

Chemical bonding

Chemical bonding is one of the key concept areas in the subject and is also an area where learners are known to commonly acquire alternative conceptions. Some of these alternative ideas are considered in this chapter, with suggestions for improving the teaching of the topic. Some classroom materials included in the companion volume for eliciting and challenging students' ideas are introduced.

The full shells explanatory principle

Students are found to commonly use the octet rule - a useful heuristic for identifying stable chemical species - as the basis of a principle to explain chemical reactions (see Chapter 9) and chemical bonding.¹ According to this 'full shells explanatory principle' bonding occurs 'in order [for atoms] to try to achieve a stable structure ie 8 electrons in the outer shell of the atom'.

Students relate the 'sharing' of electrons in covalent bonds to the full shells explanatory principle, so that 'the electrons are shared to create a full outer shell', and the 'covalent bond is the sharing of electrons to complete full valency shells'. As one student wrote in a test paper;

'A covalent bond is one in which two atoms join together by the sharing of electrons. Each of the atoms achieves noble gas configuration in the process of covalent bonding.'

Ionic bonding is similarly understood as 'where you, donate, or gain electrons, to form a completed outer shell'. The full shells explanatory principle may also be invoked in students' explanations of metallic bonding, so that one 18 year old student described how 'metals haven't got full outer shells, then by electrons moving around, they're getting, a full outer shell, but then they're sort of losing it, but then like the next one along will be receiving a full outer shell'. Another post-16 student conceptualised metallic bonding as being 'formed by the one, two or three valent shell electrons being donated to lattice, so a noble gas configuration is achieved'.

The full shells explanatory principle is inherently anthropomorphic, as no physical force is invoked to explain why systems should evolve toward certain electronic configurations. Rather, it is assumed that this was what atoms 'want', and so they act accordingly: 'in all cases what an atom is trying to do is become stable'. For example, students may suggest that dimers of aluminium chloride form so that the aluminium atoms would 'think' that they have the octets they need to be stable.

Given that the starting point for many students' thinking about bonding is the atoms' perceived need to achieve a full shell, it is not surprising that often:

- students see chemical bonding and forces within chemical structures as largely unrelated; and
- students limit their category of chemical bond to types of interactions that can be readily conceptualised in terms of the full shells explanatory principle.

It would be an exaggeration to claim that all aspects of students' alternative ideas about chemical bonding derive from this invalid extension of the octet rule. Nevertheless, the full shells explanatory principle can be seen to be extremely influential in guiding the thinking of many students.

When is a chemical bond not a force?

We could define a chemical bond as that which holds the parts of a chemical structure together. Of course, the problems of working from definitions were discussed in Chapter 2, and as with so many other concepts in chemistry the experts (ie chemists and chemistry teachers) can feel they know what they think a chemical bond is, without necessarily being able to provide a rigorous definition.

However, in view of the range of alternative conceptions, some quite tenacious, that have been uncovered in this topic, it is sensible to start this chapter by exploring what a chemical bond is. I would suggest it is a force which holds chemical species together.² This force can usually be considered as an electrical interaction. The source of this interaction is the nature of chemical species themselves - composed of positively charged nuclei and negatively charged electrons.

It will be noted that this description would appear to be just as applicable to individual atoms or ions as to molecules and lattices. A sodium ion is held together by such electrical interactions, just as a hydrogen molecule, or an ice crystal is. By convention we do not usually refer to the interactions within a mononuclear species (a single ion or atom) as chemical bonding, but perhaps this is unfortunate as research shows that students commonly see the chemical bond as being unrelated to the electrical forces within a mononuclear species. Students often do not count 'intermolecular' interactions as chemical bonding either.

If there were no quantum restrictions on where electrons could be located in chemical systems, then teaching about and studying chemical bonding would be much simpler - we would not have different categories of chemical bond such as covalent and hydrogen bonding. (However, all the matter in the universe would collapse into 'neutron stars', so there would a high price to pay for this simplicity.) These quantum restrictions limit the number of electrons in particular 'shells' around nuclei, and lead to the 'complications' of, inter alia, electronic configurations, valency ('combining power'), electronegativity, and patterns in ionisation energies.

Unfortunately this complexity also commonly leads us to present the topic of chemical bonding to students in such a way that they do not appreciate the underlying electrical nature common to the different types of bonding.

Indeed a teacher might sometimes refer to bonding as a 'chemical bond', and other occasions refer to it as 'an interaction' or talk of, say, 'the force' holding molecules together. Yet different students may well have their own idiosyncratic ways of using these various terms that teachers are not aware of.

For example, one of the classroom resources provided in the companion volume is a probe called Interactions which asks students to describe and explain the interactions in a range of chemical species. It also asks students whether the interactions depicted would be classed as 'attraction', 'force', 'bonding' and 'chemical bond'. This may seem a strange and even pointless question, but interviews with college students have suggested that learners may use these terms in different ways - that is, differently to the teacher, and differently from one another.

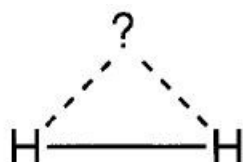


Figure 8.1 Not a chemical bond?

When this probe was undertaken by a class of 16-17 year olds similar responses were made. One student described the interaction shown in Figure 8.1 as 'covalent', and reported that 'the two hydrogens are attracted to each other, so they are bonded together'. This particular student classed this interaction as 'bonding', but not as a 'chemical bond': but did think the interaction was both an 'attraction' and a 'force'. A classmate who agreed that this interaction was 'bonding' was unsure if it was also a 'chemical bond', but did not think it was either an 'attraction' or a 'force'. Another student in the class thought that this 'covalent bonding' was both 'bonding' and a 'chemical bond', however this student recognised the interaction as an 'attraction', but not as a 'force'.

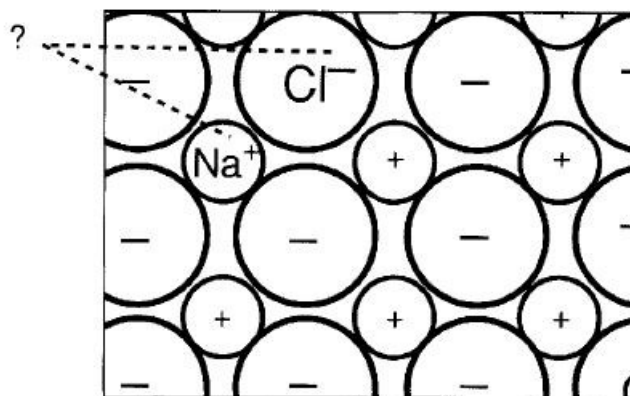


Figure 8.2 Not chemical bonding?

This same student described the interaction in 'part of a layer in a sodium chloride lattice' (Figure 8.2) as an 'attraction', but not a 'force', and neither as 'bonding' or a 'chemical bond'. A diagram showing 'iodine molecules in solid iodine' (Figure 8.3) was considered by this student to show 'van der Waal's [sic] forces' - interactions which were 'force', but were not classed as 'attraction'. Another classmate thought that the interactions in the sodium chloride were 'bonding' but not 'chemical bonds'.

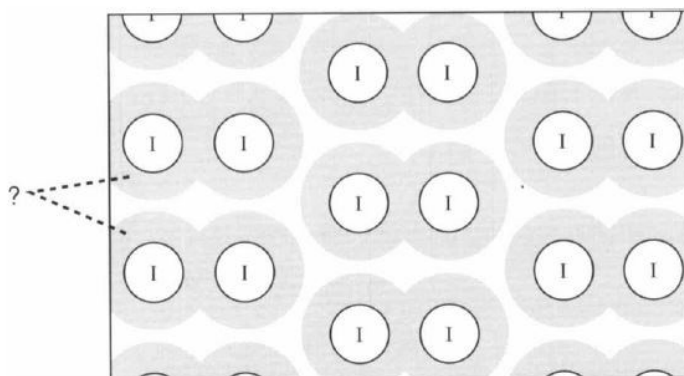


Figure 8.3 A force, but no attraction?

It is difficult to see any obvious pattern in these types of responses, except that where teachers may use terms such as 'attraction', 'force', 'bonding' and 'bond' without much thought (and often interchangeably), students are often drawing subtle and idiosyncratic distinctions.

Multiple models for chemical bonds

Chemical bonding is one of those topics where we use a variety of different models to understand different aspects of the phenomenon (see the discussion of models in Chapter 6).

For example, it is possible to teach about chemical bonds in terms of the electrical attractions between different species without mentioning orbitals - a topic that is often found difficult (see Chapter 7). Yet in post-16 level courses the idea of molecular orbitals, formed by the overlap and 'linear combination' of atomic orbitals is often introduced. A student who has learnt to conceptualise the covalent bond as a pair of electrons between, and attracted to, two nuclei, may find this new image of orbital overlap as something completely unrelated. Of course both of these 'pictures' are partial models of the same bond, but this will not be obvious to many students.

Again this is an area where helping students to appreciate that our descriptions and diagrams are just models will make learning easier.

Students' multiple conceptual frameworks for bonding

Just as teachers will use multiple models of bonding to help learners appreciate the abstract ideas involved, so students may develop manifold conceptions of chemical bonds. At some point, successful post-16 students are able to move beyond notions of bonds as shared electrons, to see bonds as electrical interactions. Those students who go on to study the subject at University level will be expected to master models of chemical reactions and bonding in terms of orbital interactions.

Some of the alternative conceptions described in this chapter are very common among students.

However, research suggests that the alternative ideas sometimes co-exist alongside developing more sophisticated understanding⁴ or example, it is often found that post-16 students are in transition between two models of the ionic bond (see Table 8.2). These students have two conceptual frameworks for making sense of the ionic bond - and may agree with (sometimes contradictory) statements that fit either framework.

The full shells explanatory principle may be a key part of a wider conceptual framework (the octet framework - see Chapter 1), where the different key ideas are mutually reinforcing. Unfortunately, conceptual change can be very slow, and shifts in the preferred bonding explanations many take many months.⁵ This is one area of the curriculum where students' alternative conceptions can be very tenacious indeed.

Student definitions of bonds: the bonding dichotomy

Research suggests that students at the end of secondary education commonly know about two separate categories of chemical bonding, and assume that chemical bonds must be like one of these two prototypes, (Table 8.1).⁶

Covalent	Ionic
Electrons are shared between non-metal atoms	Electrons are transferred from metal to non-metal atoms

Table 8.1 A dichotomy of bond types

Students seem to acquire this dichotomous classification of bonds readily, and when they do it means that they do not see bonding as primarily an electrical phenomena. Once this scheme has become established the student finds it difficult to appreciate bonding that is intermediate (polar bonds) or falls outside (eg hydrogen bonding) this narrow definition of bonding. Now clearly covalent and ionic bonds are very significant bond types, as many important substances can be understood to have - to a first approximation - either ionic or covalent bonding. However, the effect of pupils in school learning about bonding as a dichotomy of these two types, is to act as an impediment to later learning.

The covalent bond as electron sharing

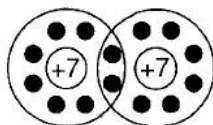


Figure 8.4 A molecule with a covalent bond

From an electrical perspective the bond in a fluorine molecule (eg Figure 8.4) comprises a pair of electrons situated between the positive cores of the atoms. This arrangement does not make sense from a purely electrical standpoint - as negative electrons would not be expected to be paired as they will repel each other.⁷ However few students question the 'pairing' of electrons, as they do not tend to see a covalent bond in electrical terms - rather they conceptualise the bond as a pair of electrons shared between two atoms to allow the atoms to have full electron shells (or octets).

Often, for the student, the bond is the sharing - and this is not necessarily meant figuratively (see Chapter 10). Students will report that atoms want to, and indeed need to, obtain full shells, and will therefore share to try to achieve this.⁸ The teacher may talk of a shared pair of electrons as a shorthand for the electrical interaction - but to many students 'sharing' electrons is a technical and not a metaphorical description of the bond. It is therefore not surprising that some students completing the Interactions probe will classify covalent bonds as not being 'attractions' or 'forces'. Sharing is a 'social' process not a physical one!

The polar bond

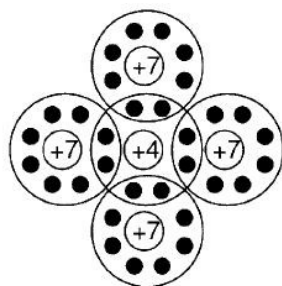


Figure 8.5 A molecule with polar bonds

If a student does conceptualise bonding in electrical terms then a polar bond can be seen as something in-between a covalent bond and an ionic bond. The electron pair is pulled closer to one atomic core (or the electron density is greater nearer that core).

Where students think about bonding in terms of the dichotomy, however, they will tend to describe a polar bond as a modified covalent bond, rather than something intermediate between covalent and ionic. However, unless the bond polarity is drawn to their attention, it is quite likely they will ignore it completely. A diagram such as Figure 8.5. is likely to be just labelled as covalent, because the way it is drawn (see Chapter 7) fits a description of 'electron sharing' better than 'electron transfer'.

This tendency to ignore bond polarity leads to other errors. For example, as students tend to classify hydrogen fluoride as covalent, rather than polar, they often describe the solvated species to be hydrogen fluoride molecules when it dissolves in water. Some students also tend to assume that bond fission will always be homolytic (with each atom 'getting its own electrons back'). Appreciating heterolytic bond fission is easier if a bond is conceptualised as polar, in terms of electrical interactions, rather than as covalent with electrons shared. Terms like 'sharing' bring associations from normal social interactions (of fairness and collaboration in this case), whereas heterolytic bond fission would be in breach of such a social contract.

One of the classroom resources included in the companion volume is a probe called Spot the bonding which presents learners with a range of diagrams of chemical systems and asks them to identify the type(s) of bonding present.

When this probe was piloted for the project by a cohort of 16-17 year olds, it was found that there was virtually no explicit recognition of the polar nature of any of the bonds represented (even when the presence of hydrogen bonding or dipole-dipole interactions were reported).

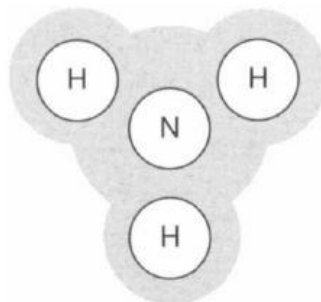


Figure 8.6 Representation of an ammonia molecule from Spot the bonding probe

For the ammonia molecule (Figure 8.6), 37 of the 39 students in the group identified covalent bonding - none suggested polar bonds (although there was one suggestion that the species had covalent bonding and 'permanent dipole-dipole forces').

In the case of the hydrogen fluoride molecule there were still no references to polar bonding, although 30 students suggested the bonding was covalent, and 4 suggested ionic. In the similar case of hydrogen chloride the split was 28 suggesting covalent and 7 suggesting ionic bonding. (Although chlorine is less electronegative than fluorine the type of representation - see Figure 8.8 - may have implied 'covalent' less strongly than in the case of hydrogen fluoride.) Even those students who suggested that there would be hydrogen bonding present (6/39), or dipole-dipole interactions (2/39), did not describe the bonding in the molecules as polar.

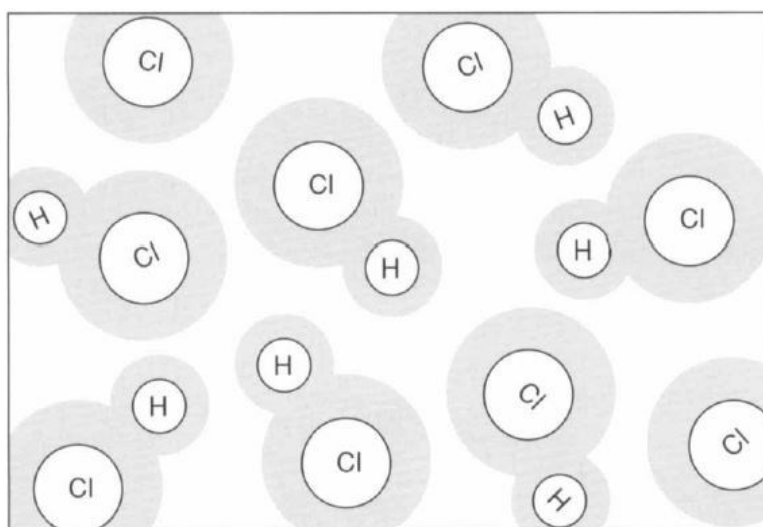


Figure 8.8 A representation of liquid hydrogen chloride from the Spot the bonding probe

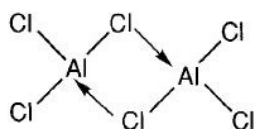


Figure 8.9 A representation of an aluminium chloride dimer from the Spot the bonding probe

Even in a case that students would find difficult to categorise according to the 'bonding dichotomy' they did not suggest the bonding might be polar. So in the case of aluminium chloride the bonds were drawn with lines (see Figure 8.9), in the way covalent bonds are often drawn, but the elements were a metal and a non-metal. 18 of the students suggested the bonding was covalent, and 11 suggested it was ionic.

Students are often taught that a dative bond is just like any other covalent bond, except that both electrons come from the same atom. This is meant to emphasise the important point that the origins of the electrons are irrelevant to their ability to bond together two atomic cores. Yet it can also obscure the fact that dative bonds are normally quite polar, with a very uneven 'sharing' of the electrons.

Bond polarity makes sense in terms of differences of electronegativity, for example in the hydrogen fluoride molecule -where the fluorine core charge is larger than the hydrogen core charge.

If the bond is conceptualised in electrical terms (the electron pair is attracted to and by both nuclei), then understanding bond polarity may be seen as developing or refining existing knowledge. Yet this is not so for learners who see the bond as a shared electron pair per se. Students tended to see bond polarity as an additional secondary characteristic of covalent bonds (rather than as a continuum running between covalent and ionic).

The ionic bond as electron transfer

Students very commonly have alternative conceptions of ionic bonding. Where students often come to see the covalent bond in terms of an inadequate image (electron sharing), they often define ionic bonding in terms of a completely irrelevant notion: electron transfer.⁹

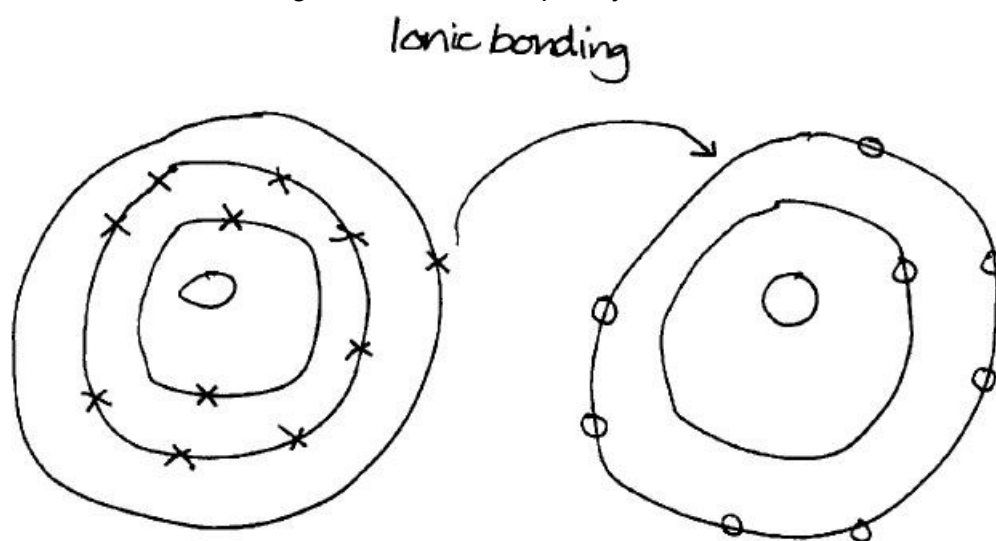


Figure 8.10 A student's diagram which has little to do with ionic bonding

Figure 8.1 0 is a student's representation of ionic bonding. This type of diagram is common, but – and I do not think this can not be emphasised too strongly - it has little to do with ionic bonding.

Diagrams like Figure 8.1 0 certainly do not show ionic bonding. Rather they represent a way in which ions might be formed from isolated atoms.

Such diagrams of 'ion formation' have very little to do with the way ionic bonds are likely to be formed in school science. If a pupil could start with isolated atoms of sodium and chlorine then such diagrams might represent how these species would interact: but sodium is usually a metal, and chlorine is usually molecular. So if sodium chloride was prepared by binary synthesis the reactants would not be in the form of discrete atoms.

In any case students are much more likely to form sodium chloride by a neutralisation process, followed by evaporation of water. In such a reaction the ions are already present in the solution, and their formation does not need to be 'explained' to discuss the ionic bonding in the product. (See the discussion of precipitation reactions in Chapter 9.)

Unfortunately for many students ionic bonding is defined as 'electron transfer' by the time they complete formal schooling, even though this is completely wrong. The strength of this conviction must reflect the way the topic is presented (by textbooks if not by teachers - see Chapter 10).

However, students seem to find it an easy image to learn as it fits with their wider understanding of the subject:

- that everything is 'made from' atoms;
- that bonds form to let atoms obtain full shells; and
- that when atoms bond, something like a molecule is formed.

If covalent bonding is taught before ionic bonding, then it is not surprising that diagrams of 'ion-pairs' are commonly seen to be like molecules, or even to be molecules. Indeed at the end of formal schooling students are as likely to hold an alternative 'molecular' framework for understanding ionic bonding as to hold the scientific view. (More precisely, many students have a confused mixture of these ideas). This alternative framework is compared with the scientific ('electrostatic') view in Table 8.2.

	Molecular framework	Electrostatic framework
Status	Alternative framework	Curricular science
Role of molecules	Ion-pairs are implied to act as molecules of an ionic substance.	Ionic structures do not contain ion pairs – there are no discrete ion-pairs in the lattice.
Focus	The electron transfer event through which ions may be formed.	The force between adjacent oppositely charged ions in the lattice.
Valency conjecture	Atomic electronic configuration determines the number of ionic bonds formed, (eg a sodium atom can only donate one electron, so it can only form an ionic bond to one chlorine atom).	The number of bonds formed depends on the co-ordination number, not the valency or ionic charge (eg the co-ordination is 6:6 in NaCl).
History conjecture	Bonds are only formed between atoms that donate/accept electrons, (eg in sodium chloride a chloride ion is bonded to the specific sodium ion that donated an electron to that particular anion, and vice versa).	Electrostatic forces depend on charge magnitudes and separations, not prior configurations of the system (eg in sodium chloride a chloride ion is bonded to six neighbouring sodium ions).
'Just forces' conjecture	Ions interact with the counter ions around them, but for those not ionically bonded these interactions are just forces, (eg in sodium chloride, a chloride ion is bonded to one sodium ion, and attracted to a further five sodium ions, but just by forces – not bonds).	A chemical bond is just the result of electrostatic forces – ionic bonds are nothing more than this (eg the forces between a chloride ion and each of the neighbouring sodium ions are equal).

Table 8.2 An alternative framework for ionic bonding¹⁰

One of the classroom resources included in the companion volume, Ionic bonding, allows teachers to test how strongly their students accept these two different frameworks for understanding ionic bonding. The probe asks student to judge the truth of statements relating to a diagram meant to show part of a layer of ions in a salt crystal (Figure 8.11). When this probe was administered to over 300 students in a range of schools and colleges it was found that 'alternative' statements were commonly judged true by students at the end of secondary schooling (who had studied the topic of ionic bonding), as well as by post-16 students before and after revisiting the topic during their chemistry classes.

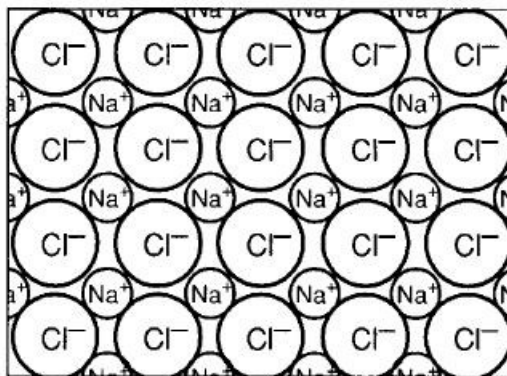


Figure 8.11 A focal diagram representing sodium chloride NaCl

For example the statement 'the reason a bond is formed between chloride ions and sodium ions is because an electron has been transferred between them' was considered correct by three quarters of respondents (69% of the sample of post-16 students who had revisited the topic, and 79% of each the other two groups). The statement that 'a chloride ion is only bonded to the sodium ion it accepted an electron from' was considered to be correct by almost half of the students (56% of the group at the end of schooling; 49% of the post-16 students relying on school knowledge; and 33% of the post-16 students having studied the topic at college level). Half the sample agreed that 'in the diagram a sodium ion is attracted to one chloride ion by a bond and is attracted to other chloride ions just by forces' (52%, 53%, 45% of the different groups respectively.) A statement that 'there are no molecules shown in the diagram' was only judged as true by a third of the sample (27%, 33% and 37% respectively).

Is the metallic bond ionic, covalent, or just damp at the edges?

Metallic bonding is not part of the 'bonding dichotomy' (see Table 8.1) commonly used by students to categorise bonds and is not usually studied in any depth before post-16 level courses. When students setting out on post-16 courses were asked about metals there was found to be four likely responses.¹¹

- there is no bonding in metals;
- there is some form of bonding in metals, but not 'proper' bonding;
- metals have covalent and/or ionic bonding; and
- metals have metallic bonding, which is a sea of electrons.

In other words, many students tended to try and find a fit between the obvious structural integrity of metals, and their simple dichotomy of what bonding must be. For many of these students chemical bonding was understood in terms of striving to obtain a full outer shell by sharing (covalent bonding) or transferring (ionic bonding) electrons. As with ionic bonding, some students assume metals are molecular, and may try to think in terms of valency to

make sense of diagrams they have seen in books (see Figure 8.1 2, which represents one college student's hybrid model based around the number of covalent bonds iron was considered to form).

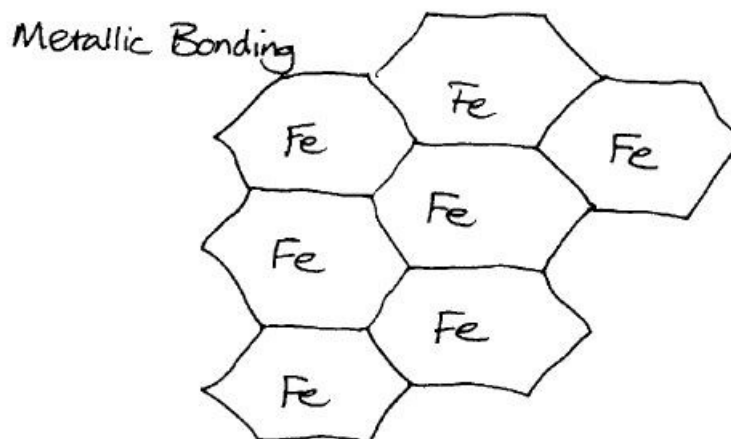


Figure 8.12 Metal molecules in iron (student diagram)

Students who are unable to make sense of metals in terms of either pattern may be at a loss to explain how metals can have bonds: leading to comments such as 'it's not ionic, and it's not covalent either, it's like, it's hard to explain this'. These students may conclude that there is no bonding in pure metals, or that there is a 'lesser' form of bonding -just a force, and not a real chemical bond.

'Ionic and covalent bonds are formed ... where atoms lose or gain electrons, or share them, whereas metallic bonding is not the sharing or loss or gain of electrons. It is just a loose association with metal ions, and electrons they have lost.¹²

As discussed above, students often do not see bonds as being forces, and so may logically consider a structure to be held together just by forces, rather than by bonding.

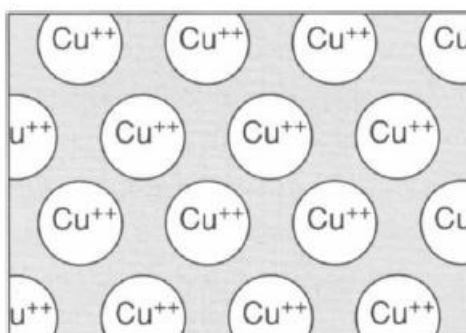


Figure 8.13 A representation of copper in the Interactions probe

The Interactions probe included in the companion volume includes an item about copper. One student who labelled the interaction present as 'metallic bonding' thought that this counted as both an 'attraction' and a 'force'; but was unsure if it could be considered as 'bonding', and did not class it as a 'chemical bond'.

Other students find ways to understand the metallic bond as a variation on the ionic or covalent case, with electrons being conceptualised as being shared, or being moved around so that the atoms take turns in having full shells (either by gaining enough, or losing enough electrons). A more appropriate model sees the electrons as having been transferred to the lattice so that the atoms could gain a full shell.

Although some students will describe the bonding in metals in terms of the 'sea' of electrons, they have often learnt the term by rote, with little understanding of this model (see Figure 8.14). Some student diagrams show the 'sea' as a vast excess of electrons (Figure 8.15), and one student who grasped that metallic bonding was 'the attraction between the +ve charge of the metal ions and the -ve charge of the electrons' went on to add that 'it also has a sea of electrons which flow around the structure.'

Like the case with the term 'sharing', the sea metaphor brings its own associations, with students referring to the metal cation 'being like an island surrounded by electrons' and 'floating in a sea of delocalised electrons'. The use of metaphors in teaching science concepts is considered in Chapter 10.

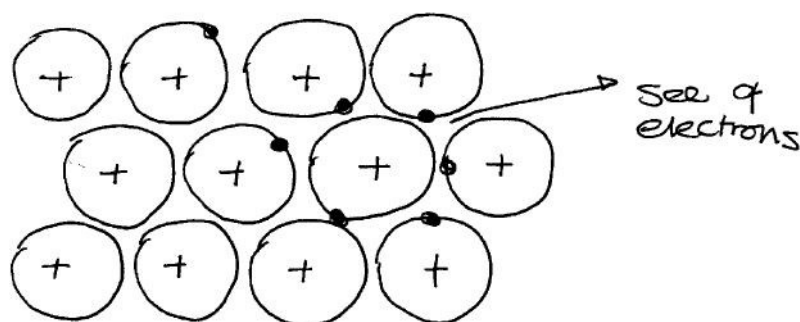


Figure 8.14 Can you 'see' the electrons? (student diagram)

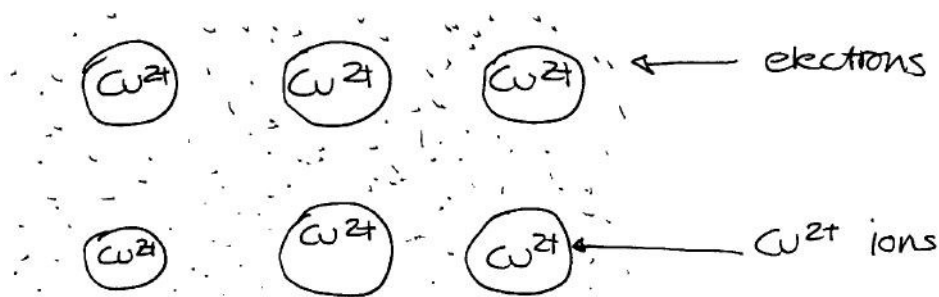


Figure 8.15 An electronegative metal? (student diagram)

Included in the companion volume is a resource, Iron - a metal, which allows teachers to diagnose some of the common alternative conceptions that students may have about metals and metallic bonding. Reference was made in Chapter 6 to how students in one group of 14-15 year olds confused the properties of the metal with properties of metal atoms. Other items in the probe allow teachers to diagnose students' ideas about the bonding in a metal. Half (9/18) of the same group of 14-15 year olds thought that 'the atoms [sic] in a metal such as iron are held together by ionic bonds'.

Intermolecular interactions as bonds, or just forces?

If many students have difficulty thinking of metallic bonding as proper chemical bonding, then the status given to intermolecular bonding tends to be even less firm. It is common for these interactions to be considered as 'just forces' rather than as 'proper' bonding. Although this may sometimes be a distinction based upon the strength of the interaction (eg van der Waals forces are too weak to be considered 'proper' bonds), this is not a full explanation when the interactions in an ionic lattice are often divided into 'ionic bonds' (resulting from electron transfer) and 'just forces' (due to electrical attraction between ions). For many students the criterion for a bond is not related to the strength of the interaction, but to whether the bond results in atoms obtaining full shells.

The full shells explanatory principle for explaining chemical bonding, that is that bonds are formed 'to try to achieve a stable structure ie 8 electrons in the outer shell of the atom' is a very common alternative conception. Although it may be expressed in various ways ('to give them all full outer shells', 'to obtain the configuration of a noble gas, and be stable', 'to form a completed outer shell'), the underlying idea is virtually ubiquitous. If a student understands bonding purely in these terms then forms of intermolecular bonding such as van der Waals forces and hydrogen bonding will not qualify, and so will be classed as 'just forces'.

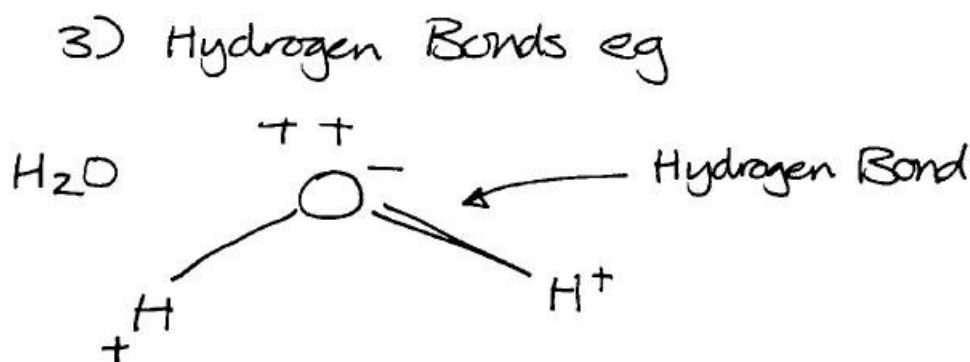


Figure 8.16 A bond to hydrogen (student diagram)

Hydrogen bonds may be misunderstood by students as simply being 'bonds to hydrogen' (see Figure 8.1 6) and just thought of as covalent bonds. This sometimes occurs when students are introduced to hydrogen bonding, without further elucidation, in biology before they have met the concept in chemistry (see Chapter 10).¹³

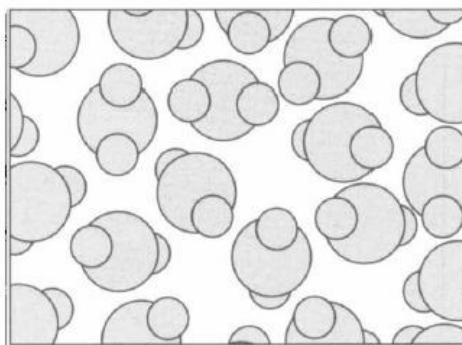


Figure 8.17 A representation of molecules in liquid water, from the Interactions probe

When students realise that this is not what chemists mean by hydrogen bonds they may often be demoted from being real bonds. Although hydrogen bonds are electrical interactions (and involve orbital interactions) they do not allow atoms to obtain full shells, which is the most common criterion used by students to characterise a chemical bond. For example, a student responding to the Interactions probe included in the companion volume labelled the interactions between molecules in water (Figure 8.1 7) as 'hydrogen bonding'. The student described how:

'The highly electronegative oxygen in H₂O attracts each hydrogen's electron so strongly that we basically have an exposed H nucleus in the H₂O molecules giving the H a permanent δ^+ and the Os a permanent δ^- charge. The neighbouring molecule H δ^+ are attracted to the O δ^- thus holding the molecules together in the liquid.'

This student certainly appreciated the electrical nature of the interaction, but did not think that this amounted to a chemical bond.

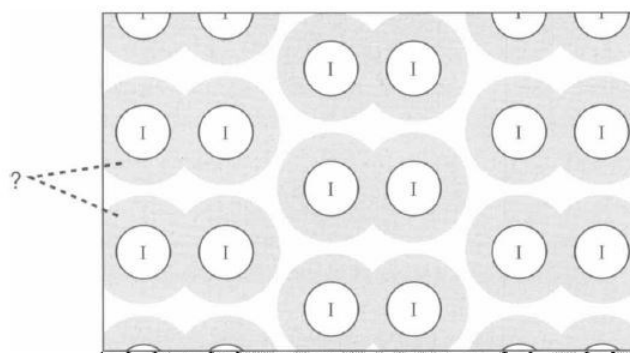


Figure 8.18 A representation of molecules in solid iodine, from the Interactions probe

In a similar way a student who recognised 'van der Waals forces' in Figure 8.18, where 'molecules [are] attracted by induced dipoles which are instantaneous and attract molecules together' thought this did not count as 'bonding' nor as a 'chemical bond'. The same student described the 'electrostatic attraction' that gave rise to solvation (Figure 8.1 9) where 'oxygen in water attracts Ag⁺ and H in water attracts NO₃⁻', but again this was not categorised as 'bonding' or as being a 'chemical bond'.

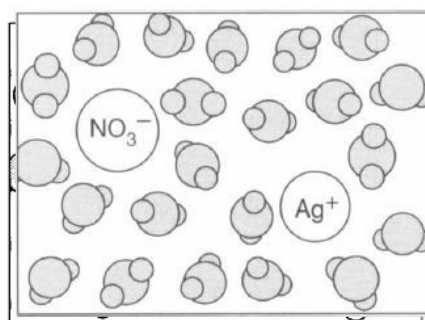


Figure 8.19 A representation of solvation from the Interactions probe

In these cases the students seemed to have a fair understanding of the interactions, even if they did not consider them as bonds. Indeed it may well be the label 'bond' which stands in the place of understanding for some students -the same individual who was able to describe van der Waals forces in iodine in terms of induced dipoles and solvation in terms of attractions, described the bond in a hydrogen molecule as a 'shared pair of electrons

between the two hydrogen atoms in a covalent bond', without any reference to what made this hold the molecule together.

However, the students who completed the Spot the bonding probe often failed to spot intermolecular bonding in the diagrams represented. Only 14/39 students suggested there would be any type of intermolecular bonding in solid iodine; 15/39 in liquid oxygen; 15/39 in liquid hydrogen chloride and 16/39 in sodium nitrate solution.

Developing understanding of chemical bonding

So to summarise, many students define bonding as a way of getting full shells (by electron sharing or electron transfer). Students who think this way may tend to ignore other types of bonding which do not fit this scheme (polar bonds, metallic bonds, intermolecular bonds), and may be slow to move beyond inadequate definitions of covalent and ionic bonding. However, this does not mean that students can not move beyond their limited models of the bond.

The octet framework (see Chapter 1) is a substantive learning block (see Chapter 4) to many students, but is not necessarily insurmountable. Students will use multiple conceptual frameworks, as discussed earlier in this chapter, and successful post-16 students do slowly learn to adopt models of chemical bonds in terms of electrical forces and even orbital interactions, and can be gradually weaned away from relying on the full shells principle.

Some of this may happen without any deliberate intervention by teachers. For example, students who had studied Raoult's law and had been taught that solutions deviated from ideal behaviour when the mixture had different types of bonding to the pure liquids were found to refer to intermolecular bonding and solute-solvent interactions as being bonds in that context.

As students often have a poor appreciation of the nature of models they may need to be encouraged to accept that they can use alternative types of explanations. One way of doing this is to encourage students to see their chemical ideas as components of a mental toolbox from which they need to select the best tool to tack a particular 'job' (explanation, examination question etc)¹⁴

Students may feel more comfortable with acquiring and trying out manifold conceptions if they see explanations as being like stories -where they must find the best fit between the phenomena and the explanation.

Figure 8.20 represent this figuratively. This diagram uses a 'profile' to show the extent to which different types of bonding explanations were given when a student on a two year post 16 chemistry course was asked about aspects of chemical bonding.^{16,17,18}

In this diagram the 'target' is an understanding of chemical bonding based around electrical and some orbital ideas, but the student commences the course habitually using the idea of 'full shells' as the main way of discussing bonds.

During the course this student developed a new principle for bonding (that 'bonds form to minimise energy') as well as the intended ideas. By the end of the course the student had moved some way towards the way of thinking that the teacher was trying to encourage, but still commonly used 'full shell' explanations, which were so well established.

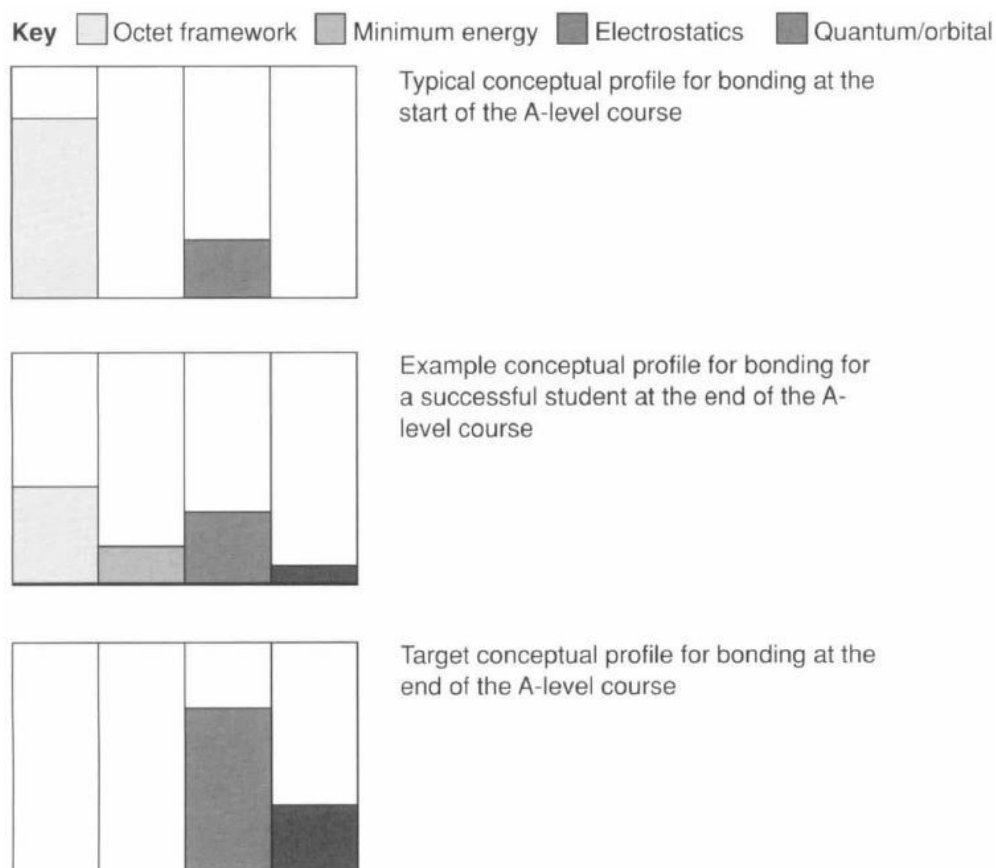


Figure 8.20 Progression in understanding bonding during post-16 chemistry¹⁵

It is worth noting that in Figure 8.20 the 'minimum energy' principle is shown as a discrete type of explanation. From a scientific point of view the idea that bonds form as chemical systems evolve to a lower energy is closely related to the idea that electrical forces act to give chemical structures at equilibrium. The 'lowest energy' point is the equilibrium configuration where the forces are balanced.

However, for the student who is profiled in the diagram, the 'minimum energy principle' was a completely separate idea. Once again a student failed to appreciate what the teacher took for granted, and so - for this learner - explanations in terms of minimising energy, and those in terms of electrical forces, were stored and accessed as separate stories for explaining bonds. This fragmentation of knowledge (see Chapter 4) was an impediment to the student developing a deeper understanding, but given more time and more experience of using his different stories it is quite possible this student may have come to integrate these different accounts. Teachers who wish to encourage their students to develop their ideas about chemical bonding should:

- emphasise the nature of bonds as electrical interactions;
- avoid using anthropomorphic language, but rather explain bonding in terms of forces; and
- avoid talking about electron transfer (ion formation) when considering ionic bonding.

Notes and references for Chapter 8

1. K. S. Taber, An alternative conceptual framework from chemistry education, *International Journal of Science Education*, 1998, **20** (5), 597-608.
2. Or, to be precise, it is an equilibrium of forces, which holds the chemical system in balance; and which means that work has to be done to move the system from that equilibrium.
3. Students' spelling of van der Waals can be somewhat variable. In the responses one group made to the **Spot the bonding** probe the variations presented included: Van der waals forces, Vanderwaals, Wan der Waals' forces, Wam der Vaals, van der walls and van der balls forces.
4. K. S. Taber, Multiple frameworks?: Evidence of manifold conceptions in individual cognitive structure, *International journal of Science Education*, 2000, **22** (4), 399-417.
5. K. S. Taber, Shifting sands: a case study of conceptual development as competition between alternative conceptions, *International journal of Science Education*, 2001, **23** (7), 731-753.
6. K. S. Taber, An alternative conceptual framework from chemistry education, *International journal of Science Education*, 1998, **20** (5), 597-608.
7. The concept of 'quantum mechanical spin' is not normally introduced in secondary level chemistry.
8. K. S. Taber & M. Watts, The secret life of the chemical bond: students' anthropomorphic and animistic references to bonding, *The International journal of Science Education*, 1996, **18** (5), 557-568.
9. K. S. Taber, Misunderstanding the ionic bond, *Education in Chemistry*, 1994, **31** (4), 100-103.
10. K. S. Taber, Student understanding of ionic bonding: molecular versus electrostatic thinking?, *School Science Review*, 1997, **78** (285), 85-95.
11. 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).
12. K. S. Taber, Understanding Chemical Bonding, unpublished PhD thesis, University of Surrey, 1997, 354.
13. K. S. Taber, Understanding Chemical Bonding, unpublished PhD thesis, University of Surrey, 1997, 838-839.
14. K. S. Taber, An analogy for discussing progression in learning chemistry, *School Science Review*, 1995, **76** (276), 91-95.
15. This figure is taken from K. S. Taber, An explanatory model for conceptual development during A level chemistry, 1999, available via Education-line, at <http://www.leeds.ac.uk/educol/> (accessed October 2001).
16. This is an approach based on the ideas of the French philosopher of science, Gaston Bachelard (see note 17) and proposed as a useful technique to follow conceptual development (see note 18).
17. G. Bachelard, *The Philosophy of Science: A Philosophy of the Scientific Mind*, New York: Orion Press, 1968.
18. E. F. Mortimer, Conceptual change or conceptual profile change?, *Science and Education*, 1995, **4**, 267-285.