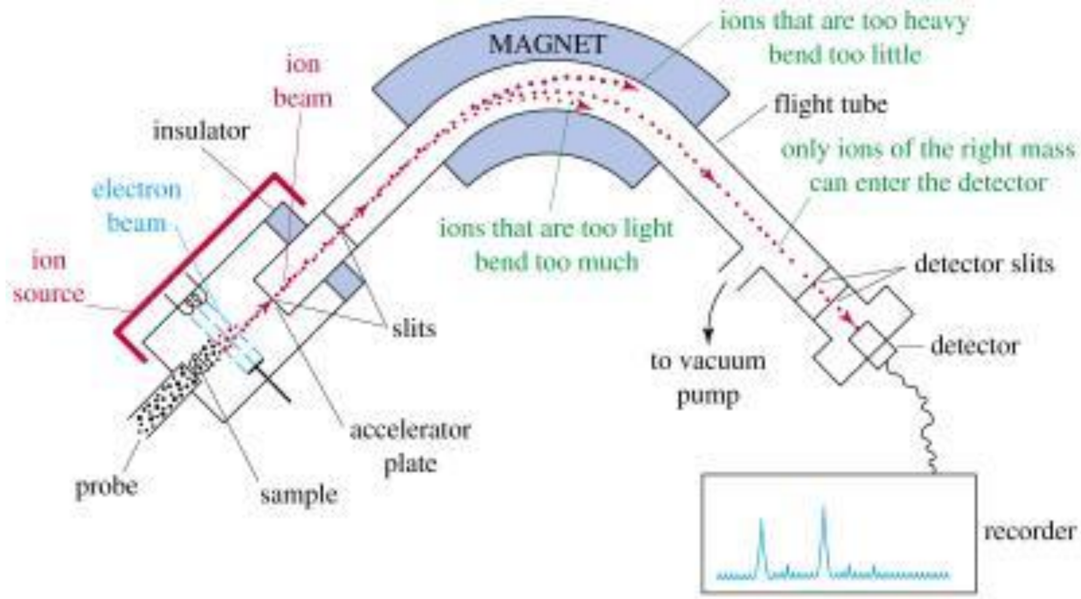
Atomic Structure

I. The Atom
    A. Atomic theory: Devised in 1807 by John Dalton, states that all matter is made up of a small number of different kinds of atoms that are indivisible and indestructible but which combine in whole number ratios to form compounds.
        i. this is still mostly true, however we now know that atoms are not indivisible and are actually made up of much smaller subatomic particles.
        ii. We now know that the valence shell electrons are crucial to chemical interactions show knowing the atomic structure of atoms allows us to understand how atoms join to form compounds and why different atoms act in different ways
    B. Subatomic Particles: There are three important subatomic particles, the proton, the neutron, and the electron
        i. Mass :The proton and neutron have a significantly larger mass than the electron, nearly two thousand times larger (1840)
            a. The protons and neutrons are tightly bound together to form the nucleus; therefore, the nucleus contains nearly all of the mass of the atom
        ii. Charge: the proton carries a charge of +1 and the electron carries a charge of -1, whilst the neutron is electrically neutral
    C. Atomic notation
        i. Atomic number: the atomic number is equal to the number of protons in an element. All isotopes of an element have the same number of protons and thus have the same atomic number
        ii. Mass number: the sum of the protons and neutrons is the mass number. The mass number can by isotope.
        iii. The atomic number and the mass number of an element may be indicated by a subscript and a superscript respectively, placed before the symbol of the element
        iv. In order to preserve electrical neutrality, the number of electrons is equal to the number of protons
    D. Atoms can gain or lose electrons to form ions but they never gain or lose protons
            a. if an electron is gained, the element assumes a negative charge
            b. if an electron is lost, the element assumes a positive charge
    E. Isotopes: Many elements are composed of slightly differing types of atoms know as isotopes. They possess different numbers of neutrons
        i. Radioactive isotopes of all elements can be produced by exposing the natural element to a flux of slow moving neutrons in a nuclear reactor.

a. Radioisotopes can be used as tracers because they behave chemically, and thus biologically, in the same manner to the stable isotopes.

II. The Mass Spectrometer: the mass spectrometer is an instrument which separates particles according to their masses and records relative proportions of these




A. How it works
   i. Substance is first converted to atoms or molecules is vapor phase
   ii. These are then turned into positive ions
   iii. The ions are then accelerated
   iv. The fast moving ions are deflected- the lighter the molecule, the greater the deflection
   v. Finally, particles of a particular mass (which can be adjusted) are detected
   vi. This is all performed under high vacuum
   vii. The mass spec is used to determine the natural abundance of isotopes of a particular element

III. Electron Arrangement
   A. Emission Spectrum: The study of emission of light (or rather, electromagnetic radiation) by atoms and ions is the most effective technique for deducing the electronic structure of atoms

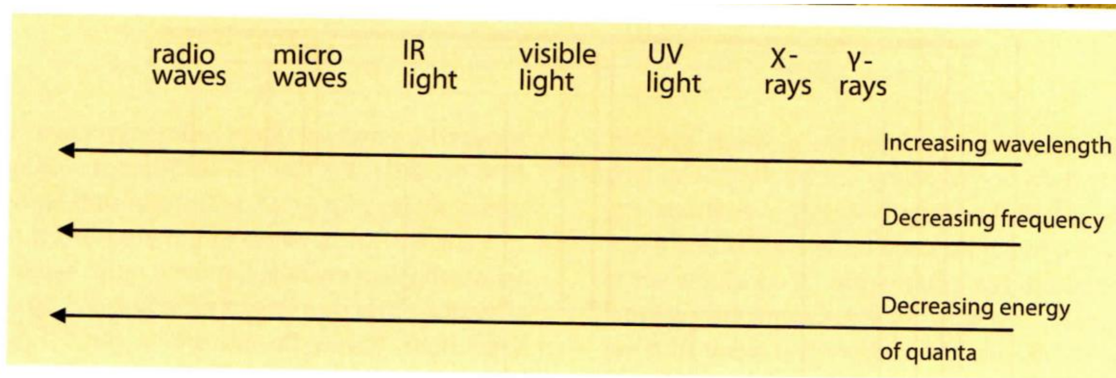



*Figure 210 The electromagnetic spectrum*

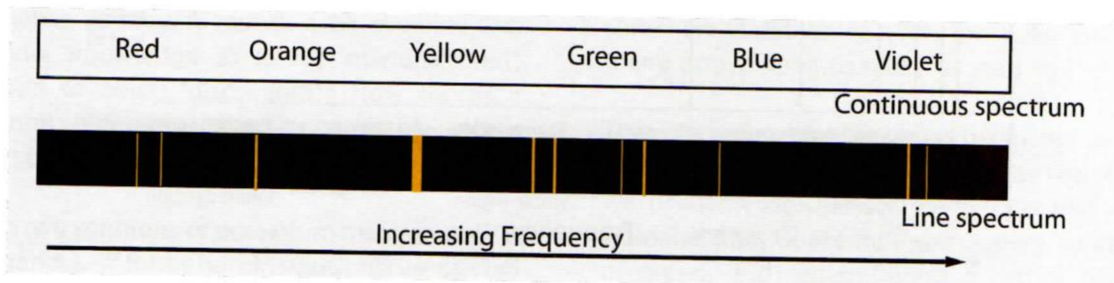



*Figure 211 Continuous and line spectra*

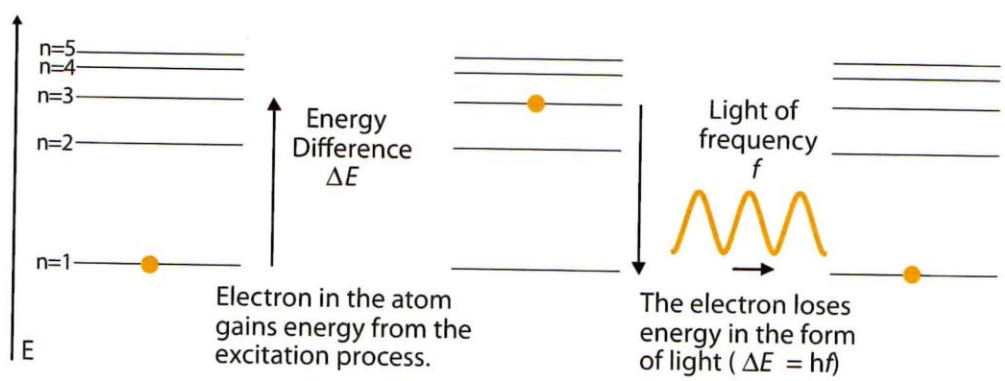



*Figure 212 The origin of line spectra*

i. This is the best evidence for the fact that electrons in an atom surround the nucleus in certain allowed energy levels or orbitals
ii. Excited electrons will often emit light of a characteristic color, this is the basis of flame tests
iii. When an electron is excited, its electrons gain energy and move to a higher energy level. In order to return to lower energy levels, the electron must lose energy, and thus will give off light.
iv. The color is determined by the frequency of the light according to the equation E=hf (where h is Plank's constant)

v. By studying the frequencies of the lines in the emission spectrum of an element, the energies of various energy levels in it atoms may be found.

B. Atomic emission spectrum of Hydrogen

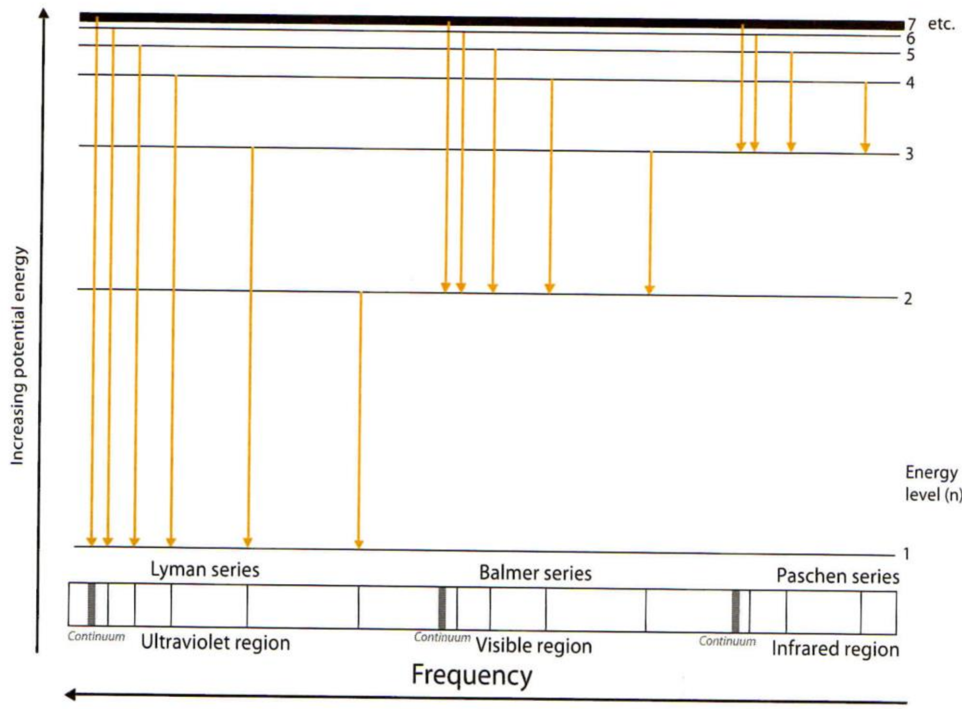



*Figure 213 Explanation of the atomic emission spectrum of hydrogen*

i. The emission spectrum of hydrogen is the simplest emission spectrum because it contains only one electron

ii. When a potential difference is applied across hydrogen gas at low pressure, electromagnetic radiation is emitted,

iii. This is not uniform, but rather at distinct, bright lines, indicating the existence of only certain allowed electron energy levels

C. Electronic Structure and the Periodic Table

i. The most stable energy levels are those that are closest to the nucleus and these levels fill before the higher levels

ii. There is a maximum number of electrons that each energy level can hold

iii. The first energy level can hold 2 electrons, the next two energy levels can each hold 8.

iv. Fill in structure for the first 20 elements

HL

IV. Electron Configuration

A. The nucleus of the atom is surrounded by electrons arranged in specific energy levels and sub levels. The different sublevels differ in the shape of the electron distribution. Each energy level is divided into orbitals each of which can contain up to two electrons, which must have opposite spin according to the Pauli exclusion principle.

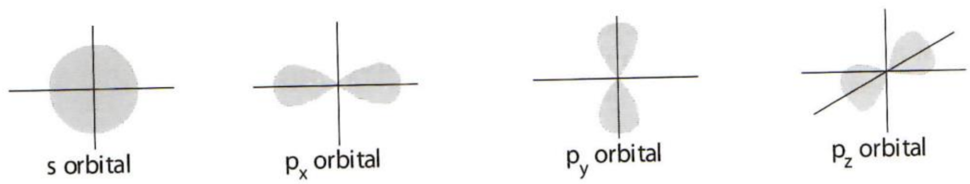



*Figure 215 An illustration of the electron distribution in s- and p-orbitals*

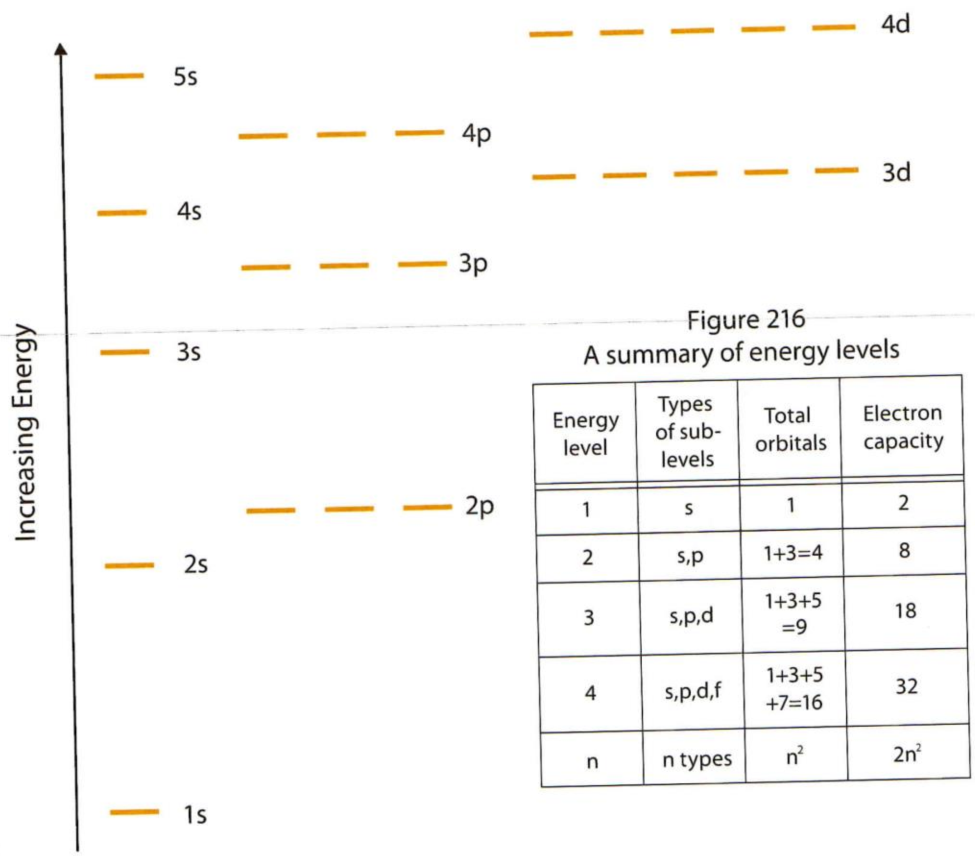



Figure 216
A summary of energy levels

| Energy level | Types of sub-levels | Total orbitals | Electron capacity |
|---|---|---|---|
| 1 | s | 1 | 2 |
| 2 | s,p | 1+3=4 | 8 |
| 3 | s,p,d | 1+3+5 =9 | 18 |
| 4 | s,p,d,f | 1+3+5 +7=16 | 32 |
| n | n types | $n^2$ | $2n^2$ |

*Figure 216 The electron energy levels in a typical atom*

B. Quantum Numbers: Electrons have a wave as well as a particle nature
   i. The wave like nature can be described by the Schrodinger Equation. This equation involves four constants, called quantum numbers, and a solution for the equation is only possible if the values of the quantum numbers lie within certain limits.
   ii. The principal quantum number (n) must be a positive integer

iii. The azimuthal quantum number (l) can have interger values from zero to n-1

iv. The magnetic quantum number (m) can have integer values from – to + (including zero), whilst the spin (s) can me = or – ½,

v. Pauli exclusion principle: no two electrons in a given atom can have the same four quantum numbers

V. Ionization Energies: the ionization energy of an atom is the minimum amount of energy required to remove a mole of electrons from a mole of gaseous atoms to form a mole of gaseous ions.

A. Ionization reactions are endothermic because energy must be put in to remove the electrons

B. The more electrons that have been removed from an atom, the greater the energy required to remove the next electron

C. When successive electrons are all in the same energy level, the energy required increases because of a reduction in the amount of electron-electron repulsion and hence a greater nuclear-electron attraction.

D. Going down the periodic table, the ionization energies decrease because the outermost electrons are in higher energy levels and hence further from the nucleus

E. Ionization energies increases across the period because the increase in the charge on the nucleus which increases the effective nuclear charge. This is not constant across the period because it can vary based on the filling of the orbitals

F. Isoelectronic species are those which have the same electronic structure.