Topic 1 – Atoms and Electron Structure
Revision Notes

1) Atomic Structure

	Relative mass	Relative charge
Proton	1	+1
Neutron	1	0
Electron	1/2000	-1

- The nucleus contains almost all of the mass of an atom because that is where the protons and neutrons are found
- The nucleus of an atom contains all of the positive charge
- The electrons are outside the nucleus and, therefore, so is the negative charge
- Atomic number = number of protons in the nucleus
- Mass number = number of protons and neutrons in the nucleus
- Number of neutrons = mass number – atomic number
- Number of electrons = number of protons (in a neutral atom)

$^{9}_{4}Be$ Mass number = 9 Atomic number = 4

4 protons, 5 neutrons, 4 electrons

2) Isotopes and ions

- Isotopes are atoms with the same number of protons but different numbers of neutrons (and different masses)
- For example, chlorine has two isotopes ^{35}Cl and ^{37}Cl. Both have 17 protons but they have 18 and 20 neutrons, respectively
- Isotopes of an element have the same chemical properties because they have the same electron arrangement
- Ions are formed when atoms gain or lose electrons.
- As an atom Cl has 17 electrons. A Cl^- ion has gained one electron so it now has 18.
- As an atom Na has 11 electrons. A Na^+ ion has lost one electron so it now has 10.

3) Relative atomic mass

- Relative atomic mass is the average mass of an atom of an element taking the mixture of isotopes into account. **However, learn the technical definition from definitions sheet**
- To calculate relative atomic mass, add together (mass number x percentage/100) for each isotope

Example:
75% of Cl atoms have a mass number of 35
25% of Cl atoms have a mass number of 37

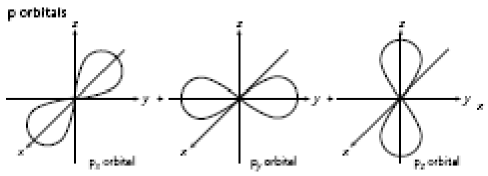
Average mass of a Cl atom = (mass no x percent/100) + (mass no x percent/100)
= (35 x 75/100) + (37 x 25/100)
= 35.5

- For simple molecules, such as O_2 and H_2O, the relative molecular mass is calculated by adding the relative atomic masses of the elements involved, giving 32.0 for O_2 and 18.0 for H_2O
- For giant structures, such as Na_2S and SiO_2, the relative formula mass is calculated by adding the relative atomic masses of the elements involved, giving 68.1 for Na_2S and 60.1 for SiO_2

4) **Orbitals**

- An orbital is a region that can hold up to two electrons with opposite spins
- Orbitals have different shapes called s, p, d, and f (but f orbitals are beyond our syllabus)
- S orbitals are spherical in shape and come in sets of one (which can hold up to 2 electrons)

- P orbitals are hour-glass or egg-timer shaped and come in sets of three (which can hold up to 6 electrons)



- D orbitals come in sets of five which can hold up to 10 electrons

5) **Energy levels (or shells)**

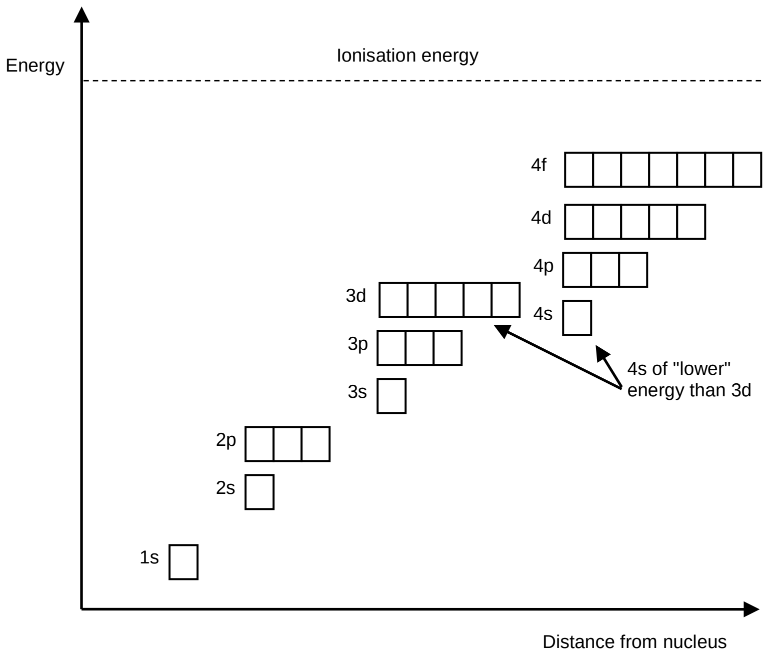
- The first energy level (or shell) only contains an s orbital, labelled 1s
- The first shell can hold up to 2 electrons
- The second energy level contains an s orbital (labelled 2s) and three p orbitals (labelled 2p)
- The second shell can hold up to 8 electrons
- The third energy level contains an s orbital, three p orbitals and five d orbitals
- The third shell can hold up to 18 electrons
- The order in which the orbitals are filled is as follows: 1s 2s 2p 3s 3p 4s 3d 4p
- Note that the 4s fills before the 3d

Some examples of electronic structures are shown below.

Hydrogen	1 electron	$1s^1$
Nitrogen	7 electrons	$1s^2\ 2s^2\ 2p^3$
Sodium	11 electrons	$1s^2\ 2s^2\ 2p^6\ 3s^1$
Sulphur	16 electrons	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
Calcium	20 electrons	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
Iron	26 electrons	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

- In a Cl^- ion, the 18 electrons are arranged $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
- In a Na^+ ion, the 10 electrons are arranged $1s^2\ 2s^2\ 2p^6$
- Transition metals, like iron, lose their 4s electrons **first** (before 3d). Fe^{3+}, with 23 electrons, is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^0\ 3d^5$

- The diagram below shows the relative energies of the orbitals from 1s to 4f



Source: www.chemsheets.co.uk

6) <u>Successive Ionisation Energies</u>

- Evidence that electrons are arranged in shells or energy levels can be obtained by measuring the successive ionisation energies of an element
- The first ionisation energy of an element is the energy needed to remove one mole of electrons from one mole of gaseous atoms i.e.

$$\mathbf{M(g) \rightarrow M^{+}(g) + e^{-}}$$

Note - State symbols are essential in ionisation equations

- In general, ionisation is easier if the nuclear charge is smaller, the electron is further away from the nucleus and there is more shielding from inner electron shells.
- For an element, successive ionisation energies get bigger because the remaining electrons are held more tightly by the unchanged nuclear charge.
- Jumps in ionisation energies occur when going from one energy level (shell) to another. This tells you which group the element is in. The jump in energy occurs because the new energy level is closer to nucleus and less shielded.

7) <u>Blocks in the Periodic Table</u>

- An element can be assigned to the s, p or d block by working out which type of orbital its outermost electron is in
- The s block is groups 1 and 2
- The p block is groups 3 to 8
- The d block is between the s and p blocks