# Teaching Secondary Chemistry

*Chapter 4*

# Developing models of chemical bonding

---

**4.1 From holding power to chemical bonding**

↓

**4.2 Principles underpinning the bonding concept**

↓

**4.3 Modelling varieties of chemical bonding**

↓

**4.4 Relating type of bonding to changes in chemistry**

---
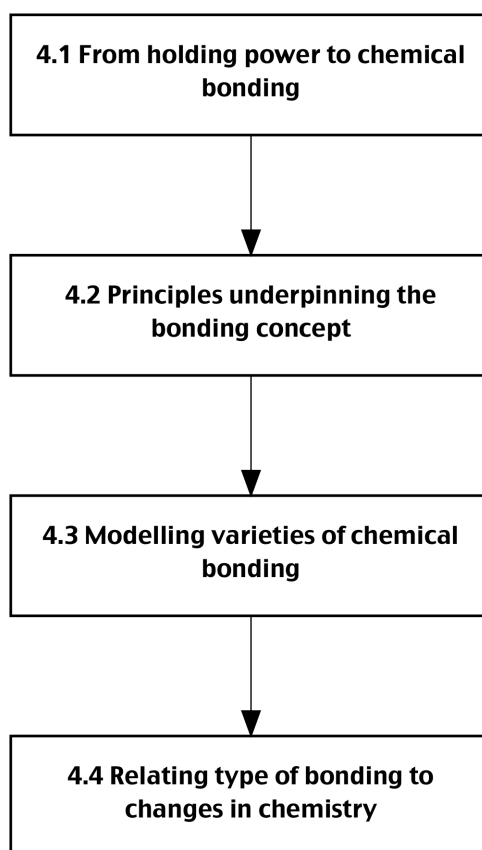
Chemical bonding is a theoretical idea that is a key part of the way matter is modelled by chemists at submicroscopic scales (as composed of extremely tiny 'quanticles', as described in Chapter 1) and so is not likely to be familiar to students from their experiences outside the classroom. However, if particle models are introduced in the manner recommended in this book (Chapter 2), students will have been introduced to

the idea that the particles of which substances are made have 'holding power' of varying strengths, which tends to cause them to clump together unless they have sufficient movement to overcome this.

## Choosing a route

The ideas presented in this chapter, then, build upon prior learning that chemists think of matter as composed myriad tiny little parts (usually called particles, but - as suggested in earlier chapters - this can be a misleading term), which have a holding power that leads to them tending to stick together. The topic of chemical bonding should therefore be seen as part of a learning progression from introductory particle theory.

One route into thinking about chemical bonding is in terms of the states of matter; the (submicroscopic) structures different solid substances take; and the properties of different substances. As Chapter 2 suggests, different substances undergo phase changes (solid to liquid; liquid to gas) at very different temperatures. We also find that substances have different behaviours if we try to dissolve them in solvents, or pass electricity through them, for example. These differences in behaviour present the chemist with phenomena to be explained, and provide the chemistry teacher with one basis for exploring the nature of the 'holding power' students learn about early in their secondary chemistry.

A second route to approach the idea of chemical bonding is through modelling chemical reactions. In the previous chapter (Chapter 3), ideas about modelling chemical reactions in teaching were discussed, and it was suggested that understanding the way chemists conceptualise chemical change requires students to think about what is happening during reactions at the level of the 'quanticles' involved: the molecules and ions. A simplified model presenting chemical reactions as involving the rearrangement of 'atoms' can offer a starting point for thinking about chemical change at the scale of atoms, but raises the question of why atoms seem to come 'unstuck' from each other during reactions, and then stick together in new combinations. This provides a second motivation for developing theories of chemical bonding.

Chemical bonding is a challenging topic, because it relates to theoretical models about entities much too small to be seen. Despite this, because of its importance in making sense of chemical phenomena, it has long been recognised as a core concept area in chemistry. The abstract nature of the topic no doubt contributes to its challenge as part of a secondary curriculum, and students commonly demonstrate serious misconceptions. The teaching sequence suggested here is informed by research into student learning difficulties. It is especially important that we emphasise something of the electrical nature of bonding interactions before students can develop false ideas about the octet rule being a 'cause' for bond formation (see the previous chapter).

Whilst the subject of this chapter is one of core theoretical ideas of chemistry, there is room for discussion of how much detail different groups of students should be taught. Many secondary students will find this material dull and difficult, and it may be that the curriculum for some groups will not prescribe much discussion of the material covered here. Conversely, some groups of students will find this material fascinating (see Chapter 12), and a full exposition at secondary level provides a very useful background for those who may go onto to study post-secondary chemistry courses. Where students are required to study chemical bonding in any depth, I would strongly recommend that the approaches offered here should be adopted. In particular, understanding bonding as primarily the outcome of the same physical forces studied in secondary physics is educationally and scientifically much more valid than allowing students to develop ideas of chemical bonding as related to mystical notions of atoms actively seeking to full their electron shells or to acquire mystical numbers of electrons.

## 4.1 From holding power to chemical bonding

As chemical bonding is an abstract, theoretical, idea, it is important that students are motivated to learn about the topic. After all, models of different kinds of chemical bonding must seem obscure and largely irrelevant to most learners. This first section of the chapter then builds upon earlier chapters by setting out what it is that the chemical bonding concept can help us understand and explain. Providing some

background of this type is important for all students, but some groups will need more persuading than others.

## Properties of substances

It is useful to remind students that samples of substances can exist as solid, liquid or gas depending upon the conditions (see Chapter 2) – and in particular temperature. This could be a useful context for setting up a data handling exercise. For example, Table 4.1 presents the melting and boiling temperatures of the group 2 ('alkaline earth') and group 7 ('halogen') elements in periods 2-5 of the periodic table (see Chapter 1).

| Period | Group 2 element | Melting temperature/K | Boiling temperature/K | Group 7 element | Melting temperature/K | Boiling temperature/K |
|--------|-----------------|-----------------------|-----------------------|-----------------|-----------------------|-----------------------|
| 2 | Beryllium | 1556 | 2750 | Fluorine | 53 | 85 |
| 3 | Magnesium | 923 | 1390 | Chlorine | 172 | 239 |
| 4 | Calcium | 1123 | 1765 | Bromine | 266 | 331 |
| 5 | Strontium | 1043 | 1640 | Iodine | 387 | 456 |

Table 4.1: Melting and boiling temperatures for some elements

There are various ways this data could be displayed, including the very useful form of diagrams for comparing the temperatures of phase changes of different substances introduced in Chapter 2 (see figure 2.6), but drawn to a specific scale. For some groups of students, asking them to produce a set of these figures to compare between elements will be a suitably challenging task (depending upon the time available and the speed at which the students work, they could be asked to produce a set of figures

for one or other group, or a pair to compare the two elements in the same period). A key point to check understanding is to ask students which state these different elements are found in at room temperature, and how readily they can be frozen/melted etc. A more capable group of students can be asked to find alternative ways to display the information shown in Table 4.1 (e.g. line graphs showing tends, bar charts, etc), and to offer a commentary on the relative strengths and weaknesses of different ways of presenting the data.

Having reminded students about the different states in which substances are found, and how this depends upon their melting and boiling temperatures, another useful actively focuses on solids. Students can be presented with a sample of different solids in sealed test tubes, and asked to suggest how these different solid samples of substances may have different properties.

A suitable selection might include:

Sulfur power

Graphite rod

Copper turnings

Hydrocarbon wax shavings

Iron filings

Magnesium ribbon

Zinc sheet

Copper sulfate crystals

A piece of glass rod

Some salt crystals

Polythene beads

It is best to avoid composite materials (such as wood), and to just use samples that can be considered to be particular substances. This activity can be set up as group discussion work, with students exploring and pooling their ideas in small groups. Ask the students not only to base suggestions on their observations, but also their background knowledge of these (and other) substances. After providing a period for discussion, ask each group to share suggestions with the class, and list these on the board. All sensible suggestions should be considered acceptable.

After compiling suggestions, highlight to the students those features of particular interest to chemists. One important teaching point is that samples of the same substance may take different forms. So although graphite was presented as sticks, it is also available as a powder. Similarly, iron filings have a relatively small particle size, but single discrete blocks of iron of considerable size are possible. So some material properties are not inherent in a substance, but are contingent on how a particular sample is prepared.

Some relevant ideas that students might present (and which you may wish to point out otherwise) are:

- Some substances are translucent or transparent, where others are opaque;

- Some substances have (various) colours;

- Some substances are shiny, whilst others are dull;

- Substances have different densities at the same temperature;

- Some substances are obviously crystalline (regular shaped solids with flat faces);

- Substances will dissolve in water to different extents;

- Substances will dissolve in organic solvents (e.g. a vegetable oil) to different extents;

- Some, but not all, substances often feel cold to the touch;

- Some substances are good conductors of electricity, whereas others are insulators;

- Some substances are attracted to magnets, but most are not.

- Some substances are hard, and some soft;

- Some substances are elastic (they deform and then return to their original shape) and others plastic (they deform, and keep the new shape) or brittle (they readily snap);

- As pointed out earlier, some solids will melt at much lower temperatures than others.

There may be scope here for including some class practical work, or some demonstration practicals, to illustrate some of these properties, if time allows. These types of properties are often referred to as 'physical properties' to contrast them with 'chemical properties' (the profile of chemical reactions different substances undergo).

## Structure of materials

As a science, chemistry seeks to find explanations for chemical phenomena. So chemists are interested in the reasons why different substances have such different properties. This leads chemists to try and find out about the structure of different substances at the submicroscopic scale: how the component quanticles are arranged into structures. However, the scale of this structure is far too small to be investigated with ordinary light microscopes, so chemists have adopted a wide range of special techniques to investigate structure at this level.

As these techniques are generally too advanced to be explained in a secondary course, it is probably sensible to present them largely as 'black box' tools that student are not expected to understand – similar to the approach taken in Chapter 1 to the chemist's special 'elemental analyser' (a mass spectrometer). A variety of modern chemical analytical techniques have been used to explore structures, including X-ray diffraction and various forms of spectroscopy. These techniques depend upon physics (of waves) that some students will study at upper secondary level, and finding out about these

techniques might be a suitable extension or homework activity for some (but probably not most) students. These techniques provide information that tells chemists about the special arrangements of the atomic centres in different substances – the geometry of the structures, and how close together the atomic cores are.

Crystal structures have very regular patterns of atomic centres. Although students are not normally expected to know details of crystal structures at secondary level, many students will be interested to see models of common structures. A model of NaCl structure will show pupils that the sodium and chlorine atomic cores have a very regular arrangement, which builds up into a cubic pattern. A model of diamond structure will show an extensive regular arrangement of the same (carbon) unit with a tetrahedral geometry. Common metallic structures are based around each atomic core having 8 or 12 near neighbours.

A very important finding from structural studies in that sometimes a solid has an extensive arrangement of atomic cores, all held closely to all their neighbours, whereas sometimes there is a different type of structure with small numbers of atomic cores closely bound together – but with these small units being more loosely bound with others like themselves. This suggests that the idea of the 'holding power' (see Chapter 2) that holds particles together into structures, is not straightforward.

Ideas and evidence about chemical structures

Not only are analytical techniques that reveal the structures of substances at the submicroscopic scale complex, but the data they produce often needs considerable interpretation. This may be another good reason to consider these techniques as 'black boxes' – as tools chemists use, but which we do not need to understand at secondary level.

However, with some groups of students (see Chapter 12), it might be appropriate to look at some examples of the output of some of these techniques: such as infra-red (IR); (ultra violet) UV-visible; NMR (nuclear magnetic resonance) spectra; and X-ray diffraction patterns. Many examples can be found on the world-wide-web. What all of these techniques have in common is an output that is quite *different* from the

structures being elucidated, offering indirect information. That is, there is a major task of interpreting features of different diffraction patterns or spectra to identify the information needed to specify the structures themselves (the film 'Life Story' offers an excellent dramatization of this – see the list of 'other resources' at the end of the chapter). This makes the important teaching point that often there is no simple path from the data obtained in scientific experiments to an obvious and clear conclusion, and that often the development of new techniques not only involves building the machines - the physical apparatus - but also developing (constructing, testing, confirming) a mathematical and analytical apparatus that allows us to draw inferences from the data.

Whilst this offers students a confusing and complex image of the nature of chemistry, it is an image that is more realistic than is likely to be gained from many school chemistry practicals that have been set up to give readily interpreted results. This is important, as a key aspect of scientific literacy is appreciating why there is often intense debate between scientists about the interpretation of evidence (e.g. the level of risk in using nuclear power; the extent to which there is human-caused global warming), and why accepted scientific ideas are sometimes found to be wrong. One of the core aims of science education, including chemistry education, should be to allow students to come to appreciate science as a reliable source of knowledge based on logical analysis of evidence, whilst also recognising that scientific knowledge is never absolute and beyond revision.

## Explaining holding power

Chapter 2 highlighted the way in which a sample of a substance could be solid, liquid or gas depending upon the 'holding power' of the 'particles' and on how fast they are moving: the higher the temperature of a sample, the faster the movement of its component 'particles', and the less effective the 'holding power' at holding them together.

Bonding is the term used in chemistry as a formal label for the holding power. In the previous section it was suggested that structural studies report that this holding power

or bonding does not always act in the same way – so sometimes it seems to produce organised arrangements that keep particles very close together, and sometimes it leads to more haphazard, more weakly bound, clumps. Chemists therefore seek an explanation not only for the origin of the holding power, but also for these different types of effects. It transpires that chemical bonding is a very complex topic, but that much can be explained by the simple principle that electrical charges attract and repel each other:

| | positive charges (such as on atomic cores, cations, nuclei) | negative charges (such as on electrons, anions) |
|---|---|---|
| positive charges (such as on atomic cores, cations, nuclei) | repel | attract |
| negative charges (such as on electrons, anions) | attract | repel |

We also need the idea of chemical bonding to produce more detailed models of what is happening (at submicroscopic scales) when chemical reactions occur, as we have seen this involves some previously bound combinations of quanticles breaking apart, and new arrangements forming (see Chapter 3). Again this is a complex topic, but a useful rule of thumb is that reactions allow rearrangements that lead to the component quanticles holding together more tightly – or being more strongly bound (or bonded) together. When we consider this idea in relation to bonding being an electrical effect,

we can think of chemical reactions occurring when they allow the charges present – the atomic cores and the outer shell electrons - to rearrange into a more stable pattern. As we will see later in this chapter, that is not the whole story, but it provides a useful starting point for an explanation of what is going on that - if not complete - is at least scientifically sound and can be understood by upper secondary level students. To summarise, this is a topic that may not be studied in any detail by some secondary students (see Chapter 12); but is important enough that when it is part of the curriculum, we should teach it using scientifically sound models in ways that help students see the nature of chemistry: i.e. as a science that develops theoretical models that help us explain phenomena.

**The nature of the chemical bond**

At the level at which we talk about bonding in secondary school chemistry, we can think of it as an electrical phenomenon. The basic components of matter are electrically charged (the nuclei containing positive protons, and negative electrons), and in chemistry it is often most useful to think of chemical change in terms of reconfigurations of positive atomic cores surrounded by negative electrons. Similar charges repel (atomic cores repel each other; electrons repel each other) but opposite charges attract (cores attract and are attracted by electrons), and the attractive forces pull the cores and electrons together into arrangements until similarly charged components are close enough for the repulsion to balance the attraction. This forms an equilibrium arrangement, which is stable because any small disturbance from this position will be resisted by the electrical forces. This applies to atoms, ions, molecules, lattices etc.

When students first meet the idea of the chemical bond, they have limited experience on which to construct a model in terms of electrical forces. Lower secondary students usually have limited appreciation of forces between charges, and they commonly think of bonds as not just being physical (a force) but actually material. They may think in terms of glue, sticks, springs, and various other physical connectors – and some of these may well derive form the physical models we use in teaching (where plastic spheres are connected by such material means). It is difficult for students to think in terms of bonding that is not due to some material link, even if they can

logically appreciate that at the level of electrons and ions and molecules, it is not possible for there to be smaller material links holding them together. Handling this as a teacher requires some subtlety. When teaching abstract ideas, it is important to find connections with students' existing experience and ideas, however the notion that the chemical bonds are material links is clearly a serious misconception. As teachers we need to welcome students making comparisons with what is familiar – the bond *acts like* glue, or *like* an elastic band joining two things together – whilst trying to ensure that this analogy does not become seen as an identity (e.g. not that the bond *is* a spring between two atoms).

Some students may think of the bonding in the atom as being like that due to magnets. This is not a perfect comparison (as magnets have two poles, unlike charged particles), but students will be familiar with magnetism acting as a physical force that can occur without material connection, and in this sense this may be a more appropriate analogy. However it remains important to emphasise that the bonding is electrical, and acts *like* magnetism. (At higher levels of study those students who continue with chemistry will find that actually there is a magnetic component to atomic structure and bonding, but at the introductory level this is an unhelpful complication. Indeed, ultimately, magnetism is an electrical phenomenon, but again that is an unhelpful complication here.)

**Quantum theory**

The electrical model is not the whole story, as it does not explain why the stable arrangements produced by the electrical forces so often result in species with particular electronic configurations (two electrons in the first shell, and often eight in the outermost shell of a species). This is not normally tackled until more advanced levels of study, as it is usually considered to be a difficult topic. In principle, however, this relates to how at the tiniest scale, everything is quantised: such quantities as energy and angular momentum are found to exist in minimally-sized packets. This is very similar to the way matter and charge are quantised, giving us basic units such as electrons that cannot be divided into anything smaller. That also seems a very counter-intuitive idea, but one that students usually come to accept given plenty of opportunity to explore and use our particle models of matter.

A strong clue that the electrical model cannot be the whole story is how in atomic and molecular structures, electrons are often considered to be found in pairs – pairs of electrons form covalent bonds between atomic cores, and non-bonding pairs of electrons on some atoms form hydrogen bonds with hydrogen atoms bonded to other atoms (see below). If electrical forces were the only important factor, then electrons would not be expected to act as paired-up. Yet students rarely seem to raise this as an issue, something that indicates how difficult it is to persuade them to think primarily in terms of electrical interactions. A teacher who finds students raising this objection can probably consider they are doing better than most in helping students think about matter at the submicroscopic scale in the way chemists do.

(The electrons do repel, but the existence of quantum mechanical spin can in effect reduce this effect for pairs of electrons. One model of this is to think of electrons as tiny magnets which can be arranged anti-parallel - N-S and S-N - so that they have a magnetic attraction to counter the electrical repulsion.)

## 3.6 Modelling varieties of chemical bond

At a basic level, the different types of bonds discussed in chemistry can all be understood as due to the electrical interactions between different species, attracting them together until the repulsions between similar charges balance the attractions between opposite charges (see Figure 3.7). There is often a net attraction between neutral species such as molecules, for example, until they get very close when this is balanced by repulsions. This is an essential prerequisite for matter as we know it to exist – otherwise the universe would probably either have no condensed matter, or would be one large neutron star!
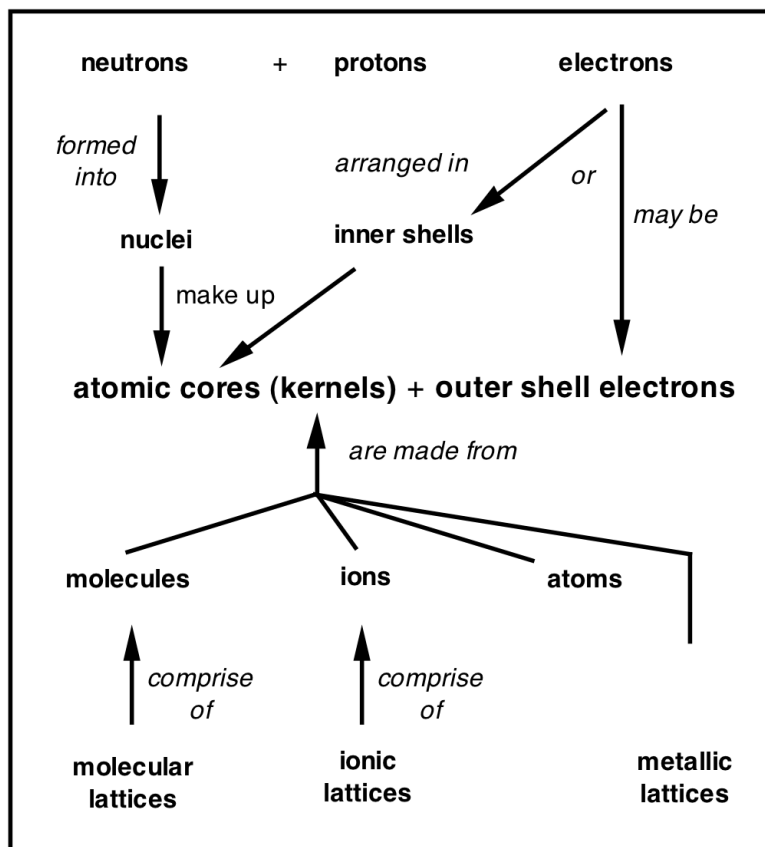
**Figure 3.7: A way of thinking about the structure of matter at submicroscopic levels**

There is some ambiguity among chemists and in different textbooks over whether the term 'chemical bond' should be applied to all or just some of these interactions:

| usually agreed to be chemical bonds | not always seen as chemical bonds |
| --- | --- |
| covalent (including dative) | intermolecular interactions (such as van der waals' forces) |
| hydrogen | |
| ionic | solvent-solute interactions |
| polar | |
| metallic | |

Such a distinction is usually based on considering those interactions labelled as bonds as being stronger than the others. However, that is not a clear distinction, as we see when something like NaCl (with strong chemical bonding) readily dissolves in water

– where the strength of the interactions between the ions and the polar water molecules are a major factor in its solubility.

Even the term 'intermolecular' bonding needs to be applied carefully. Hydrogen bonding (usually considered to be a chemical bond, although sometimes that status has been questioned) can be intramolecular (being very important in determining the shape of proteins and nucleic acids, for example) or intermolecular.

Dative bonding *could* be considered intermolecular as well, so if gaseous ammonia ($NH_3$) and hydrogen chloride (HCl) are mixed, they form a solid ($NH_4Cl$). This may be considered to be an 'adduct', the name given when two already stable molecules are able to join into a new larger molecule (i.e. $NH_4Cl$ might also be written as $NH_3.HCl$). If the product here, ammonium chloride, is gently warmed then it thermally decomposes back to hydrogen chloride and ammonia.

$$HCl \quad + \quad HN_3 \quad \rightleftharpoons \quad NH_3.HCl \quad (or \quad NH_4Cl)$$
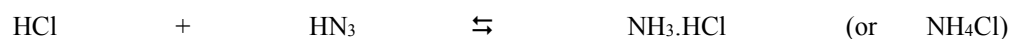
The formation of the ammonium chloride is easily demonstrated in the lab: a long glass tube with cotton wool at either end can be used (see Chapter 2). As little energy input is needed to bring about the decomposition, there is a tendency to see the bond formed as intermolecular, and consider the adduct as not being a fully stable molecule. Clearly such complexities are unhelpful when introducing the topics of chemical bonding to students. Moreover, individual students have been found to use terms such as bond, bonding, attraction etc., in their own idiosyncratic nuanced ways.

It is more sensible to consider all of the electrically based interactions that hold together molecules (internally), lattices, and complexes such as adduct and solvated species as forms of chemical bond, and make it clear that some bonds are much stronger than others, and so much more energy is needed to disrupt the stronger bonds.

**A recommended teaching sequence**

Students are prone to see all materials as made up of molecules, and in particular to see ionic compounds as containing discrete molecule-like entities. This tendency is encouraged when introducing covalent bonding first. Here it is recommended that you take metallic bonding as a starting point for thinking about bonding, moving on to first ionic, then covalent bonding. It is important to be explicit that what you are teaching are *models* that chemists use to explain properties of different substances. These 'first-order' models offer a good deal of explanatory power, but do not explain everything, and do not fit all cases.

**Metallic lattices**

Metals are elements, and so having relatively simple structure provide a good place to start modelling bonding. At the simplest level, the structure of metals can be modelled as a regular arrangement of atomic cores, that is metallic cations, which despite all having a positive charge are held in their lattice positions because of a large number of electrons able to move around and between them, acting as a kind of 'electrical glue'. This is of course a metaphor, and should be used carefully.

A very common metaphor used to describe this arrangement is that the electrons form a 'sea' in which the cations are immersed. However, students using the 'sea of electrons' idea sometimes understand it to mean that there is a *vast excess* of electrons, whereas the stability of the metallic lattice depends upon its overall neutrality. In teaching this model it is important to stress that the number of electrons per cation is the same as the magnitude of the core charge, as the overall neutrality of the lattice is an important factor in its stability. So in the example of magnesium (see Figure 3.8), the electronic configuration of the atom is 2.8.2, which can be modelled as a positively charged atomic core, plus two valence electrons.
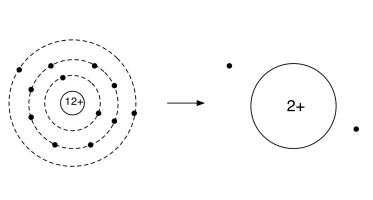
**Figure 3.8: magnesium atom as an atomic core (Mg²⁺ cation)**
**plus two valence electrons**

The metallic lattice consists of a vast array of these cores, $Mg^{2+}$ cations, with the associated electrons. We can think of the lattice forming when a vast number of magnesium atoms come close enough together for their outer shells to overlap and merge so the electrons in them can move throughout the array. This is shown in Figure 3.9, where a sectional slice is shown – with the electron positions shifting from one moment to the next. It is useful to also use three-dimensional models to emphasis that the structure is not just ordered in two dimensions.
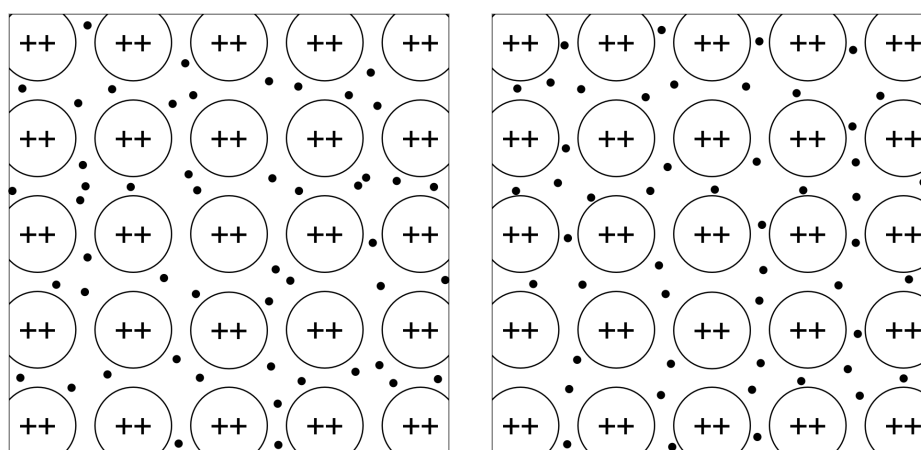


**Figure 3.9: Metallic bonding holds together the lattice in metals**
**such as magnesium – the bond comprises of the mutual**
**attraction between the metallic atomic cores, and the delocalised**
**electrons able to move around the structure**

A simple way to reinforce the electrical neutrality of the metallic lattice would be to present students with a series of images (similar to those in Figure 3.9) showing core charges of +1, +2, and +3, some showing a balance of charge, and some with substantially too few, or too many electrons. Students could be asked to work in

groups to identify which figures represent plausible structures, and to justify their decisions. This activity might be especially suitable for less highly achieving students.

The simple arrangement in Figure 3.9 reflects a 'cubic' arrangement of cations. This is just a model. Although some metallic structures are cubic, most are based on each ion in a slice of the lattice being surrounded by six others in a hexagonal arrangement (see figure 3.10), as this actually gives better packing ('close' packing) when the next layer is offset.
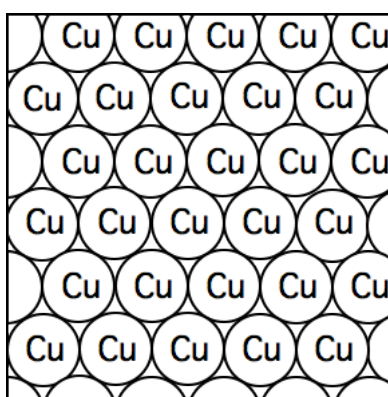


**Figure 3.10: A representation of close-packing in copper**

Modelling this with spheres (such as marbles, or expanded polystyrene balls) shows that there actually two regular ways of building up such layers, depending upon whether the third layer sit directly above the first ('ABAB'), or is offset from both the first two layers ('ABCABC'). This can provide the basis of simple practical modelling activity for students to build models of the two different arrangements using suitable spheres (such as expanded polystyrene balls one size). Although secondary students are not usually expected to know about these specific structures, the modelling activity provides a suitable group practical that can reinforce teaching about the regular nature of the metallic lattice. With some students, close instructions for building the model will be appropriate, whereas for others they could simply be issued the challenge of producing two non-identical structures with all the spheres close packed.

The apparent ambiguity between figures such as figure 3.9, which shows space between the atomic cores for the electrons, and physical models showing the close packing provides an important teaching point about the difficulty of modelling

quanticles – atoms and ions do not actually have clearly defined surfaces of boundaries, but rather become more tenuous further away from the nucleus.

A more advanced way of thinking, usually only met at college ('sixth-form') level explains metallic structure in terms of a more complex orbital model, where the metallic bonding is formed by the overlap of atomic orbitals, and the outcome is an enormous number of 'molecular' orbitals that will have a complex pattern of geometries. However, they will also form a virtual continuum of energy levels (the 'conduction band'), so although each particular molecular orbital may put restrictions upon occupying electrons, the available thermal energy is sufficient for electrons to readily move between orbitals in the band.

Although this is a more complex picture, and will only met by advanced learners, it is important to ensure introductory teaching will not act as an impediment to later progression for those students who do continue with the subject. So, in teaching in terms of overlap *of shells* we should be careful to stress that this is a model, and somewhat simplified, so that students who may study chemistry at higher levels do not become too committed to that particular picture. Given this proviso, the overlapping shells model can act as a much simpler 'version' of the molecular orbital/conduction band model to explain the delocalisation of the valance electrons – the 'conduction' electrons so important to the properties of metals.

The key points when introducing metallic bonding is that the cations form a regular pattern, and are bound by the attraction between the positive cations and the negative electrons. As always, our diagrams need to be presented to students as representations designed to emphasise certain points, and not as realistic images of how metals actually are.

**Ionic lattices**

Somewhat more complicated than the metallic case, is that of ionic compounds. Research shows us that students very commonly misunderstand ionic bonding. Ionic bonding is often taught through a convention of considering atoms of an electropositive metal and an electronegative non-metal - often Na and Cl are used as

the example - and then considering how they might interact to form ions, Na$^+$ and Cl$^-$, which would then bond together (see Figure 3.11).

This does not reflect a likely chemical process, but is just a kind of 'thought experiment' in how ions might be formed from atoms, in the unlikely event that such unstable species should be around to happen to interact. However, students often think such a scheme is actually how ionic bonds are formed, and - worse - often think that the electron transfer depicted in such schemes 'is' the bond!
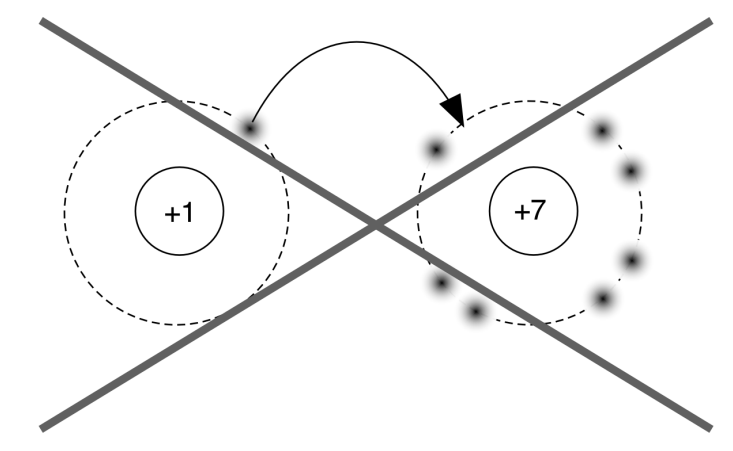


**Figure 3.11: A common student misconception of the ionic bond**

However, we *do not have to consider how ions come about to explain ionic bonding*: indeed presenting such schemes only encourages students to think of all chemical processes as starting with atoms, rather than more feasible reactants. In the natural world there are many materials that already contain ions like Na$^+$ and Cl$^-$, but we never find *atomic* sodium or chlorine under natural conditions. (Indeed, there are very few materials that contain discrete atoms: samples of the noble gases being the obvious exceptions.) In the case of metallic bonding it was useful to start thinking about atoms, but thinking in terms of atoms in explaining ionic bonding is an unhelpful mind-set.

It is recommended that in introducing ionic bonding to students, it is more useful to think in terms of a more feasible chemical context, such as in terms of reactions they

are expected to be familiar with. One example might be the neutralisation of an acid and an alkali, such as

---

hydrochloric acid    +    sodium hydroxide    →    sodium chloride    +    water

---
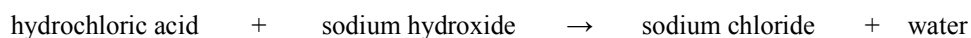
In this case the reactants are solutions containing ions, and the actual chemical reaction is between hydrogen ions and hydroxide ions to form more water. This leaves sodium and chloride ions in solution: $Na^+_{(aq)}$, $Cl^-_{(aq)}$. Solid sodium chloride does not form because the ions are too strongly hydrated (they are each bonded to a sheath of solvent molecules, which forms an ad hoc complex in the solution). However, if the solvent, the water, is allowed to evaporate (see Chapter 1), then this leaves the ions, which organise into a regular array of cations and anions because of the mutual attraction between oppositely charged ions (see figure 3.12).
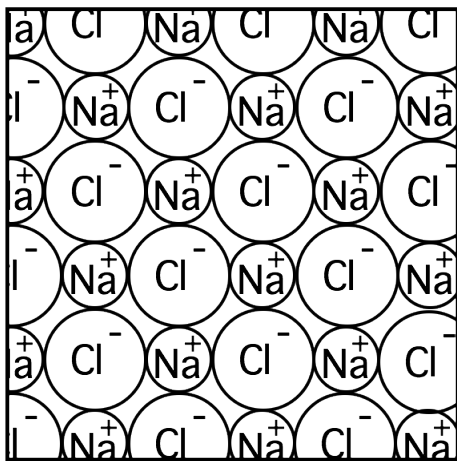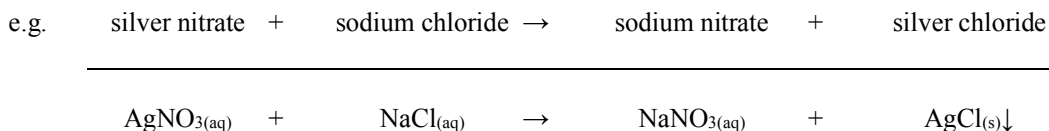


**Figure 3.12: Representation of the ionic sodium chloride lattice
in two dimensions**

An alternative context for introducing ionic bonding would be a double decomposition reaction, because the students can see the formation of the ionic product immediately. So if silver nitrate solution is added to sodium chloride solution, the formation of the precipitate, silver chloride, is immediate.

|  | e.g. | silver nitrate | + | sodium chloride | → | sodium nitrate | + | silver chloride |

$$AgNO_{3(aq)} \quad + \quad NaCl_{(aq)} \quad \rightarrow \quad NaNO_{3(aq)} \quad + \quad AgCl_{(s)}\downarrow$$

It is important to stress to students that although we call the reactant solutions 'silver nitrate' and 'sodium chloride' solutions (as they are solutions of these compounds), the solutes do not exists as bonded compounds in the solution, but rather as ions which are mixed into the solvent. However, when the two solutions are mixed together so that it contains silver, sodium, chloride *and* nitrate ions, the attraction between silver and chloride ions is strong enough that when they collide in the mixture they bind together, eventually forming large clumps that settle from the solution to give a solid silver chloride 'precipitate' (see figure 3.13).
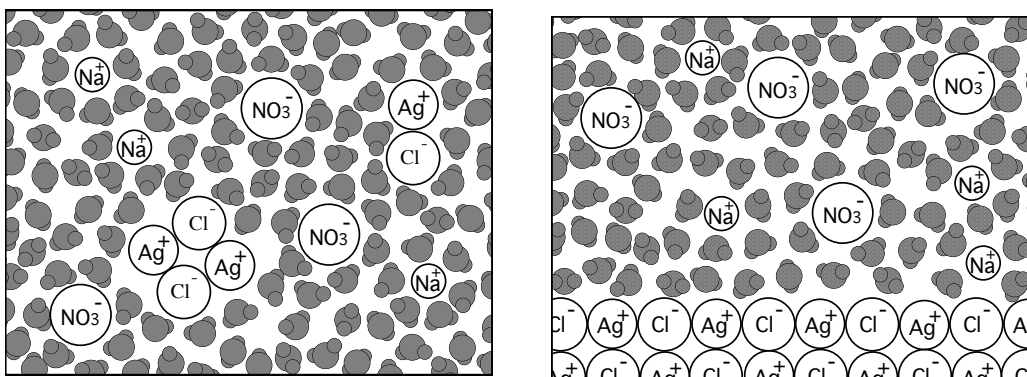


**Figure 3.13: When solutions of silver nitrate and sodium chloride are mixed, the silver and chloride ions bond together forming a solid with an ionic lattice.**

In this process an ionic lattice is formed because of the attraction between the oppositely charged ions. In essence, that *is* the bond. There is no need to explain how the ions came to be – they exist in materials available to chemists, in substances we call salts ($NaCl$, $NaBr$, $KCl$, $K_2SO_4$, $Ca(NO_3)_2$ etc). Most of the material on earth comprises of the output of nuclear processes that took place in stars, where the temperature is much too high for individual atoms to exist (and so matter is in the form of plasma, a kind of gas comprising of separate nuclei and electrons). After this material was exploded into space, at the end of the star's 'life' cycle, it cooled and

formed into more stable combinations of nuclei, and electrons: sometimes atoms, but sometimes ions or molecules. Very little of the matter that formed the earth is of the form of discrete atoms. The formation of the ionic lattice due to ionic bonding does not involve any mysterious 'electron transfer', and if ionic bonding is taught this way there is no need for students to think in those terms.

Of course the formation of sodium chloride by binary synthesis *does* involve the formation of chloride ions (arguably the sodium ions are already present in the metallic lattice), but from molecules, not individual atoms. Students may see this reaction demonstrated, but they are unlikely to carry it out (on safety grounds); and neither metallic sodium nor chlorine gas are common laboratory reagents. Whilst the binary synthesis route certainly offers an exciting demonstration of a vigorous reaction, forming sodium chloride by neutralisation followed by evaporation is much more practicable way of producing sodium chloride.

This is important because research tells us that students often think that in the NaCl lattice, for example, there are NaCl molecules, or at least discrete ion-pairs, which are bound because they have a history of having transferred electrons: so despite the symmetry of Figure 3.12, students often interpret such figures as a collection of NaCl molecules which have ionic *bonds* within them, but are then just attracted to each other *just by forces*. This is unfortunate as such misconceptions (a) cannot help students understand why NaCl is hard, and has a high melting temperature, and (b) often lead them to expect NaCl molecules to be the solvated species when a solution is prepared. These misconceptions all seem to derive from teaching the ionic bond through fictitious electron transfer events between isolated atoms, which actually are quite irrelevant to the chemistry! This idea appeals to students so much that they will sometimes explain precipitation reactions, such as our AgCl example in terms of

- the silver ion getting its electron back from the nitrate, and

- the chloride ion giving back the sodium atom its electron, so that silver and chlorine are back to being atoms;

- allowing the silver atom to then give an electron to chlorine atom, to reform the silver and chloride ions (which were already present, of course) with an ionic bond between them.

This scheme is actually much more complicated than the scientific model of silver ions sticking to chloride ions because of their opposite charges, and illustrates just how tenacious some misconceptions can be, once they have a hold of a student's imagination.

**Covalent bonding**

Covalent bonding tends to occur between non-metallic elements, and is often described using the metaphor of 'sharing' electrons. A covalent bond is understood as occurring when the valence shells of two atoms overlap so that one (or more) pairs of electrons fall within the valence shells of both atoms. At more advanced levels this is described in terms of the interaction of atomic orbitals on different atoms forming molecular orbitals (and the pair of electrons occupying the lower energy 'bonding' molecular orbital), but at an introductory level, atoms are often represented as having overlapping outer electron shells.

We can represent covalent bonds in a variety of ways, which can be confusing for students. Experienced chemists and teachers see past these differences in representational formalism, but the reasons for different ways of drawing the same thing may seem arbitrary to learners. Indeed, when looking at a range of student textbooks there is often no obvious reason for the preferred forms of diagrams used. These images represent 'quanticles' – entities that are fuzzy and often better thought of as clouds of charge than as like tiny billiard balls with definite surfaces. This may not be what our common forms of representation suggest (for example, drawing electron shells may give the impression that they are solid structural elements of atoms), so it is up to the teacher to emphasise the limitations of images which are often a compromise between what we can easily draw, and particular points we wish to represent in particular images.

Some examples of how covalently bound molecules may be represented are illustrated here. Figure 3.14 shows a molecule of fluorine where the covalent bond is represented as the pair of electrons where the outer shells overlap.
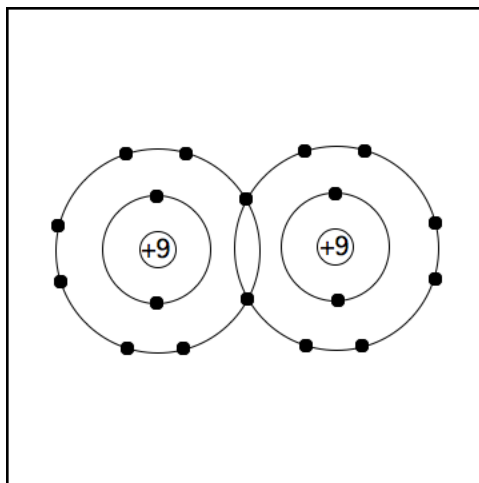


**Figure 3.14: One possible representation of a molecule of fluorine ($F_2$)**

In figure 3.15, the chlorine molecule is represented in an alternative way, with only the valence shell electrons shown, and the bonding pair of electrons shown inside the overlapping atoms. It is not sensible to ask which of these pictures is a more accurate representative of the molecule (which is too small to be visible, three dimensional, with electrons in motion) itself, but rather it is important to explain to students which features are being foregrounded in different forms of representation. In terms of the covalent bond itself, the essential feature is that the pair of electrons is electrically attracted to, and by, the positive nuclei, and so act as a bond.
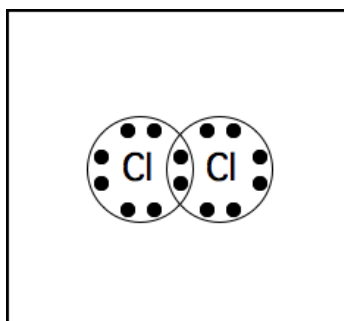


**Figure 3.15: One possible representation of the chlorine ($Cl_2$) molecule**

Figure 3.16 shows some alternative ways of representing the methane ($CH_4$) molecule. The first of these uses a 'dot and cross' style of showing valence shell electrons. This can be helpful for students in *keeping account* of electrons by showing the electrons from different atoms in a different style. However, in the molecule, there is no difference between the electrons, and the interactions between the electrons and the positive nuclei are completely impendent of where the electrons derive form. This should be stressed, as some students assume that each electron in the bond is more strongly attracted to its 'own' atom, and that on bond fission, the electrons will always go back to the atoms they came from (which will interfere with later learning about heterolytic bond fission). Students should be nudged from thinking in terms of the 'ownership' and 'history' of electrons, to instead thinking simply in terms of the electrical forces acting between the different charges present.
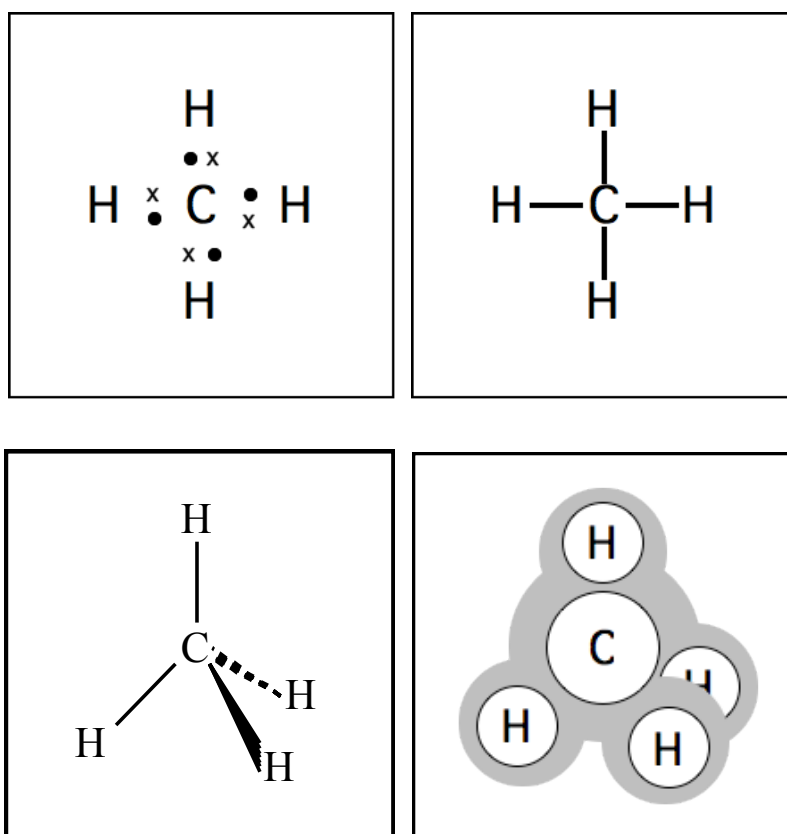


**Figure 3.16: several representations of the methane ($CH_4$) molecule**

The second representation in Figure 3.16 simply shows the bonds as lines connecting the parts of the molecule together. This is an easy representation to draw (which can

be important with more complicated molecules), but is probably best only used once students are used to the idea that the bond is the electrical interaction due to the electrons found between the different atomic nuclei. The third representation shows how the molecular structure is not flat like the surfaces where we draw, but takes up a three dimensional arrangement (tetrahedral in this case, due to the mutual repulsion of the electron pairs in the bonds). The final image replaces the representation of discrete bonds, with a representation of the 'clouds' of electron density around the atomic cores – i.e. how we can imagine a time-averaged image of where the electrons are to be found. (From a quantum mechanical perspective, the structure of molecules is better described in terms of the probability of finding electrons in particular positions, but it is useful to think of this as the electrons moving about and so smearing out their charge density.) A more sophistication version could show the variations in electron density – more like a contour map.

As I have suggested throughout the chapter, it is not helpful to talk about *which* of the various possible representations is 'best'. Students should be made aware molecules and other 'quanticles' are not easily drawn, and that scientists will model them through representations that stress particular relevant features. There is a key issue here for teachers, as it usually makes good sense to adopt particular conventions in teaching and then to consistently use them to limit learning demand for students. Yet figures students see in books and on-line will be diverse, and reliance on one form of representation can lead to students treating that form as a realistic image of how molecules actually are. A sensible compromise would seem to be in order: that is initially using a preferred form to represent molecules but later (once students are used to seeing and drawing images of molecules) introducing variation where context makes others forms useful to make particular teaching points (e.g. about molecular shape or the presence of double bonds).

**Bond polarity**

At an introductory level, students tend to be given the impression that bonding in compounds is covalent or ionic, as if this is a dichotomy:

| Covalent bonding | Ionic bonding |
|---|---|
| In non-metallic elements and compounds of non-metals | In compounds of metals with non-metals |

Students often come to see these different forms of bonds as fundamentally very different, making it difficult for them to later appreciate how few bonds are 'pure' covalent and indeed no bonds are 'pure' ionic.

The ionic bond as represented in introductory chemistry texts is an ideal, and most compounds thought of as ionic are actually someway from having fully ionic bonding. (At advanced levels, students will learn how tables can be used to estimate the percentage of ionic and covalent character, depending upon the electronegativity difference between the elements.) Fully covalent bonds only usually exist between atoms of the same elements, and strictly then only where those atoms are not themselves bonded to very different atoms – so the C-C bond in ethanol is not pure covalent, as one of the carbon atoms is bonded to an electronegative oxygen atom, which will influence the carbon-carbon bond through an 'inductive' effect, distorting the geometry of the electron density in the carbon-carbon bond (Figure 3.17).

$$\begin{array}{ccc} \text{H} & & \text{H} \\ | & & | \\ \text{H} - \text{C} \rightarrowtail \text{C} \twoheadrightarrow \text{O} - \text{H} \\ | & & | \\ \text{H} & & \text{H} \end{array}$$
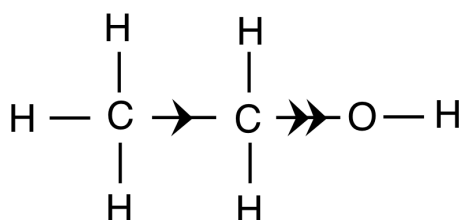
**Figure 3.17: The C-C bond in ethanol is not a 'pure' covalent bond, because the oxygen atom (core charge +6) attracts electron density and distorts the bond indirectly.**

These issues are ignored in introductory treatments, but some options for representing molecules offer a better starting point for later progression in students' thinking. So Figure 3.18 shows a representation of a tetrafluoromethane ($CF_4$) molecule showing the core charge and valence electrons. If students are taught about bonding as an electrical interaction, it seems clear that the electron pairs in the bond will – all other

things being equal - be pulled closer to the +7 charge (fluorine atomic core) than the +4 charge (carbon atomic core). So the bond here will be polar rather than purely covalent.
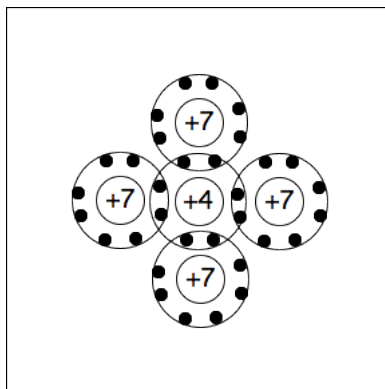


**Figure 3.18: a simple representation of the tetrafluoromethane (CF₄) molecule**

Figure 3.19 shows a representation of an interhalogen compound, ClF, using the same format. Here both halogen atoms have the same core charge, +7, but because of the difference in size of the two cores (for fluorine, nucleus {+9} and one shell of 2 electrons; for chlorine, nucleus {+17} and two inner shells {2.8}), the equilibrium position for the bonding pair of electrons is nearer the fluorine nucleus. So again, this is a polar bond.

At advanced levels there are various ways of showing bond polarity (electron position, the use of $\partial+$ and $\partial-$ symbols to indicate 'partial' charges on atomic centres etc.) Whilst such detail is not needed in introductory treatments, it is important to teach bonding as primarily an electrical interaction, to help students later appreciate how ionic and covalent bonds can be understood as extremes on a continuum, and not a simple dichotomy where all bonds in compounds easily fit one or other category.
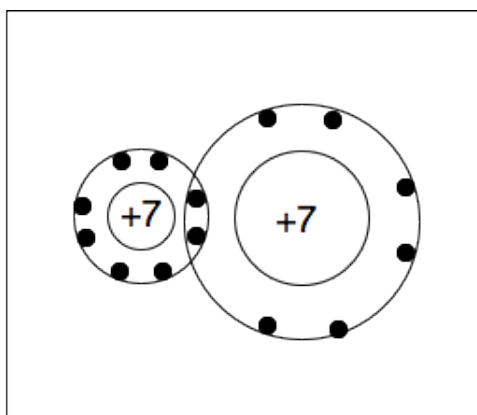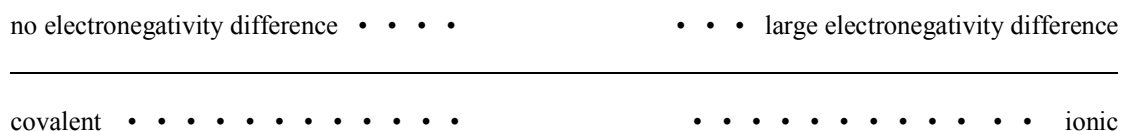
**Figure 3.19: a representation of the Cl-F molecule**

The extent to which a bond will be polar then depends on the difference in electronegativity of the elements involved:

no electronegativity difference  • • • •                    • • • large electronegativity difference

_____

covalent  • • • • • • • • • • •                    • • • • • • • • • • • •  ionic

**Other forms of bonding**

In secondary chemistry it is common to limit explicit discussion of bonding to the metallic, ionic and covalent cases. However, students will have come across the idea that substances in the solid state are held together by some form of bonding when learning about the basic particle model of matter (see Chapter 2) and so are likely to assume that this bonding will be ionic in ionic compounds, metallic in metals _and covalent in material with covalent bonds_. The latter assumption would be correct in carbon, silicon and other substances with giant covalent lattices. However, there are many materials with discrete covalent molecules, which exist in the solid state at room temperature because of the weaker interactions between molecules. For example, wax and polyethylene (polythene) contain molecules that have covalent intra-molecular bonding, but are attracted to each other by a weaker form of (inter-molecular) bonding.

So Figure 3.20 shows the molecules in sulfur ($S_8$), which is in the solid state at room temperatures, and can be melted by heating in a test tube over a Bunsen flame (care if

this is demonstrated – some of the sulfur may burn, leading to noxious vapour being released).
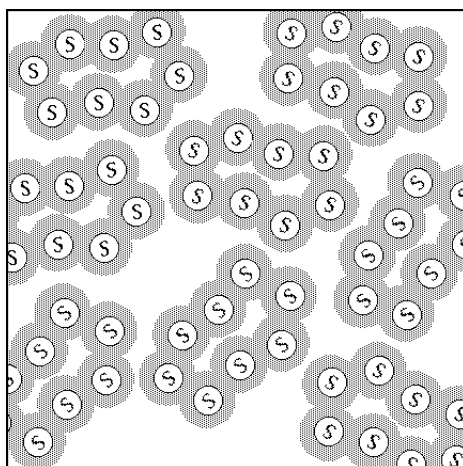


**Figure 3.20: Molecules in sulfur**

Sulfur comprises of molecules in the form of rings of eight atomic centres, with each atomic core bound to two neighbours by a covalent bond. However, sulphur exists in the solid state at room temperature because the molecules are attracted to each other (and actually fit together to give a crystalline structure). This is despite there being no electron 'sharing' *between* molecules, or any ions present. This is normally explained in terms of the electron movements in adjacent molecules becoming synchronised to give 'fluctuating transient dipoles' that allow areas of higher electron density (overall negative charge) on one molecule to attract and be attracted by areas of lower electron density (overall positive charge) on an adjacent molecule.

This is difficult to visualise, and is not usually discussed with students in introductory chemistry. However it is important students appreciate that neutral molecules will attract together because of the charges present, even if no details are offered. This will help to avoid the common misconceptions that so-called 'molecular solids' have covalent bonds throughout, and the corollary that, as many of these substances melt readily when in the solid state, covalent bonds are often quite weak.

Students need to appreciate that bonding effects are generally explained in terms of interactions between charges (and it may be worth pointing out that the forces attracting molecules of sulfur together can not be explained in terms of forming octets or full shells!) Figure 3.21 develops this principle (summarised earlier in Figure 3.7)

by showing how different chemical structures are built up of different configurations of atomic cores and valence electrons. Although it is certainly *not* sensible to present such a scheme to be learnt by students first meeting bonding ideas, it can usefully inform teaching. Teaching consistent with this way of thinking is more likely to help students appreciate the principles common to different forms of bonding, and will better support progression for those who go on to more advanced study.
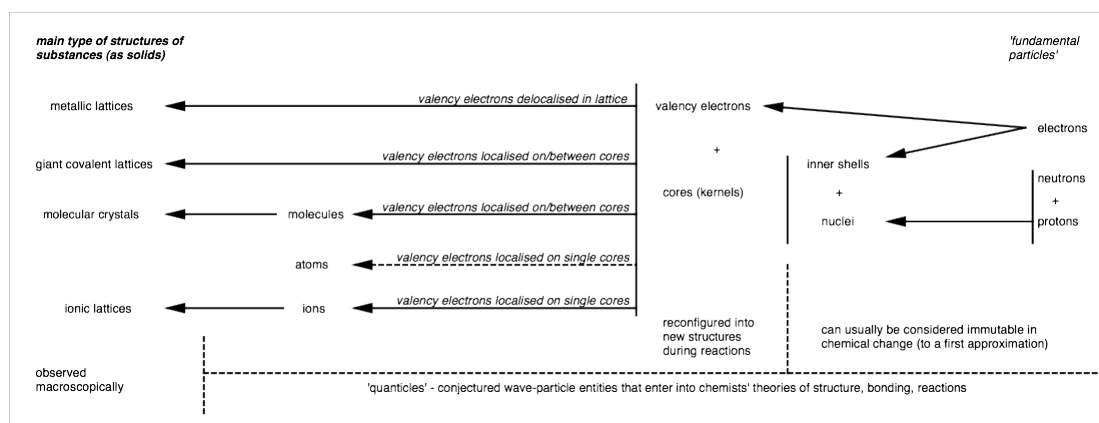


**Figure 3.21: Macroscopic structures reflect different ways of binding atomic cores together with valence electrons**

These types of interactions, often called van der Waals' forces, are not the only important types of bonding beyond metallic, ionic and covalent bonds. Solvent-solute interactions may be due to transient dipoles, but can also often involve permanent (rather than just transient) polarity on molecules. This is why water is a good solvent for ionic materials.

An important example of an interaction due to bond polarity, is the hydrogen bond, which is so important for the properties of water, protein and nucleic acids. Unfortunately, students sometimes come across hydrogen bonding discussed in biology lessons (where the nature of the bond may not be explained) before it has been introduced in chemistry, leaving them to infer what is being referred to. (A common guess in this situation seems to be that it is just a covalent bond to a hydrogen atom.) Hydrogen bonds form between electronegative atoms (usually O, F, N, sometimes S, Cl; all of which have one or more pairs of non-bonding electrons, so called 'lone pairs') and hydrogen atoms that have polar bonds to other electronegative

atoms. Figure 3.22 shows hydrogen bonding between two molecules (a dimer) of ethanoic acid (as found in 'glacial', i.e. solid, ethanoic acid).
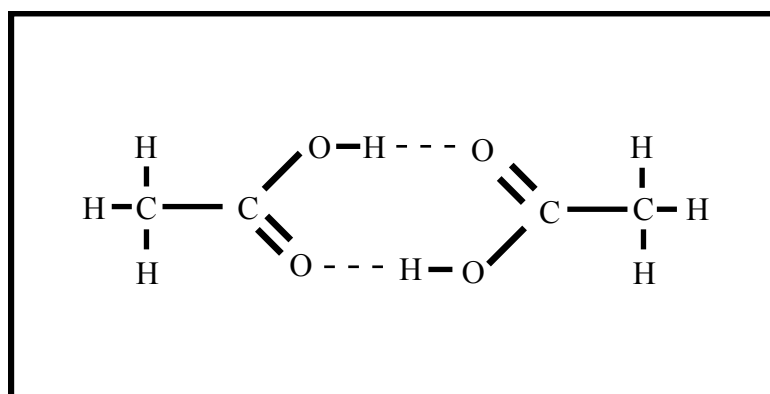


**Figure 3.22: hydrogen bonding in ethanoic acid**

Hydrogen bonds can be quite strong for intermolecular bonds (as a hydrogen atom bonded to an electronegative atom is a positively charged proton which has very limited electron density covering its 'rear' and so is readily attracted by/to 'lone pairs' of electrons on other atoms.) Researching into the nature of the hydrogen bond could be a useful extension activity to differentiate for more able ('gifted') students in a mixed-ability class. For example, hydrogen bonds have a specific geometry (the hydrogen bond being approximately opposite the polar bond), whereas most intermolecular bonds are non-directional: something that could be followed upon by a gifted student by exploring orbital models of bonding. A small group of more able students could be set the task of building a model of ice structure (or even part of a nucleic acid molecule) showing the importance of the intermolecular bonds in the crystal structure.

## *3.7 Relating bond type to changes in chemistry*

Sometimes the weaker forms of interaction, such as van der Waals' forces, solvent-solute interactions, and even hydrogen bonding, are considered to not really count as chemical bonds. However, as the strength of most types of interactions vary considerably, it makes more sense to think of there being a range of different types of bond, some of which are *usually* stronger, and some of which *tend* to be weaker. So

sulfur, held together by van der Waals' forces, has a higher melting temperature than mercury, although the latter has metallic bonding.

**Melting and boiling**

A simple change of state, such as melting, has different consequences for substances depending upon the type(s) of bonding present.

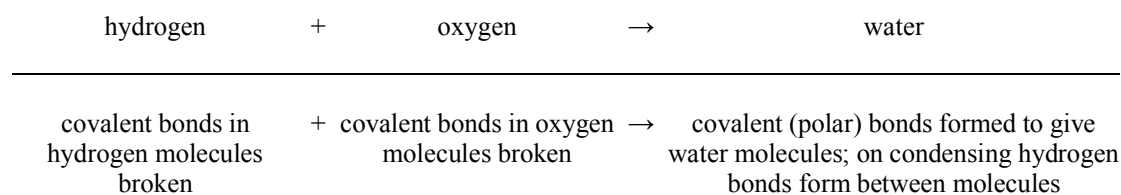| Metals | Metallic bonds are not strongly disrupted by melting (e.g. the metallic bond is not dependent upon a particular geometry, so the metal is still held together by metallic bonding in the liquid state, and – for example - continues to conduct electricity), and some metals have quite modest melting temperatures (e.g. Na c.98°C), but boiling requires overcoming the metallic bonding completely (so sodium has a boiling temperature of 890°C). Group 2 metals have higher melting temperatures than group 1 metals (as would be expected from having greater core charges and more delocalised electrons). |
| --- | --- |
| | Transition metals tend to have higher melting temperatures than main group metals, and their bonding is said to include some 'covalent character'. (Here our model that bonding only involves the outermost shell of electrons is found to have limitations.) |
| | Boiling a metal is basically a form of atomisation – although metal vapours may also include small clumps of atoms. |
| Ionic solids | Ionic solids tend to have high melting temperatures, as the ionic bonding depends upon the ions being arranged in the lattice so cations are next to anions, and not other cations. Vaporisation of ionic materials usually leads to vapours containing discrete ions and some clumps (such as ion-pairs). |
| A substance with a lattice of covalent bonds, such as diamond (C or $C_\infty$) or silica ($SiO_2$) | As covalent bonds are directional and strong, these substances tend to have high melting temperatures. Melting the material requires breaking (not just weakening) of the bonds. |
| Molecular solids | Molecular solids tend to have relatively low melting and boiling temperatures (e.g. nitrogen, methane, carbon dioxide, ammonia etc are in the gaseous state at room temperature), as the bonds between molecules are weak, and the bonds within molecules do not need to be broken for the change of state. |
| Hydrogen bonded solids | Solids with hydrogen bonding tend to have higher melting and boiling temperatures than other materials with similar size molecules. In ice, water molecules form a lattice with each molecule hydrogen-bonded to four others. This is disrupted on melting, although there is a constant flux of hydrogen bonds being formed and broken in the liquid. |

**Dissolving**

Dissolving of a substance in the solid state involves breaking of bonds in the solid, disruption of bonds in the solvent, and formation of new interactions between solvent and solute.

| example | involves | feasible |
|---|---|---|
| alloying | Disruption of original metallic lattices and formation of new lattice – however retains delocalised electrons between positive cores. | Often feasible – some mixtures allow a better 'fitting' lattice than in pure metals. |
| salt in oil | Would require breaking strong bonds between ions, but the ions would not bond strongly to non-polar molecules. | No significant dissolving occurs. |
| salt in water | Requires breaking strong bonds between ions, and disrupting hydrogen bonding in liquid water, but ions often become strongly hydrated as polar water molecules are attracted to the ions. | Some, but not all, ionic solids are very soluble in water. |
| wax in water | The forces between wax molecules tend to be modest, but hydrogen bonding in water is quire significant, and only weak interactions are formed between water molecules and wax molecules. | No significant dissolving occurs. |
| wax in oil | The interactions formed between non-polar solvent and non-polar solute molecules are similar in kind and strength. | Mechanical agitation of the solid by collisions from solvent molecules is sufficient to allow a solution to form. |
| glass in oil or water | Strong covalent bonds in materials such as glass are difficult to break, and would not be compensated for by solvation interactions. | No significant dissolving occurs. |

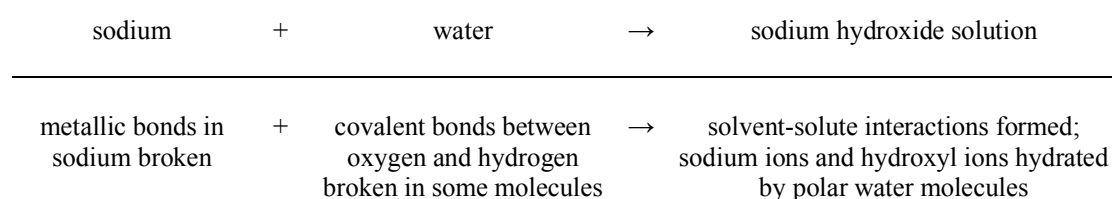**Bonds and chemical reactions**

Chemical reactions seldom occur between substances in an atomic form. Although some reactions of the noble gases have been achieved, these are exceptions. Generally a chemical reaction, such as the various ones described in the chapters of this Handbook, involves both the breaking of bonds in the reactants, and the formation of bonds in the products. Consider a few examples:
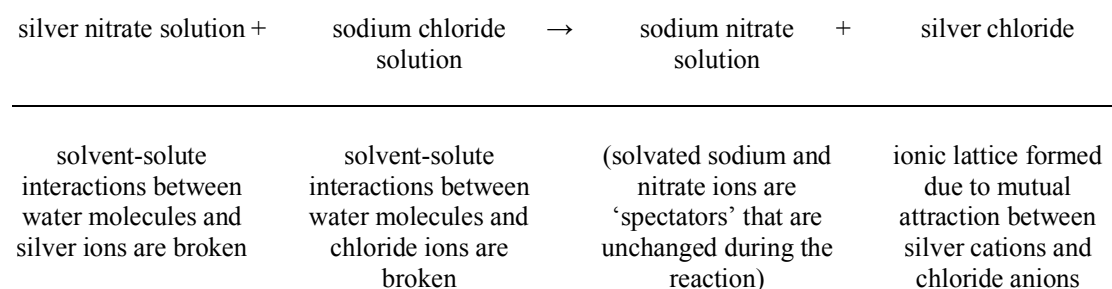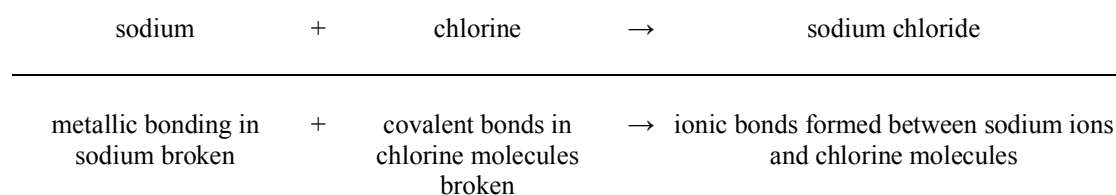
e.g. 1: covalent bonds broken, and formed:

| hydrogen | + | oxygen | → | water |
|---|---|---|---|---|
| covalent bonds in hydrogen molecules broken | + covalent bonds in oxygen molecules broken | | → | covalent (polar) bonds formed to give water molecules; on condensing hydrogen bonds form between molecules |

e.g. 2: metallic and covalent bonds broken:

| sodium | + | water | → | sodium hydroxide solution |
|---|---|---|---|---|
| metallic bonds in sodium broken | + | covalent bonds between oxygen and hydrogen broken in some molecules | → | solvent-solute interactions formed; sodium ions and hydroxyl ions hydrated by polar water molecules |

e.g. 3: ionic bonds formed:

| silver nitrate solution | + | sodium chloride solution | → | sodium nitrate solution | + | silver chloride |
|---|---|---|---|---|---|---|
| solvent-solute interactions between water molecules and silver ions are broken | | solvent-solute interactions between water molecules and chloride ions are broken | | (solvated sodium and nitrate ions are 'spectators' that are unchanged during the reaction) | | ionic lattice formed due to mutual attraction between silver cations and chloride anions |

e.g. 4: metallic and covalent bonds broken, ionic bonds formed:

| sodium | + | chlorine | → | sodium chloride |
|---|---|---|---|---|
| metallic bonding in sodium broken | + | covalent bonds in chlorine molecules broken | → | ionic bonds formed between sodium ions and chlorine molecules |

Once students have been taught about bond types, it is useful when discussing different reactions in various topics to ask students about the types of bonds broken and formed. This will reinforce learning, help to shift them from thinking of bonding and reactions in 'octet' terms, to electrical interactions, and encourage them to try to visualise what is occurring at the submicroscopic level.

Keith S. Taber                    36

## *Further reading and resources*

The Royal Society of Chemistry (RSC) publish a range of recourses to support chemistry teaching in schools. *Chemical Misconceptions - Prevention, Diagnosis and Cure* (2 Volumes, Taber, K. S., 2002). Includes resources for probing student thinking, and finding out whether students have acquired common misconceptions on a range of topics relevant to this chapter. The probes themselves can also be downloaded form the RSC site:

http://www.rsc.org/Education/Teachers/Resources/Books/Misconceptions.asp

A range of simulations of downloaded from the Chemistry Experiment Simulations and Conceptual Computer Animations page from the Chemical Education Research Group at Iowa State University. These include a simple simulation of a particle model of NaCl dissolving in water, and a simulation showing hydrogen bonding between water molecules in the liquid phase.

http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/simDownload/index4.html

The simulations available from The Concord Consortium are mostly suitable for more advanced learners, but some of them will useful for supporting learning with some secondary students. For example, there is a simulation showing how the electron density around two hydrogen atoms are distorted as one is move towards and then away from the other; and a model showing that the distinction between ionic and covalent bonds reflect the extremes of a bond between two atoms where their electronegativities may be varied by 'sliders':

http://www.concord.org/activities/subject/chemistry