Chapter 3

Topic 3 - Atomic structure, ions and metals

Topic 3 of the Nuffield course extends the model of atomic structure from that met at GCSE. It includes some complicated and involved ideas, and you may want to sit down for an hour with a nice cup of tea, the text book, a pencil and some paper. The chapter presents evidence which is explained by a more sophisticated model of the atom (the Bohr model).

The chapter opens with a discussion of mass spectrometry and the use of this in identifying the makeup of an atomic nucleus. Questions you might be asked would include looking at a mass spectrum and explaining that the peaks come from the occurrence of different isotopes. Write down the definition of an isotope below.

Carbon has two main isotopes. These are $^{12}_{6}\mathrm{C}$ and $^{13}_{6}\mathrm{C}$. There is 99% of the former and 1% of the latter. Explain what the numbers in the symbols mean for each isotope, and work out the relative atomic mass of carbon.

3.1 Periodicity

Periodicity refers to the periodic variation of properties of elements in groups or periods of the periodic table. We plot or tabulate data to show regular patterns of behaviour. You should recognise a pattern in the variation of atomic volume with atomic number. It is important to be familiar with the data we looked at in the first session in the computer room - The variation in **boiling** and **melting points** of the first 18 elements. The sharpest drops in these values occur between carbon and nitrogen, and between silicon and phosphorus. Why should this be?

3.2 Flame colours and emission spectra

When a granule of an ionic compound (or drop of solution) is placed in a flame different colours may be observed. You should know the experimental procedure for experiment 3.3. Different metals give different colours. We were able to see this with more sophistication using the spectrometer, which analysed the light produced by a hydrogen lamp, a sodium lamp, and a mercury lamp.

Movement of electrons from one energy level, or quantum shell, to another is caused by the absorption or emission of a packet of energy or **quantum** (a photon of light for instance). The energy of this photon is equal to the energy gap between the two energy levels or quantum shells. The energy levels in an atom have definite separations. These are called **discrete** energy levels. Different atoms have different spacings of their energy levels, and the energy levels are filled by different numbers of electrons. This means that the energy transitions are specific for individual elements. A large energy gap means that the photon emitted has a larger energy, and so the light is further from the red end of the spectrum towards the violet.

Flame colours are created by the strong emission lines in the spectrum of an elephant. An emission spectrum is produced by the transitions of many atoms of the same type. In summary: an emission spectrum is produced when electrons that have been excited to higher energy levels emit energy as they return to lower energy levels. For example in the visible part of the electromagnetic spectrum, the emission spectrum is seen as bright coloured lines that are specific for an individual element. You should know the flame colours of the elements we tested.

3.3 Ionisation energy

You need to know the equations for first second, third etc. ionisation energies.

$$\mathrm{M_{(g)}} \longrightarrow \mathrm{M^{+}_{(g)}} + \mathrm{e^{-}} \tag{3.1}$$

Write down equations similar to equation 3.1 for the second and third ionisation energies of M.

$$\tag{3.2}$$

$$\tag{3.3}$$

You should also know the equation for electron affinity (the energy change when a species gains an electron).

$$\mathrm{X}_{(\quad)} \longrightarrow \tag{3.4}$$

The patterns in ionisation energies give us evidence for the electronic structures of atoms. You may be asked to interpret the sequence of successive ionisation energies for an element and work out what group it is and explain the reasons for any ‘jumps’ in the data. For instance, a group 4 element will have a jump between the fourth and fifth ionisation energies because there is a change in the quantum shell from which the electrons are being removed, and electrons in the next quantum shell are closer to the nucleus and therefore experience a greater effective nuclear charge. You may also be asked to interpret the pattern of ionisation energies across a period, or down a group. This leads to the arrangement of electrons in quantum shells and sub-shells.

3.4 Electronic structure

The electron distribution in atoms is quite complex. There are the major energy levels, which we give the catchy symbol n. These range from $n = 1$ to $n = \infty$. (Question: What is it called if we move an electron of an atom to the energy level where $n = \infty$?) Each of these major energy levels is sub-divided into quantum sub-shells. Each energy level has n sub-shells. So, $n = 1$ has one sub-shell, $n = 2$ has two, $n = 3$ has three, and so on. The letters used to label the sub-shells are rooted in the history of physical chemistry and don’t really follow a pattern. They are, in order, s, p, d and f[1]. You can remember them with a mnemonic such as ‘Sid produces delightful fragrances’.

You should know that energies of the sub-levels go $s < p < d < f$. The **bulding-up principle**, also called the *Aufbau* principle from the German for building up,

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[1] These letters come from the appearance of lines on photographic film. Over 100 years ago, when the spectra were first examined closely, nobody knew why the appeared as they did. Scientists grouped the lines according to appearance. Some were **s**harp, some were similar to the **p**rincipal lines in the hydrogen spectrum, some were **d**iffuse, and later some **f**undamental ones were found.

states that levels of lowest energy are filled by electrons first. This means that we start by putting electrons in $1s$, then $2s$, then $2p$ and so on. The order of filling up is:

$$1s\ 2s\ 2p\ 3s\ 3p\ 4s\ 3d\ 4p\ 5s\ 4d\ 5p\ 6s \tag{3.5}$$

You need to know this order, or be able to work it out from the triangle thingy in your notes. You also need to know how many electrons each sub-shell can hold. The blocks in the periodic table will tell you this (the s block is two elements wide, the p block is six elements wide and so on).

s can hold _____ electrons
p can hold _____ electrons
d can hold _____ electrons
f can hold _____ electrons

You should be able to describe the electronic arrangement of atoms *and ions* up to an atomic number of about 56 (which as we all know is barium). The electronic structure of lithium is $1s^2\ 2s^1$. That of phosphorus is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$. You should be able to say whether an element is in the s block or p block and so on. From the examples for Li and P we would say that Li is an s block element (its outermost electron is in an s sub-shell) and P is a p block element (its outermost electron is in a p sub-shell). The d block metals have a specific name, which is. . . ?

3.5 The ionic model

Ions and the ionic model are not much different from the discussion at GCSE. You should (still) be able to draw 'dot-and-cross' diagrams to show the formation of ionic compounds from their atoms. You might like to do The Most Famous Example in the World…Ever below (sodium chloride):

You might be asked to demonstrate an understanding of how someone might tell that a substance contains ions. The classic experiment for this is to test for conductivity, and you should know experiment 3.5.

Because of the large regions surrounding a nucleus, ions are not hard spheres, rather they are a bit 'squashy'. In different environments (for instance, in compounds with different counter-ions) an ion may have a different shape and radius. It is difficult to say where an ion in a structure ends and where another begins, and this again causes problems when we want to talk about the sizes of ions. In fact the usual procedure is to choose a standard ion and measure everything else on a scale defined by that one ion. Unfortunately, the idea of 'standard' is ambiguous and different people use different standards!

It is unlikely that you will be asked to work out the ionic radius of an ion, but you may be asked to select which ion has the largest radius out of a group of ions. You should be able to justify why losing an electron from an atom makes the ionic radius smaller than the atomic radius. You should be able to justify why adding an electron to an atom makes the ionic radius greater than the atomic radius.

3.6 Metallic bonding

You should know that a metallic bond is the result of a giant structure composed of positive metal ions in a 'sea' of mobile electrons. The sharing of this 'sea' of electrons binds the ions tightly into the regular lattice, giving metals their characteristics of

- __________ melting point and boiling point
- __________ electrical and thermal conductivity.

3.7 Conclusion

A lot of the material in this section is new and a bit abstruse. I have tried during lessons to tell you the truth about atomic structure as it is understood now, and that is why some parts have been rather esoteric. You do not need to know all of the detail we have covered in discussion. Remember that this model of the atom is only a model. It has limitations, the primary one being its overwhelming complexity (the maths is *very* hard).