7. Chemical structure

This chapter considers a key area of chemistry, that of chemical structure. It also reviews some of the research findings about learners' ideas about atomic structure and other chemical structures, and introduces some related classroom instruments included in the companion volume.

The structure of the atom

During their secondary education students are expected to learn about the structure of the atom, or – more correctly – to learn about a particular model of the structure of the atom (see the comments on models in Chapter 6).

The usual model of the structure of the atom met at this level consists of the nucleus at the centre of one or more shells of electrons. The electrons are usually shown (in two dimensional diagrams) as being placed on these circular shells (*eg*, see Figure 7.1).
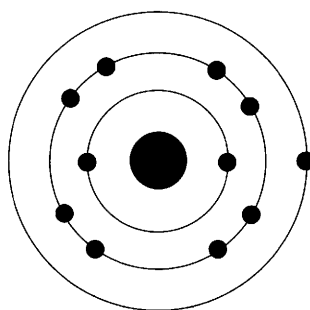

Figure 7.1 A typical representation of an atom

Although this model is perfectly appropriate at this level, those students who take their study of chemistry further (at post-16 level) will need to accept more detailed models (*eg* where electron positions are described by orbitals, which make up sub-shells). It is useful, therefore, for secondary teachers to emphasise that such a diagram only represents a model, and is one of several models that together help us understand matter at the atomic scale.

It is helpful if learners are familiar with a range of representations for molecules and other structures, as this will reinforce the modelling aspect. A mixture of different types of diagram have been deliberately used in preparing classroom materials for the companion volume.

The principles determining atomic structure

Clearly the atomic model is abstract, and a long way from learners' everyday experiences of the world. Students have never directly perceived individual molecules – except perhaps by smell, and that does not provide any insight to molecular structure. The terms 'proton' , 'neutron' and 'electron' are (initially) unfamiliar technical terms, and so need to be learnt by rote. As teachers are well aware, students may often confuse these labels while still mastering the basic model.

More significant than such errors are students' alternative ideas about how and why atoms, molecules and other chemical structures are formed and maintained. From the scientific viewpoint there are three main sets of principles involved:

■ the nucleus is held together by nuclear forces;

■ systems of nuclei and electrons (*ie* atoms, molecules etc) are held together by electrical forces; and

■ the tendency for these forces to minimise the energy of the systems is limited by quantization which restricts the allowed configurations.

RS•C

The nuclear interactions are usually taken for granted in chemistry, and only studied in physics. The importance of quantum restrictions is not usually referred to in the teaching of either chemistry or physics topics until post-16 level. Although the electrical nature of interactions may well be discussed, research suggests that students do not always appreciate the nature of the electrical forces involved. This means that atomic structure is normally taught without reference to two of the three scientific principles on which it is based, and that the one key principle which is considered may not be emphasised strongly enough.

In the absence of a sound physical basis for understanding chemical structures, it is not surprising that learners often develop their own alternative ideas.

Learners' ideas about the atomic nucleus

The term 'nucleus' itself may sound quite similar to 'neutron' and this may be a source of confusion. More significantly, students will be familiar with the use of 'nucleus' in biology and may sometimes – hard as it may seem to appreciate – confuse atoms and cells. (It is reported that a significant minority of students may consider atoms to be alive, perhaps viewing them as something like amoeba.[1])

In one sense such a comparison is impressive: cells are sometimes considered to be the 'building blocks' of organisms, and atoms are often said to be the 'building blocks' of matter (even though this simplistic view is problematic, see Chapters 6 and 10). The cell-nucleus–atomic-nucleus analogy can be significant. The cell nucleus is often described as a type of 'control centre' for the cell, and the atomic nucleus may be understood to be a control centre for the atom. (This may contribute to the way that some learners see the force between nuclei and electrons to be unidirectional – from the nucleus, acting on the electrons.)

Making comparisons between different ideas is an important part of developing new concepts (see Chapter 2), but learners need to be taught to look for the negative as well as the positive aspects of an analogy. An example of this – seeing the atom as like a tiny solar system – will be discussed below.

If students appreciated the major role of electrical forces in maintaining atomic and molecular structures, then they might be expected to commonly ask how the nuclei – containing several (and sometimes many) positive charges are held together. Secondary students will not normally have considered the nature of nuclear forces, and might well expect the nucleus to be forced apart by the repulsion between the protons. That few students seem to spontaneously think of this problem seems to reflect the way that atoms are not usually conceptualised in electrical terms. This is unfortunate, as students are left without an appropriate way of thinking about the nature of chemical stability (see Chapter 6) and chemical reactions (see Chapter 9).

One suggestion to explain nuclear stability, mooted by post-16 level students, was that the nucleus was held together because of some influence from the electrons.[2] One student, Annie,[3] made such a suggestion in three different interviews months apart. During the first year of her two year course she suggested 'forces from the outer ring [sic]' were 'pushing' the neutrons and protons together. In a later interview she suggested that '[be]cause the nucleus pulls in the electrons, so [I don't know] if the electron forces actually help bind the nucleus, in any way' . At the end of her course she commented that 'obviously the electrons … may sort of control what's actually happening in the nucleus. Sort of … holding the neutrons and the protons together'

Another student, Carol mused about why a nucleus would be stable:

'you would think that a nucleus wouldn't be there really because, it's all protons and they repel, 'cause they're the same charge … but, there's another force, might be to do with electrons around the outside that holds it together … acting from outside.'

These comments reflect a common finding that students are often either ignorant of basic electrical principles, or at least do not transfer them from 'physics contexts', to apply them in the 'chemical context' of atoms and molecules. As one post-16 student taking college courses in both physics and chemistry explained:

RS•C

'I can't think about physics in chemistry, I have to think about chemical things in chemistry.'[4]

This 'compartmentalisation' of learning may well be partly responsible for some of the common alternative conceptions that students hold about atomic structure.

Learners' ideas about atomic structure

This lack of application of basic electrical ideas to the atom is reflected in the way students often conceptualise the way the electrons are held in position around the nucleus.

According to accepted scientific principles:

- all electrons in an atom are attracted to the nucleus;

- the force acting on an electron (due to the nucleus) depends upon the magnitude of the nuclear charge and the separation (the distance between the electron and the nucleus);

- the attractive force between an electron and the nucleus acts in both directions: both experience the same magnitude force; and

- each electron repels the others with a force which depends upon their separation.

To help simplify more complex atoms, we often introduce the idea of 'shielding' where inner shell electrons are considered to cancel the effect of an equivalent number of nuclear protons, so we can model the atom as a positive core charge and one shell of outer or valence electrons. This is only partially valid, as the 'electron shells' are not actually shells and interpenetrate – but it remains a useful approach. However, it is difficult for students to appreciate how the concept of shielding is supposed to work unless they accept the principles above.

It may seem that these points are the domain of physics rather than chemistry, yet these principles become quite important when students study chemistry at post-16 level, and are expected to explain such phenomena as patterns in ionisation energies. It is therefore significant that considerable numbers of students may well have alternative ideas about aspects of these interactions.

Interviews with post-16 students studying chemistry revealed the following alternative conceptions:

- the nucleus is not attracted by the electrons;

- the nucleus attracts an electron more than the electron attracts the nucleus;

- the protons in the nucleus attract one electron each; and

- the electrons repel the nucleus.

A classroom probe designed to elicit these, and related ideas, from post-16 level students is included in the companion volume. The **Ionisation energy** probe was originally used with 110 students in one college,[5] and was then piloted for the present publication with responses from 334 students in 17 different schools and colleges.[6]

It was found that just over a quarter of this sample of post-16 level students agreed with each of the statements that 'Each proton in the nucleus attracts one electron' and 'The nucleus is not attracted to the electrons.' Where students did think that the nucleus would be attracted to the electrons, they tended to agree most with statements suggesting the force on the electron would be larger than the force on the nucleus.

Half of these students agreed that 'Electrons do not fall into the nucleus as the force attracting the electrons towards the nucleus is balanced by the force repelling the nucleus from the electrons.'[7]

RS•C

Conservation of force: a common alternative conception

The interview study with post-16 students referred to above also found that students commonly held an alternative view of the way the nucleus held electrons in the atom. According to science the force between the nucleus and an electron depends upon the size of their charges and their separation.

Yet according to the alternative view the force was due entirely to the nucleus, and the size of the nuclear charge determined the total amount of force the nucleus could 'provide'. Force was seen as originating out of the nucleus towards the electrons. As the total nuclear force was fixed it would be shared by the electrons in the atom. This common alternative conception is known as the 'conservation of force' conception.[8]

Although this idea is incorrect it can be strongly held by students. In part this might relate to notions (referred to above) of the nucleus being the atom's control centre. However, perhaps the main reason for this was that the principle could be used to make correct predictions:

■ the larger the nucleus the more strongly electrons are attracted (true: helium has a higher ionisation energy than hydrogen);

■ the less electrons the more strongly they are each attracted (true: when an atom is ionised the second electron becomes more difficult to remove).

Although the reasoning is not quite correct, the use of this principle is reinforced when it seems to work successfully.

Several items in the **Ionisation energy** diagnostic instrument relate to this alternative conception, and have been found to be accepted by most students in the sample from 17 institutions:

■ 'If one electron was removed from the [sodium] atom the other electrons will each receive part of its attraction from the nucleus' – 55% of the sample agreed.

■ 'The third ionisation energy [of sodium] is greater than the second as there are less electrons in the shell to share the attraction from the nucleus' – 57% of the sample agreed.

■ 'After the [sodium] atom is ionised, it then requires more energy to remove a second electron because once the first electron is removed the remaining electrons receive an extra share of the attraction from the nucleus.' – 61% of the sample agreed.

Indeed, the clearest statement of this 'conservation of force' principle – 'The eleven protons in the [sodium] nucleus give rise to a certain amount of attractive force that is available to be shared between the electrons' – was judged to be true by 72% of these college level chemistry students, and was considered false by only 15%.

Learning by analogy – the example of the atom

In Chapter 4 the notion of learning impediments was discussed. Meaningful learning relies on the learner interpreting new information in the context of what they already know – so that it 'makes sense' to them. It was suggested that sometimes when students fail to learn the science that is presented to them, this may be due to understanding differently, when they relate the new material to alternative conceptions they already have. Much of the material in this publication is concerned with helping teachers explore students' alternative conceptions.

However, it was also suggested that sometimes students simply fail to make sense of teaching because they cannot relate what they are hearing and seeing to any existing knowledge. Where such 'null impediments' occur, teachers need to find ways to bridge between the new knowledge and what the learner does already know. Often the pre-requisite learning is in place, and the teacher simply needs to make the connections more explicit.

When the new ideas are too abstract to be directly related to existing ideas, teachers often call upon comparisons with other more familiar contexts. Atomic structure is clearly a topic which is highly abstract, as students are expected to learn about the internal structure of a conjectured entity which is

much too small to be directly experienced. We saw in Chapter 6 that many students had difficulty enough making sense of the molecular model of matter – the interactions within an atom are a further step from their everyday experience.

Not surprisingly analogies are often used when teaching about the structure of the atom. Various fruit placed at the centre of a large room may be used to give some feel for the scale of the nucleus within the atom. (Alternatively references to balls in various sports venues could be used – see Chapter 10.) A common comparison that is made is that 'an atom is like a tiny solar system'. The relationship between the nucleus and electrons is here modelled on the sun and planets. It has been suggested that although this may be useful for giving students an image of the atom, it is an approach that can go wrong.[9]

The use of this teaching analogy relies upon a number of assumptions;

1. that an atom is in some ways like a solar system;

2. that the students are familiar enough with the solar system to make use of the comparison; and

3. that the students will recognise in which ways the atom is like a solar system, and in which ways it is not.

None of these points are straightforward. The 'planetary' model of the atom is only of limited use once students move into post-16 courses, when they will be expected to see atoms in terms of orbitals (rather than orbits) and 'clouds' of electron density. In any case students may well already be familiar with the image of the planetary atom as it is a cultural icon that they will have seen in many contexts. Students' understanding of the solar system may not be orthodox, and they may not be clear about which features are to be transferred onto the atom.

These criticisms are not intended to suggest that such analogies should not be used. However, research does suggest that currently teachers do not always help students understand which features of an analogy or model do or do not match the target.[10] The significance of highlighting similarities and differences for forming new concepts was discussed in Chapter 2, and practice in such comparisons may be very useful to students.

One of the classroom activities included in the companion volume, **An analogy for the atom**, provides teachers with a chance to explore their students' use of this teaching analogy. This activity has two parts. The first worksheet, **The atom and the solar system**, provides a probe for eliciting students' ideas about the forces at work in these two systems.
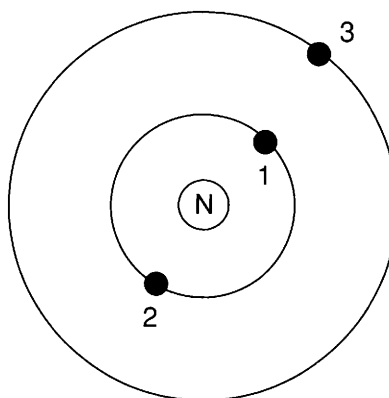

Figure 7.2 A simple representation of an atom

When this was piloted for the project, it was found that students often held alternative ideas about both the atom and the solar system. For example, one student in a class of 14–15 year olds who had studied atomic structure reported that the type of force attracting the electron towards the nucleus was a 'pull' force. An electron further from the nucleus (electron 3 in Figure 7.2) would be attracted by a stronger force as 'it is further away and therefore it will need a stronger force to draw the

RS•C

electron towards the centre'. He thought that there was no force acting on the nucleus due to an electron (as 'the electron is drawn to the nucleus instead of the nucleus [being] attracted [to] the electron'). He also thought there would be no force between the electrons as 'they are fixed on an axis and they have to have a fixed distance away from each other'. This student seemed to have quite firm ideas about atoms, albeit ideas at odds with the scientific model.

Other students in the same class suggested that the force attracting the electrons toward the nucleus was 'gravity', 'magnetic' or 'pole force'. Some classmates agreed that the nucleus would not be subjected to a force as 'the nucleus is the only thing that can apply a force'. Some students thought that there was a force attracting the nucleus towards the electrons, although it would be smaller (than the force attracting the electrons to the nucleus) as the 'nucleus is bigger + [has] bigger mass so bigger force'. Some students thought the electrons could not be interacting with each other, as they were interacting with the nucleus, or there was 'no relationship between them', whilst others acknowledged a gravitational 'reaction'. Some students did know electrons would repel each other, and one suggested 'they repel each other around the nucleus'.
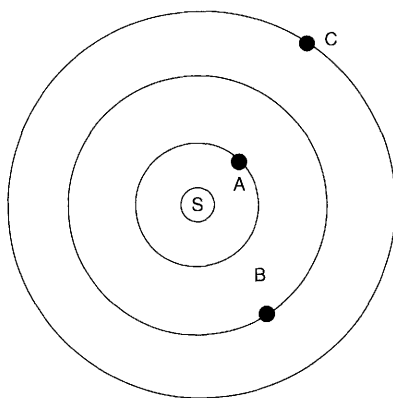

Figure 7.3 A simple representation of a solar system

Most of the students in this class recognised the role of gravity in attracting planets to the sun, although the 'pole force' also put in an appearance. Most of the group did not think the planets exerted a force on the sun, and about half thought there were no forces between planets. Some of the comments reflect the answers to the questions about the atom. The force acting on the planet with the largest orbit had to be greatest as 'it's further away so the force to stay with [the sun] is much bigger', 'the planets are attracted to the sun, not the other way round' and that 'planets [are] only attracted to the sun' and not each other.

Although the similarities in response suggested that learners might well see similarities between the two systems, the high proportion of alternative notions suggests that using the comparison as a teaching analogy could simply transfer incorrect ideas about one system to the other.

The second part of the exercise, **Comparing the atom with the solar system**, asks students to list the similarities and differences between the atomic system (Figure 7.2) and the solar system (Figure 7.3). To some extent this activity is scaffolded (see Chapter 5), as the questions asked on the **The atom and the solar system** worksheet provide some cues for suitable comparisons. The first worksheet can be seen as organising the students' existing knowledge to prepare them for the later task. It acts as a scaffolding PLANK, providing a conceptual platform for developing new knowledge.

When this exercise was piloted for the project it was found that some students found it very difficult to suggest more than a couple of similarities or differences (and many did not make the 'obvious' point about the atom being a good deal smaller than a solar system). This may suggest that this is a skill which needs more explicit practice (see the exercises on **Chemical comparisons** discussed in

RS•C

Chapter 2). The class of 14–15 year olds which produced the responses discussed above did have a fair attempt at spotting similarities and differences. Some good suggestions were made for both the similarities, and the differences:

Similarities:

■ Both the atom and the solar system have centres that attract the surrounding planets or electrons

■ Both have forces involved

■ They both rotate around a centre point

Differences:

■ More than one thing on the ring in atoms

■ Planets have no charge but electrons are negatively charged

■ Electrons have virtually no mass and planets have a large mass

■ The solar system is a lot bigger

■ Different forces

■ Planets can be seen with the naked eye, electrons can't

■ The planets rotate around

However, unsurprisingly in view of the alternative notions revealed in the first part of the activity, some of the points of similarity and difference suggested were:

Similarities:

■ They are both [electrons, planets] held in orbit by gravity

■ They [electrons, planets] are not attracted to each other

■ They both have energy sources in the centre

■ Neither the nucleus or the sun are attracted to the planets or electrons

Differences:

■ The planets move around [implying that the electrons do not]

■ The rings are closer together around the sun

■ The force attracting the particles is pull but the force attracting the planets is gravity [implying gravitational force is something other than a pull]

■ There is force between [the planets – implying no force between electrons]

Learners' ideas about orbitals

Post-16 level students are often expect to move beyond ideas about electron shells, to learn something about electronic orbitals. Some observers feel that orbital ideas are unhelpful prior to university level study. It has been suggested that the notion of 'electron pair domains' is simpler, and sufficient for school and college level study.[11] However, examination stipulations may require learners to tackle orbital concepts.

This has found to be a topic that students often find difficult.[12] This should not be surprising because orbital ideas are highly abstract, and so students may find difficulty making sense of them in terms of existing knowledge (ie there may be a 'null learning impediment' – see Chapter 4). Where students have a naive appreciation of the roles of models in science (see Chapter 6), they may well be committed to the idea of electron shells, and this existing school science knowledge may interfere with the intended learning (ie there may be a 'pedagogic learning block' – see Chapter 4).

RS•C

Students often come to this topic with an image of the electron shell (usually represented as a circle around the nucleus) as a kind of electron orbit, and so the term 'orbital' may initially be acquired as an alternative term with much the same meaning.[13] When the idea of sub-shells is introduced, this may again become confused with shells, orbits and orbitals. Further confusion is likely as other related concepts are introduced. So orbitals may be confused with energy levels, and conventional diagrams representing orbital 'probability envelopes' may be read as showing orbital boundaries.

Until students have managed to differentiate between some of these related, but distinct, concepts, the conventional labelling for atomic orbitals (1s, 2s, 2p, 3s, ...) is meaningless, and unlikely to be mastered.

Such abstract ideas are demanding even for able students, and time, reinforcement and practice are needed if learners are to show (and maintain) a good understanding of the orbital topic. Unfortunately, the pressures of covering the course material often mean that before this can occur, students have already been introduced to further complications: hydbridization and molecular orbitals. The latter may be σ or π, and may be bonding or anti-bonding. The folly of expecting some students to make sense of these ideas after a limited exposure is reflected in the way one student in New Zealand defined anti-bonding orbitals as 'silly things' that just stuck out![14]

One particular problem that some students have is recognising that rehybridization of atomic orbitals (a formal mathematical process used in trying to understand nature, not a process in nature itself!) gives a set of atomic orbitals – some of which will no longer be present once a molecule is formed. It may help us to understand the bonding to think of a carbon atomic system undergoing rehybridisation to give sp^3 atomic orbitals suitable for overlap. There are no electrons in a methane molecule, however, best described as being in sp^3 hybrid orbitals – those electrons are now in σ-molecular orbitals.[15]

In a molecule as simple as oxygen (O_2) there will be orbitals that are effectively atomic orbitals unchanged from those in a ground state atom, orbitals which are effectively hybridised atomic orbitals, and molecular orbitals which are unlike any of the precursor atomic orbitals. The level of treatment expected in post-16 chemistry is likely to consider how rehybridisation allows two partially occupied orbitals on each atom to overlap to form molecular orbitals. The overlap of sp^2 hybrids allows the formation of the σ-bond, and the overlap of unhybridized p orbitals allows the formation of the π-bond.
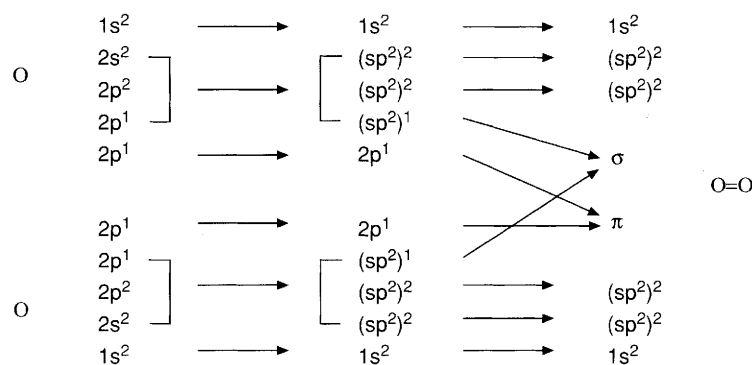

Figure 7.4 A scheme to show the orbitals in a simple molecule

Figure 7.4 shows the (hypothetical) process of moving from ground state atomic orbitals on two oxygen atoms, through rehybridised atomic orbitals, to the orbitals expected in the double bonded molecule. Even in such a relatively straightforward example as a diatomic molecule of an element, it is little surprise that such a scheme proves confusing to many students. It is therefore rather ironic that oxygen is found to be paramagnetic – and must therefore have unpaired electrons in its ground state. The model used at post-16 level does not even predict the correct electronic structure in this case.[16]

RS•C

At this level radicals are considered to always be highly reactive species, the diradical nature of oxygen[17] is not normally discussed, and students assume that electrons are paired up in the oxygen molecule.

Depicting molecular structure

In this publication a deliberate attempt has been made to use a variety of ways of representing molecules and other structures. (For example, in Chapter 5, when suggestions are made about introducing hydrogen bonding through the structure of the water molecules, the example worksheet included three distinct representation of the molecules.) This is because all of our diagrams are limited ways of representing various aspects of our mental models of molecules. The importance of models in chemistry, and the limited way in which they are appreciated by students, was considered in Chapter 6. Whenever we chose to use a particular type of diagram we are (consciously or not) emphasising certain aspects of our models of molecules. Often some aspects of the diagram are irrelevant or even unhelpful, and it may not always be clear to students which aspects of a diagram (or a solid model or computer animated model) they are meant to be attending to, and which aspects are less relevant in a particular context. It has been found that students' ability to solve chemical problems is often related to their ability to interpret different representations of chemical systems.[18]

It is suggested that teachers should both use a variety of diagrams and other models, and make a habit of reiterating which aspects of the model are and are not significant in a particular context. Reference was made in Chapter 2 to some research where students were asked to make discriminations between different chemical species (molecules, ions, atoms etc) represented by pictures from textbooks. It was found that some of the students largely used criteria which were based purely on the way the species were drawn, rather than their chemical attributes. That is, some learners 'seemed to discriminate between [chemical species] on the basis of the way they were represented, rather than what was represented'.[19] Where some students focused on features of the chemical species themselves, others would comment upon 'the different conventions used in chemistry textbook diagrams to represent various aspects of the species drawn' such as whether electrons were shown as 'e' or '•'.

Noticing such different conventions does not always equate to considering them significant, but does remind us that aspects of diagrams that have effectively become invisible to 'experts' due to familiarity may draw the attention of relative 'novices'. Such aspects may act as distractions, and take up some of the limited 'slots' in the students' 'mental scratch pad' (see Chapter 5). This is another example of why it is important for the teacher to learn to see the material presented at the 'resolution' available to the learner.
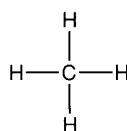
Consider the following diagrams:

$$\begin{array}{c} H \\ | \\ H-C-H \\ | \\ H \end{array}$$

Figure 7.5 A representation of a methane molecule

$$\begin{array}{c} H \\ H \ \overset{\times}{\underset{\times}{C}} \ H \\ H \end{array}$$
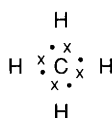
Figure 7.6 A second representation of a methane molecule
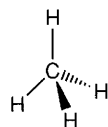
RS•C

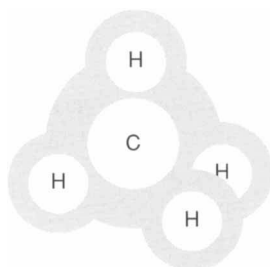**Figure 7.7 A third representation of a methane molecule**

Figure 7.8 A fourth representation of a methane molecule

The four diagrams in Figures 7.5–7.8 are different ways of representing a methane molecule. In Figure 7.5 the bonds are shown as lines. This is a type of diagram that students can easily learn to draw, but it may not always be the most appropriate diagram to use. Younger students are known to sometimes think of chemical bonds as being physical connections between atoms – and to think of bonds as (and not just analogous to) sticks, springs or adhesive.[20,21] This type of diagram could reinforce such a view.

Figure 7.6 actually could help to avoid this as it represents the bonding electrons. The dot and cross formalism is very common and is meant to help students with their electron-accounting by showing where the electrons (are conjectured to have) originated. However, this type of figure may encourage common alternative conceptions that electrons from different atoms are different and are only (or are more) attracted to their own nucleus, and will always return to the original atom when the bond breaks. Another source of confusion is that the same dot and cross symbols may be used in a distinct convention to indicate the two different spin states of electrons rather than their atomic origins.[22]

Figures 7.5 and 7.6 may both give the impression that molecules are flat, whereas Figures 7.7 and 7.8 should avoid this interpretation by representing the three dimensional shape of the molecule. Figure 7.7 again uses lines for bonds, where Figure 7.8 gives an impression of the electron cloud. This type of diagram is a little more sophisticated, but gives a better impression of the way bonding electrons are part of molecules (rather than belonging to specific atoms) and the way molecules have quite 'fuzzy edges' rather than being like marbles with a distinct edge.
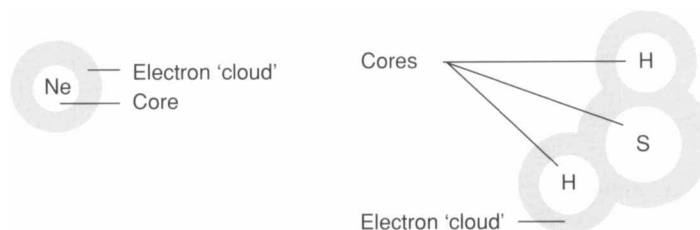

Figure 7.9 Representing molecules with electron clouds

This type of representation (see Figure 7.9) has been used quite a lot in this publication, and is used in some of the materials intended for 11–14 year old students. It might be objected that the idea of an electron cloud is far too abstract for students at this level, but all of these representations are abstract, and pictures with electron 'clouds' need be no more problematic than diagrams showing concentric electron 'shells', or representing bonds as lines or as dot and cross pairs. While students just beginning to learn about particle theory are not ready to fully appreciate what is meant by an atomic core or an electron cloud, these types of diagrams did not seem to be problematic for the students. It

could be argued that showing electron clouds as shading is actually less abstract than the use of 'C' to represent the carbon nucleus and the K electron shell in Figure 7.5 – something we routinely expect pupils to accept.
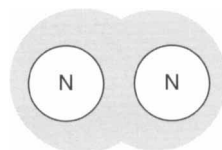
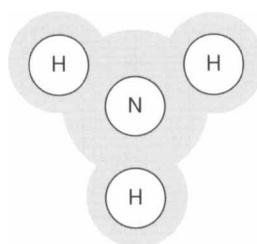**Figure 7.10 A representation of a nitrogen molecule**

Figure 7.11 A representation of the pyramidal ammonia molecule
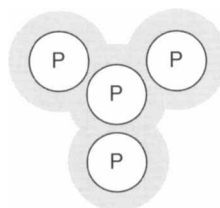
Figure 7.12 A representation of the tetrahedral phosphorus molecule

However, I am certainly not suggesting that this type of diagram is a panacea. Figure 7.10 does not clearly indicate the multiple nature of the nitrogen bond, and Figures 7.11 and 7.12 do not discriminate between the two different arrangements present. The type of diagram we select to use at any particular time should reflect the specific molecular features we are trying to emphasise.
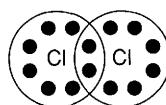
Figure 7.13 A representation of a chlorine molecule

A diagram which is effective at making some points, may be less useful on other occasions. We also need to be very careful in checking what students read into diagrams. Figure 7.13 shows a common type of diagram used in chemistry books. The overlap between atomic shells is meant to indicate the covalent bond, with the bonding electron pair within the overlap. One college student suggested that the bonding electrons in such a diagram were more restricted than the non-bonding (lone pair) electrons as they could only move within the area of overlap, whereas the other electrons shown could move throughout the shell.[23]

As this student started to learn a more sophisticated model of the electronic structures of atoms and molecules, she over-interpreted the more simplistic model of a molecule so familiar from her school science. This is not an argument for not using such diagrams, but as with teaching analogies (see earlier in this chapter), and all other forms of teaching models (see Chapter 6), we need to help students see which features are important, and which are just conventions of graphic artists. Selecting different types of diagrams of the same species, to emphasis different teaching points, can help our

RS•C

students realise which aspects of these representations are actually meant to reflect the abstract features of the species chemists represent.

Learning about lattice structures

It has been found that when students commence university study of chemistry they may often have a very limited understanding of lattice structures.[24] This is unfortunate as a key objective of learning about chemical structure is to be able to explain the properties of materials – hardness, electrical conductivity, cleavage planes in salts, malleability etc – and this requires an understanding of both chemical bonding (see the next chapter) and larger scale structure. One common problem is that, having learnt about molecules, students often assume that all materials are molecular.[25]

In ionic materials, the formula may seem to imply molecules. So NaCl is often thought to comprise of 'NaCl' molecules arranged in a lattice (see Chapter 8). Students who have learnt to use the idea of valency or combining power (H=1, N=3, C=4 etc) to work out the stoichiometry of covalent molecules (NH_3, CH_4 etc) will often extend this idea to the ionic case using electrovalencies (Mg=+2, F=−1, so magnesium fluoride will be thought to comprise of $Mg^{2+}(F^-)_2$ molecules). The diagrams used in many student textbooks are likely to encourage this false interpretation (see Chapter 10).

Similarly, in metals, the student may assume that the electronic configuration of the metal atom determines how many other metal atoms it bonds with, and therefore how many metal atoms are present in the 'molecules'.

In the case of substances with covalent bonding students may become confused as they need to distinguish between molecular materials bound by van der Waals forces, and giant covalent structures. Students sometimes suggest that covalent bonds are relatively weak because (for example) sulfur is readily melted and sugar easily dissolved. They often fail to realise that the intramolecular bonds may be unaffected in such processes.

Although there are only a limited number of familiar substances which have giant covalent lattices, this is an important type of structure. Learners may consider this type of material to contain discrete molecules with strong intermolecular forces, or may even consider solid carbon to comprise of discrete atoms, something that may in part derive from the 'molecular' formula of carbon being commonly given as 'C' and taken to imply 'C_1', when 'C_∞' might be more appropriate.

One of the classroom resources included in the companion volume, **The melting temperature of carbon**, is designed to challenge this alternative conception, and reinforce the relationship between melting temperature and structure. This study activity has two parts: **Predicting the melting temperature of carbon** and **Explaining the melting temperature of carbon**.

Predicting the melting temperature of carbon presents students with a table of the melting temperatures and relative molecular masses for a number of non-metallic elements (reproduced as Table 7.1). Although carbon is included in the table, its melting temperature is omitted. Students are asked to spot the pattern in the data provided (ie that melting temperature increases with increasing relative molecular mass) and to predict the melting temperature of carbon.

RS•C

Element	M_r	Melting temperature/K
Helium, He	4.0	4
Carbon, C	12.0	
Neon, Ne	20.2	25
Fluorine, F_2	38.0	53
Chlorine, Cl_2	71.0	172
Bromine, Br_2	159.8	266
Iodine, I_2	253.8	387
Sulfur, S_8	256.8	392

Table 7.1 Comparing molecular mass and melting temperature

When the exercise was piloted a typical value predicted by many students was about 14 K. If students recall that carbon (graphite or diamond) has a high melting temperature, this will already present them with a puzzle to be solved.

Some students will not be aware of the puzzle, either not having learnt, or not bringing to mind, the fact that carbon has a high melting temperature. They are told that carbon has a melting temperature of 3823 K at the beginning of **Explaining the melting temperature of carbon**, so that they will all be aware of the anomaly.

The students are then asked to consider diagrams showing discrete molecules of neon, chlorine and sulfur alongside a similar style of diagram showing part of a 2-dimensional representation of a diamond lattice. The first three diagrams show discrete species (eg see Figure 7.13). The figure for carbon is clearly different, showing only part of a more extensive structure (see Figure 7.14).
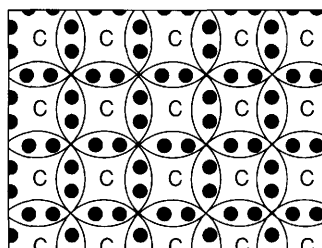

Figure 7.14 Part of the giant covalent structure of carbon

The questions in this part of the exercise ask students if the diagrams help them explain why carbon has a high melting temperature.

A school student in a class of 14–15 year olds predicted that carbon would melt at about 13 K as the 'molecular mass of carbon is roughly half that of neon so the melting temperature is roughly half', and then explained the high melting temperature from the diagram, as 'every single carbon atom is bonded to 4 others carbon atoms meaning it takes a lot of energy to break the bonds'.

One post-16 student who predicted that the melting temperature of carbon would be 14 K because 'its molecular structure is a single atom...', later – after considering the diagrams – explained that 'you can see that the atoms are densely packed together with one carbon atom attached covalently to four [other] carbon atoms...so it means a lot of atoms bonded together'. Another student in the same post-16 class who predicted a melting temperature of 13.5 K explained that 'the C atoms carry on continuously bonding to other C atoms and the molecule does not stop' and that it was 'difficult to break the bonds and therefore difficult to melt'.

RS•C

Challenging students' expectations

This exercise is certainly contrived, being designed to encourage students to make an incorrect prediction, which then needs to be explained away. It also leads them to the type of explanation needed through the diagrams provided. For those students who are well aware of the high melting temperature of carbon, and the reasons for it, it will seem an artificial task, but provides useful practice in a number of areas (calculating molecular masses; finding patterns in data; explaining properties in terms of structure).

For those students who do not make a clear distinction between substances with covalent bonding which are simple molecular, and those which have giant lattices, the task presents a genuine anomaly.

Evidence from historical studies suggest that it is the anomalies between theoretical predictions and actual observations which often act as the incentive for conceptual revolutions. In a similar way, conceptual change in learners is believed to be encouraged by making them see how their ideas are inconsistent with actual observations (see Chapter 10). It is said that conceptual change may be brought about by 'dis-equilibrium' or 'cognitive dissonance' (where new information does not match existing ideas). As rational beings, we try to find ways to make our ideas more coherent.[26]

In laboratory situations phenomena are often complex, and in practice students often 'see' what they expect. By providing a data based exercise **The melting temperature of carbon** provides a definite anomaly that needs to be explained.

Notes and references for Chapter 7

1. A. K. Griffiths & K. R. Preston, Grade-12 students' misconceptions relating to fundamental characteristics of atoms and molecules, *Journal of Research in Science Teaching*, 1992, **29** (6), 611–628.

2. K. S. Taber, Understanding Chemical Bonding – the development of A level students' understanding of the concept of chemical bonding, Ph.D. thesis, University of Surrey, 1997.

3. An assumed name. As is standard practice, the names given to individual students discussed in the text have been changed to protect anonymity.

4. K. S. Taber, The sharing-out of nuclear attraction: or I can't think about physics in Chemistry, *International Journal of Science Education*, 1998, **20** (8), 1001–1014.

5. K. S. Taber, Ideas about ionisation energy: a diagnostic instrument, *School Science Review*, 1999, **81** (295), 97–104.

6. Details are available at the project website: **http://www.egroups.co.uk/group/challenging-chemical-misconceptions** (accessed October 2001).

7. This statement contains two major errors. Firstly it is implied that forces acting on two separate bodies can somehow cancel each other. Secondly, the positive nucleus is said to be repelled by the negative electrons.

8. The label 'conservation of force' refers to how a set amount of force is considered to 'emanate' from a nucleus of a particular charge regardless of the number and arrangement of electrons it is bound to. Of course force is not conserved in this way, although both charge and energy are conserved. There is some evidence from interviews that students do sometimes confuse the charge, with the force acting due to the charge; or the force with the work done by the force. (See note 2).

9. K. S. Taber, When the analogy breaks down: modelling the atom on the solar system, *Physics Education*, 2001, **36** (3), 222-226.

10. S. Sizmur & J. Ashby, *Introducing Scientific Concepts to Children*, Slough: National Foundation for Educational Research, 1997.

11. R. J. Gillespie, Bonding without orbitals, *Education in Chemistry*, 1996, **33** (4), 103–106.

12. K. S. Taber, Building the structural concepts of chemistry: some considerations from educational research, *Chemistry Education: Research and Practice in Europe*, 2001, **2** (2), 123–158, available at **http://www.uoi.gr/conf_sem/cerapie/** (accessed October 2001).

13. This could be another example of what Prof. Hans-Jürgen Schmidt calls a label acting as a hidden persuader – see Chapter 1.

14. R. K. Coll and N. Taylor, Alternative conceptions of chemical bonding for upper secondary and tertiary students, *Research in Science and Technological Education* **19 (2)**, 171–191.

15. The habit of describing electrons in molecules as though they are still in atomic orbitals may be related to the tendency to consider bonding electrons as still belonging to specific atoms.

16. The molecular structure of oxygen is better understood in terms of the nitrogen molecule, formed by overlap of sp hybrids and unhybridized p orbitals in two planes. In nitrogen this gives a triple bond (with a s-component and two p-components) and one non-bonding pair of electrons in sp-hybrids on each nitrogen atomic centre. In oxygen there are two more electrons to be accommodated, which are placed in degenerate p-anti-bonding orbitals. Although there is a triple bond structure, the overall bond order is 2 due to the anti-bonding orbitals. See N. W. Alcock, *Bonding and Structure: Structural Principles in Inorganic and Organic Chemistry*, Chichester, Ellis Horwood Limited, 1990.

17. J. Barrett, *Understanding Inorganic Chemistry: the Underlying Physical Principles*, Chichester, Ellis Horwood Limited, 1991.

18. G. M. Bodner & D. S. Domin, The role of representations in problem solving in chemistry, paper presented to the *New initiatives in chemical education* on-line symposium, available at **http://www.inform.umd.edu/EdRes/Topic/Chemistry/ChemConference/ChemConf96/Bodner/Paper2.html** (accessed October 2001)

19. K. S. Taber, Can Kelly's triads be used to elicit aspects of chemistry students' conceptual frameworks?, 1994, available via Education-line, at **http://www.leeds.ac.uk/educol/** (accessed October 2001).

20. T. Wightman, P. Green & P. Scott, T*he Construction of Meaning and Conceptual Change in Classroom Settings: Case Studies on the Particulate Nature of Matter*, Leeds: Centre for Studies in Science and Mathematics Education, 1986.

21. A. G. Harrison & D. F. Treagust, Learning about atoms, molecules, and chemical bonds: a case study of multiple-model use in grade 11 chemistry, S*cience Education*, 2000, **84**, 352–381.

22. A. J. Catterall, Electron structure in molecules: a new approach, *School Science Review*, 1967, **48** (166), 798–802.

23. K. S. Taber, Building the structural concepts of chemistry: some considerations from educational research, *Chemistry Education: Research and Practice in Europe*, 2001, **2** (2), 123–158, available at **http://www.uoi.gr/conf_sem/cerapie/** (accessed October 2001).

24. D. Cros, R. Amouroux, M. Chastrette, M. Fayol, J. Leber & M. Maurin, Conceptions of first year university students of the constitution of matter and the notions of acids and bases, *European Journal of Science Education*, 1986, **8** (3), 305–313.

25. P. G. Nelson, To be a molecule, or not to be?, *Education in Chemistry*, 1996, **33** (5), 129–130.

26. Whilst it is undoubtedly true that we are driven to make sense of the world, and that coherence is an important criterion, we are also capable of holding apparently inconsistent views. Although being rational has its uses, we live in a world where information may be partial, incorrect and incoherent: so we have evolved brains that can cope with this muddled input. One consequence is the ability to hold manifold conceptions, or multiple conceptual frameworks, for the same topic – see Chapter 8.

RS•C

This page has been intentionally left blank.

RS•C