Oxygen	$1s^2 2s^2 2p^4$
	Fluorine	$1s^2 2s^2 2p^5$
	Neon	$1s^2 2s^2 2p^6$
	Sodium	$1s^2 2s^2 2p^6 3s^1$

The order of filling orbitals can be determined by using the diagram below:



Electronic Structure and the Periodic Table

Chemical similarities amongst elements in the same group arise because of their similar electronic configurations in their outer quantum level. The Noble Gases all have completely filled *s* and *p* orbitals in their outer quantum level, e.g. He $1s^2$, Ne $1s^2 2s^2 2p^6$. The Alkali Metals all have an outer electronic configuration of s^1, the Halogens $s^2 p^5$.

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