

This is an author's manuscript version.

The version of record is:

Taber, K. S., & Coll, R. K. (2002). Chemical bonding. In J. K. Gilbert, O. de Jong, R. Justi, D. F. Treagust, & J. H. Van Driel (Eds.), **Chemical Education: Research-based Practice** (pp. 213-234). Dordrecht: Kluwer Academic Publishers BV.

Chemical Bonding

Introduction

Chemical bonding is a key concept in chemistry. It is also a topic area where understanding is developed through diverse models - which are in turn built upon a range of physical principles - and where learners are expected to interpret a disparate range of symbolic representations standing for chemical bonds.

Chemistry, as a subject, may be said to be concerned with the properties and reactions of substances. Substances are often understood in terms of aggregations (or combinations) of particles, and the nature of the bonding between those particles is used to explain many of the chemical and the physical properties of the substance - including such gross aspects as whether the substance is a solid, liquid or gas at a given temperature and pressure. Chemical change involves the rearrangement of the constituent particles that make up the reactants, to give new configurations that are characteristic of (and, indeed, determine) the products. In other words, chemical reactions involve the breaking and forming of chemical bonds. The relative bond strengths in reactants and products are also a key determinant of the thermodynamics of a reaction process, and therefore of the equilibrium position in practice, for many reactions, this means whether or not the reaction will proceed.

The central importance of chemical bonding to the subject is well recognised (Fensham, 1975), and is such that it is considered a core topic in many chemistry curricula at school, college and

university level. However, it is also a topic where learners commonly develop a wide range of alternative conceptions (Taber 2001a).

Learners come to science classes holding a wide range of alternatives to established scientific ideas (e.g., Pfundt & Duit, 1998; Driver et al., 1994), which may act as impediments to the intended learning. In some topics within science, these alternative ideas may be understood to originate largely from 'intuitive' interpretations of early experience (e.g., Gilbert & Zylbersztajn, 1985), or to be acquired from folk-science (e.g., Solomon, 1993) and everyday meanings of technical words (e.g., Watts & Gilbert, 1983). Common alternative conceptions in such topics as forces and motion, plant nutrition and energy may be understood to derive from such sources. However, these types of explanations are less convincing when considering how learners develop alternative ideas about something as abstract as chemical bonding - which is not within their direct experience, nor something they will hear discussed in everyday discourse.

The range and sophistication of the scientific models used by chemists to make sense of chemical bonding is one factor which contributes to learners finding this topic difficult. Many of the ideas used to understand chemical bonds would not be accessible at an introductory level. Instead, curricula models need to be used which simplify the topic (Gilbert, 1998). Ideally learners will develop their 'tool kit' of bonding concepts as part of their progression in learning about the subject (Taber, 1995).

When teaching models are developed they should reflect an optimum level of simplification (Taber, 2000), that is they should be kept as simple as is possible whilst still being scientifically 'authentic' (so that they provide a suitable basis for being developed at a later stage in the learner's chemical education). Where attempts to simplify a complex subject are *not* planned so as to provide for such progression, then they can act as impediments to effective learning, and may contribute to the development of some common alternative conceptions. Some of the alternative conceptions that learners commonly acquire about chemical bonding may be understood to arise in this way.

Structure of the chapter

Learners' alternative conceptions which are considered to largely derive from the way they have been taught have been labelled as pedagogic learning impediments (Taber, 2001b). This chapter is organised around four such pedagogic learning impediments relating to chemical bonding:

- learners hold an incorrect and inappropriate rationale for *why* bonding should occur;
- learners see all bonding as involving discrete molecules, and do not understand the nature of ionic and metallic bonding and of giant covalent structures;
- learners may discount from the category of 'bonding' anything which does not seem to fit the description of 'electron sharing' or 'electron transfer';
- learners may be unable to make sense of intermediate bond types (e.g., polar bonding).

A key research finding which provides an overarching perspective on these learning impediments is the claim that learners' thinking about bonding is often influenced by a common alternative conceptual framework which has been labelled the 'octet' framework. We therefore begin the chapter by describing the nature of this alternative framework, and considering why it seems to be so widely adopted. In particular, we consider the tendency of students to focus on separate atoms when thinking about chemical reactions, and the common use of anthropomorphic language to discuss chemical processes. The four pedagogic learning impediments listed above are then considered in terms of the following principles:

- The chemical bond is due to electrical forces;
- Bonding need not imply molecules;
- Not all chemical bonds are covalent or ionic;
- Bonding may be intermediate between covalent and ionic.

In considering each of these five themes, relevant literature on learners' alternative conceptions about bonding is reviewed, and the teaching models and approaches considered to encourage such ideas are considered. The chapter considers recommendations for the way in which the topic should be taught to minimise the incidence of these alternative conceptions, and to encourage learners to develop models of chemical bonding which are more authentic, and which consequently have more explanatory power.

The Octet Framework

Research has shown that by age 16 students commonly adopt as an explanatory principle the notion that atoms *want* to have ‘octets’ or ‘full outer shells’, and that chemical processes often occur to allow atoms to achieve this. (It should be noted that students, and some school textbooks, incorrectly refer to eight electrons in the third or higher shells as a full shell.) In other words, reactions are said to occur to allow atoms to obtain octets of electrons by forming bonds (Taber, 1998).

An atomic ontology

The octet framework may not seem a likely candidate for an explanatory principle to rationalise bonding, as all the chemical reactions met at an elementary level in school science have *both* products *and* reactants which have ‘satisfied’ the ‘need’ for full shells or octets. The fecundity of the octet rule depends upon a particular, but widely found, mind-set (Taber, 2001a). The notion that ‘everything is made from atoms’ seems to be accepted by learners as a creed, and is taken to have a much more encompassing meaning than is justified. In practice, very few common materials contain anything approximating separate atoms, and the few that do are of limited chemical significance. Yet learners seem to acquire an ‘atomic ontology’ where all molecular level species are considered to derive from and comprise of atoms. Learners tend to think of the starting materials of chemical processes as being single unbound atoms, even though this is hardly ever the case.

Anthropomorphic explanations

The octet framework not only draws upon this ‘assumption of initial atomicity’, but also relies upon the way learners may make sense of abstract chemical processes by using anthropomorphic language (Coll & Taylor, 2001a, 2001b; Taber, 1998; Taber & Watts, 2000). Learners find that they can use the language they apply to psychological states (‘needs’) and social situations (‘sharing’) to their

chemistry. Atoms are said to *own, donate, accept* and to *share* electrons. They are said to *desire, need* or *want* electrons, and to be *happy* when they *think* they have full shells or octets of electrons. Sometimes even more extreme language (such as talk of atoms being *jealous*) may be used without seeming incongruent to the learner (Taber & Watts, 1996). Although at one level the students are aware that the atoms are not living conscious entities, they may rely on the analogy between physical atomic interactions and human social interactions to make sense of the chemical models. It has been found that younger pupils may even suggest that atoms *are* alive, perhaps because they seem to move under their own volition (Griffiths & Preston, 1992; Wightman et al., 1986).

The term 'sharing' used to describe the covalent bond often keeps its *social* connotations when used by students. For example, the shared electrons may be seen to still 'belong' to specific atoms, and so bond fission is often assumed to be homolytic - as each atom would *want* to get 'its own' electron back (Taber, 1998).

Taber and Watts (1996) discussed the way such anthropomorphic language is used so readily in these contexts and tentatively suggested that it should be indulged while students were becoming familiar with the scientific particle models of matter. Such ways of conceptualising molecular interactions could be useful in the initial stages of learning about this unfamiliar and abstract world, thereby enabling learners to make sense of this new realm. The social metaphor provides a way of 'anchoring' the new learning on the bed-rock of established knowledge (Taber, 2002).

However, anthropomorphic explanations are at odds with the fundamental physical concepts which explain the existence of chemical bonds, so *unless* learners used this way of thinking as a platform for developing a more physical understanding of molecular processes, such anthropomorphic ways of talking about atoms and molecules could become 'fossilised', and so become a barrier to further learning (Taber & Watts, 1996). Indeed, it could be argued that the uncritical and unthinking use of terms like 'sharing electrons' by teachers and in text books are not helpful to learners.

The source of the alternative conceptual framework

A number of factors combine to produce this alternative conceptual framework (see for example, Taber, 2001a). For one thing there is little discussion at secondary school level of *why* chemical

reactions do occur - as this is felt to be too difficult a topic. This leads to an 'explanatory vacuum', which is perhaps especially potent as the whole style of presentation of abstract theoretical chemical concepts gives students a tacit (and reasonable) expectation that the theory exists to explain the phenomena being studied. A related second feature is the presence of a feasible option for an explanatory principle. The focus on electronic configurations, and the high profile use of the octet rule (to determine likely formulae for molecular and ionic materials), provides a suitable candidate.

If the learner makes an 'assumption of initial atomicity' about reactants then the octet rule can seem to provide a rationale for reactions to occur. Such an assumption is encouraged by text book (or teacher) diagrams purporting to represent genuine chemical reactions, but showing interactions between discrete atoms (Taber, 2002). So, for example, four separate atoms of hydrogen and one of carbon may be shown to make methane; and isolated atoms of sodium and chlorine may be shown as forming a pair of ions which represents sodium chloride. The failure to represent the reactant molecules or lattice structures concerned is a simplification which encourages students to develop alternative conceptions.

The emphasis in introductory secondary school chemistry courses on the octet rule, the use of chemically unrealistic diagrams for bond formation, anthropomorphic descriptions of chemical processes, and the common lack of any discussion of why reactions do occur all contribute to the conceptual landscape that is fertile ground for the development of the common view that *atoms form bonds in order to obtain full shells of electrons* (Taber, 2001c).

The chemical bond is due to electrical forces

While electrical forces cannot be used to explain all aspects of chemical bonding (e.g. the influence of the 'spin' of the electron does not neatly fit into this scheme), they do provide a sound basis for starting to make sense of bonding phenomena. Chemists use more sophisticated models drawing upon quantum mechanics (Ogilvie, 1990), but an authentic teaching model used to introduce chemical bonding, at an optimum level of simplification, would be based upon the effect of electrical forces (Taber, 2002).

Lack of Appreciation of the Scientific Perspective

Yet there is research evidence from a number of studies around the world that learners often do not appreciate this scientific perspective, whether at secondary school (Wightman et al., 1986), high school/college (Griffiths & Preston, 1992; Harrison & Treagust, 2000; Taber, 2001a) or even at undergraduate level (Cros et al., 1986). Indeed, the younger learners may imagine bonds to be very small springs or lengths of string, and Butts and Smith (1987), for example, suggest that the ubiquitous use of ball-and-stick models may reinforce such notions.

Boo (1998) suggests that a belief that the chemical bond is a physical entity arises from a world-view that building a structure requires energy input, whereas destruction involves release of energy - that is, learners believed that bond breaking releases energy and bond making involves energy input (c.f. Hapkiewicz, 1991). However, this tendency - to see bonds as just as substantive as the materials which they are supposed to hold together - may also be seen as part of a more general problem of students misunderstanding the way in which scientists use the particle theory of matter to represent and to explain macroscopic phenomena.

Shifts Between Levels of Representation

Learners do not easily follow shifts between the macroscopic and molecular levels (Lijnse et al., 1990; Jensen, 1995; Johnstone, 1991; Selley, 1978; Tsaparlis, 1997). Where the scientist invokes a *molecular level feature* (such as the electrical interactions between parts of a molecule) to explain a macroscopic property, the learner may often simply transfer the feature to be explained (hardness, melting temperature, lustre etc.) to the molecular species (Taber, 2001a). There are many examples of this sort of thinking in the literature, such as the belief that atoms in a metal are hard, but those in liquids are softer (Harrison & Treagust, 1996); or that copper is malleable because it has malleable atoms (Ben-Zvi et al., 1986).

The covalent bond as a shared electron pair

By the time learners reach 16 years of age they have usually mastered a teaching model of molecules having covalent bonds, which are seen as pairs of electrons 'shared' by two atoms. However, it has been found that students commonly have difficulties progressing *beyond* the notion of the shared pair of electrons: which is seen as something more than an image or metaphor (Taber, 2001a). For many learners the shared electron pair *is* the bond, a notion which is somewhat lacking in explanatory power (Tsapralis, 1984, c.f. Taber & Watts, 2000), and which does not provide a good basis for progression (see, for example, the comments below on polar bonding).

Taber (1998) reports that the explanatory principle that atoms form bonds in order to achieve full shells or octets is very common among 16-18 year old students - apparently replacing the younger students' ideas about string, glue and elastic. Molecules are usually assumed to arise from discrete atoms because the atoms *want* or *need* to obtain 'full outer shells' or octets (see the section above on the Octet framework). For many students a shared electron pair holds atoms together *because* it enables them to have octets of electrons, and Barker and Millar (2000) have reported that the notion of *valency* (which is seen by students as the number of bonds 'needed' to obtain octet structures) is imbued with an explanatory or causative power by some learners.

The existence of bonding which does *not* lead to atoms having full electron shells is consequently something of a mystery to many learners. These students are not able to understand - for example - why sulfur would 'want' to go beyond SCl_2 or SF_2 to give SCl_4 or SF_6 , or why the chlorine atom in AlCl_3 would 'want' to share an electron pair to form a dative bond, when it already had all the electrons that it 'needed' (Taber, 2001c).

Bonding need not imply molecules

From the scientific perspective, some materials with covalent bonding will exist in the form of molecules, but others may have extensive covalently bound lattices. Metals, and salts, are bonded, but *do not* consist of molecules. Yet research suggests that learners do not readily appreciate such structural diversity: just as they tend to assume elements in reactions are present as separate

atoms (see ‘an atomic ontology’, above), many also tend to conceptualise bonded materials as *always* being in the form of molecules.

If understanding the nature of the covalent bond in discrete molecules - the context of chemical bonding most familiar to students - can be problematic, then appreciating lattice structures is often more so (Coll & Taylor, 2001a, 2001b; Coll & Treagust, 2001; Taber, 1997, 2001a). According to one French study of students entering university to study science, the structure of crystals was “a mystery for most” (Cros et al., 1986, p.309). After one year of university study the interactions within the crystal were “somewhat less mysterious” to the students, but less than a fifth of the sample referred to electrostatic interactions (Cros et al., 1988, p.334).

A highly prevalent alternative conception for chemical bonding at secondary school and (to a lesser extent) the tertiary level is that continuous covalent or ionic lattices contain *molecular* species (Birk & Kurtz, 1999; Butts & Smith, 1987; de Posada, 1997; Peterson et al., 1989; Taber, 1994, 1998). Other research has revealed that some learners believe ionic substances such as sodium chloride possess covalent bonds (Peterson et al., 1989).

Although there are a limited number of familiar substances which have giant covalent lattices, this is an important *type* of structure. Learners may consider this type of material to contain discrete molecules with strong *inter-molecular* forces (Tan & Treagust, 1999). Students may visualise solid carbon as comprising of discrete atoms, something that may in part derive from the ‘molecular’ formula of carbon being commonly given as ‘C’ and taken to imply ‘C₁’, when ‘C’ might be more appropriate (Taber 2002, c.f. Nelson 1996).

A molecular model for ionic bonding

As suggested above, learners often appear to have little appreciation of the underlying electrostatic nature of chemical bonding (Boo, 1998; de Posada, 1997). For example, attraction between two oppositely charged species may be thought to result in neutralisation rather than bond formation - the likely source of confusion being the parallel with acid-base chemistry (Boo, 1998; Schmidt, 1997).

Year 12 Australian students asked about sodium chloride were found to often volunteer a description of how ions might be formed through an electron transfer event (i.e. from sodium atom to chlorine atom). It was also common for these students to refer to molecules of NaCl, and some believed that there were two types of bond in sodium chloride: either that the 'NaCl molecules' had internal covalent bonds, but were ionically bonded to other molecules, or vice versa (Butts & Smith, 1987). Some of the students thought that this assumed molecular nature of sodium chloride explained why the solid did not conduct electricity, believing that ions were only formed from the molecules on dissolving. One student thought that each ion would have one ionic bond, and five "physical" bonds (p. 196).

Similar findings have been reported from 16-19 year old U.K. students (Taber, 1994). It was common for these students to consider NaCl ion-pairs within the lattice *as if* they were molecules (with some actually using this term). Although these U.K. students did not tend to expect covalent bonding within the ionic lattice they often distinguished between two types of interaction: ionic bonding within the ions pairs, and 'just forces' between them. The idea of ionic materials containing molecules seems to be quite widespread among this age group according to other studies from the U.K. (Barker & Millar, 2000), Australia (Harrison & Treagust, 2000), Singapore (Tan & Treagust, 1999) and New Zealand (Coll & Treagust, 2001).

The studies of Australian and English students suggest that many think that each ion in the sodium chloride lattice could only form one bond. This relates to the 'full shells' explanatory principle that students use to explain bonding: a sodium atom is considered to 'need' to donate one electron (and therefore forms one bond) and a chlorine atom is seen to 'need' to gain one electron (forming one bond). The bond would therefore (according to students) only exist between ions that had transferred electrons. From this explanation students obviously get the idea that a single bond is formed – not a network or a lattice (Taber, 1997, 1998).

There is some suggestion that these common alternative conceptions may survive university teaching, for example, among New Zealand chemistry graduates (see Coll & Treagust, 2001). In the U.K., Oversby (1996) found that some of his post-graduate trainee chemistry teachers considered the alternative conceptions of the ionic bond reported by Taber (1994) to be an acceptable model of ionic bonding, i.e.:

- that ions could only bond to the number of counter ions allowed by their valency;
- that a bond only exists where an electron has been transferred; and
- that where electron transfer had not occurred ions were held together by forces but not bonds.

The common practice of teaching about covalent bonding and covalency *before* ionic bonding, and of representing ionic bonding in terms of (chemically unlikely) diagrams showing electron transfer (i.e. ion formation, *not* bonding) often leads to students conceptualising electrovalency as determining the number of bonds formed, and seeing ionic materials as quasi-molecular, so that in sodium chloride there are seen to be discrete NaCl units which are internally bonded, and which are sometimes considered to be the solvated particles on dissolving (Taber, 2001a).

Not all chemical bonds are covalent or ionic

From the scientific perspective there are a number of important bonding types besides covalent and ionic bonding, such as metallic bonding, and hydrogen bonding. To the chemist, the relative strengths of different types of bonds is important, but even 'weak' bonds have significant chemical consequences. As described previously, research suggests that learners may often only consider the covalent and ionic cases to represent 'proper' chemical bonds, as opposed to 'just forces'.

Metallic Bonding

Taber (2001a) reported that U.K. students starting out on a college (equivalent to senior high school) course tended to present four notions of the bonding in metals:

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

For many of these students, chemical bonding was understood in terms of striving to obtain a full outer shell, *either* by sharing (covalent bonding) *or* transferring (ionic bonding) electrons. As with

the ionic bonding case, some students assume metals are molecular (see also Fleming, 1994; de Posada, 1997; Harrison & Treagust, 2000). Students who were unable to make sense of metals in terms of either pattern might conclude that there is no bonding in pure metals, or that there is a 'lesser' form of bonding - something also reported by Coll and Taylor (2001a) - that was just a force, and not a real 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); or - more acceptably perhaps - the electrons were considered to have been transferred to the lattice so that the atoms could gain a full shell.

Taber (2001a) found that some students would describe the bonding in metals in terms of a 'sea of electrons': but that often they had learnt the term with little understanding of this model. Some students were so strongly influenced by the sea metaphor that they drew diagrams showing the 'sea' as a vast excess of electrons - giving a structure that would be highly charged and unstable.

Intermolecular Bonding

Some students do not appreciate the nature of a molecular solid, where discrete molecules are held in lattice positions by intermolecular forces, which are weaker than the intramolecular bonding (Butts & Smith, 1987). Students may suggest that a grain of sugar is a single molecule, and has a giant structure like diamond. This way of thinking can lead to the common ideas that the covalent bonds break on change of state (Tan & Treagust, 1999), and that (because molecular covalent materials usually have low boiling temperatures) covalent bonds must be relatively weak (Barker & Miller, 2000).

Students may have difficulty accepting anything that is not clearly explicable in 'octet' terms as being a chemical bond. Hydrogen bonding, and van der Waals' forces cannot be readily fitted in such a scheme, and the difference between intermolecular and intramolecular bonding is not clear to students. Where students *do* acknowledge intermolecular bonding, they may not realise these interactions are distinct from more familiar covalent bonds (Pereira & Pestana, 1991).

The status of hydrogen bonding is sometimes found to follow a trajectory from chemical bond, to 'just a force' and back. Taber reported that UK students first heard about hydrogen bonding in biology classes (at age 16-17), where they took the meaning to be 'a [covalent] bond to hydrogen' (such as that in methane). When these students came to understand that this was *not* what a hydrogen bond was, it could then lose its status as a 'proper' chemical bond (Taber, 2002, see also Barker & Millar, 2000).

While there is evidence that some learners come to appreciate the relationship between intermolecular bonding and physical properties, such as boiling point (de Posada, 1997; Peterson & Treagust, 1989; Peterson, et al. 1989; Taylor & Lucas 1997), research also reveals that some learners believe intermolecular bonding is *stronger* than intramolecular bonding (Goh et al. 1993; Peterson et al. 1989); that they invoke intermolecular bonding in inappropriate circumstances such as in ionic compounds (Taber, 2001a); or believe it is absent in polar molecular substances such as water (Birk & Kurtz, 1999; Fleming, 1994; Griffiths & Preston, 1989).

The limited consideration of the forces involved in chemical bonding and the consequent adoption by students of the perceived *need* of atoms to obtain full shells as the driving force for forming bonds, appears to make it difficult for learners to later develop a wider understanding of bonding that encompasses cases where there is neither full sharing nor net transfer of electrons (e.g. hydrogen bonding, van der Waals' forces, where shifts in electron density are more limited), or to accept compounds where the bonding does not lead to octet-like structure (e.g. BeCl_2 , SF_6 etc.), or to understand why an 'atom with an octet' (e.g. chlorine in AlCl_3) forms a dative bond.

Bonding may be intermediate between covalent and ionic

From the scientific viewpoint, most materials have bonding that cannot be considered to be 'purely' covalent or ionic (or metallic). Rather, these types of bonding are seen as ideals or prototypes, which provide useful categories for classifying many bonds to a 'first approximation'. In most materials, the bonding may be more precisely described as intermediate, with varying degrees of covalent and ionic (and metallic) character. The notion of bond polarity indicates that the covalent-ionic dimension should be seen as a continuum, and not a dichotomy.

However, research suggests that students tend to see bond polarity as an additional secondary characteristic of *covalent bonds*, rather than as something between covalent and ionic (Taber, 2001a). This is more than a pedantic quibble when senior high school level students believe that polar substances such as hydrogen fluoride exists as *molecules* in aqueous solution (Barker & Millar, 2000).

Bond polarity is explained using the idea of electronegativity, which can be understood in terms of a simple Coulombic model (i.e. a small atom with a large core charge will be more electronegative). Yet confusion over the understanding of electronegativity is widespread, resulting in a number of alternative conceptions: inability to establish the correct polarity of polar-covalent bonds, the view that non-polar molecules are formed between atoms of similar electronegativity, and that the number of valence electrons, the presence of lone pairs of electrons, or ionic charge determine molecular polarity (Birk & Kurtz, 1999; Boo, 1998; Harrison & Treagust, 1996; Peterson et al., 1989).

The apparent presentation of covalent and ionic bonds as a dichotomy (electron sharing vs. electron transfer; bonding between two non-metals, or between a metal and a non-metal), and the absence of discussion of the electrical basis of bonding (and the consequent adoption of the full shell explanatory principle), makes it difficult for learners to conceptualise bonding 'in between' ionic and covalent. The tendency to focus teaching about bonding on the ionic and covalent cases at introductory levels leads to these two archetypal models of bonding in compounds being seen by students as a dichotomy.

These introductory conceptualisations may continue to act as impediments to learning more sophisticated scientific models at university level (although less detailed evidence is currently available than for school and college levels). Even advanced graduate level (i.e. MSc and PhD) learners are unclear about aspects of polar bonding and the concept of the ionic-covalent continuum (Coll, 1999), and confuse intermolecular and intramolecular bonding (Coll & Taylor, 2001b). Despite demonstrated competence in the description and use of sophisticated mental models such as molecular orbital theory (Coll, 1999; Coll & Treagust, 2001), graduates retain naïve models such as the octet rule, only drawing on more sophisticated concepts (such as electron clouds and orbital overlap) when their simple models fail to explain the physical properties of

substances. Recent graduates, unlike more expert chemists, lack the skills to use and manipulate multiple mental models (see for example, Coll & Treagust, 2001;).

Recommendations for teachers

The identification of these problems, and the conjectured reasons for them, has led to suggestions about how to change the teaching of this topic (Taber, 2001a, 2002). At the present time these recommendations have not been adequately tested, and more research is clearly needed, but they represent principles for research-based practice derived from the current state of knowledge.

Build on physical principles

Although bonding concepts can be explained in terms of electrical forces, and the principles of quantum mechanics, we find that many students either do not share this underlying physical framework, or at least do not apply it in chemical concepts. Many seem to explain chemistry in terms of 'magical' concepts such as 'octets', and almost vitalist concepts such as the 'needs' and 'desires' of the individual atoms (c.f. Benfey, 1982).

It is therefore sensible, when teaching the subject, to remember to *explicitly refer* to the underlying principles, and *not to assume* that learners are recognising the physical forces involved. In particular, at an introductory level, the electrical interactions within and between molecules and ions need to be emphasised at all times.

Focus on molecules and ions rather than atoms

Discrete atoms are seldom featured in significant chemical processes, yet the atom has had a privileged place in the teaching and learning of chemistry. Most real chemistry involves molecules, or ions, or more extensive systems. Even when chemistry does involve atomic radicals (such as $\bullet\text{Cl}$), the atoms do not feature in the initiation step. Yet, students often assume, and books and teachers often imply, that chemical reactions occur between atoms (Taber, 2002).

If learners' atomic ontologies act as serious learning impediments, then it may be pedagogically more sound for teachers to conceptualise chemical structures from an alternative perspective better in tune with modern science, and which does not give atoms an undue emphasis in the molecular menagerie. This approach would consider chemical structures (molecules, ions, lattices etc.) to be various arrangements of atomic cores and clouds of valence electrons, rather than arrangements of atoms (Taber, 2001a). Such an approach would be a more authentic model, and a better basis for progression in learning about chemistry, but will require chemistry teachers to overcome fixed habits of mind. The research suggests that the 'everything is made of atoms' approach is a simplification that is readily learned, but not readily developed into more sophisticated understanding.

Teach bonds as electrical concepts (not magical or social concepts)

Chemical bonds may be understood as the physical forces which hold together chemical systems such as molecules and lattices. Bonds hold molecules together (both internally, and to one another), because work must be done against electrical forces to separate the parts of the system. Bonds form because of those electrical forces.

Although these processes can later be understood in more sophisticated ways (in terms of electrical potential wells or surfaces, and allowing for the effect of quantum restrictions), it is appropriate to introduce bonding as an electrical effect, at some 'optimum level of simplification' (Taber, 2000, 2002) that balances what the learner is ready to understand with what provides a valid basis for further learning.

The ionic bond is the force of attraction between an ion and those counter ions around it, yet many learners have been found not only to associate, but to *identify*, the ionic bond with an electron transfer between discrete atoms. Just as it would seem perverse to speculate about the genesis of the methane and oxygen molecules when considering the reaction between methane and oxygen, it is not appropriate to focus on *ion formation* (in terms of electron transfer) to *explain* ionic bonding. If a sodium chloride crystal forms by evaporation of a salt solution, and we are interested in the bonding in the crystal, it is of little relevance how the ions (already present in the

solution) came into being. When bonding is seen as a primarily electrical phenomena there is no need to explain it in terms of the *needs* of atoms.

If covalent, ionic and metallic bonds are explained in electrical terms, students are better prepared to accept that hydrogen bonds, van der Waals forces, solvent-solute interactions etc. are also types of chemical bonding. Where learners see covalent bonds as electron pairs attracted to two different positive cores, they have a good basis for subsequently learning about electronegativity and bond polarity.

If bonding is seen to be due to the same electrical forces that hold the individual atom together, then there is no reason for them to consider that the interaction between an atomic core and an electron in a bond is different depending upon whether or not the electron originated in that particular atom, and the student can have an open mind about where an electron will go on bond fission (so that heterolytic fission is considered possible).

The teaching of chemical bonding is dominated by processes (e.g., hypothetical schemes for bond formation), not outcomes (i.e. the nature of the bonding, and its consequences: solubility, melting temperature, etc.). This is inappropriate since scientists are interested in using chemical bonding to explain properties and predict reactivity.

Emphasise the non-molecular nature of non-molecular lattices

Just as the atom takes on a significance in learning about chemistry which is not justified in terms of the (lack of) role that discrete atoms play in chemical processes, once learners have been taught about molecules there is a tendency to apply the 'molecule schema' to all structures. The role of valency in limiting, if not exactly determining, molecular formulae, may be extended to metals and to ionic materials. Metals may be seen by students to consist of discrete molecules of similar atoms, in a similar way to iodine or phosphorus. In the ionic case, valency is seen by many students to indicate the number of ionic bonds that can be formed, and not just the charge on the ions.

One possible teaching tactic that may be employed to avoid these problems is to teach about metallic bonding first (ions, delocalised electrons - but no molecules), then ionic bonding (with the

significance of co-ordination number being emphasised as well as the need to balance charge overall), and covalent bonding last (see table 1). Even then, it may even make sense to first teach about covalently bound crystals such as diamond and silica which are *similar* to metallic and ionic crystals in having a single form of bonding holding the structure together. The case of discrete covalent molecules, which may then be bound together by another form of bond (still electrically based, but weaker) is the most complex.

Table 1: a teaching order for solid structures (from Taber 2001a)

type of structure	bonding	comments:
1. metallic crystal	metallic: cations (atomic cores) + delocalised electrons	one element present; charge on cation related to valency
2. ionic crystal	ionic: cations and anions	added complication: two (or more) elements; stoichiometry determined by charge ratios
3. giant covalent	covalent	added complications: number of bonds (and stoichiometry, if a compound) determined by valency; bonds have specific directions
4. simple covalent	covalent intramolecular, plus intermolecular (van der Waals, H-bond)	added complication: additional level of structure - need to consider discrete molecules, and arrangement of molecules in crystal.

Take care with language

A final recommendation is to take particular care in the use of language. The need, when explaining chemistry, to shift between the everyday and the theoretical molecular levels of description will always be with us, and will make demands on learners; so it is important to clearly signpost the transitions, and make sure that we *model* the way chemists use molecular notions to explain macroscopic phenomena which have distinctly different features (for example, see Taber, 2002).

We must be careful to ensure that substances, and not molecules, are said to evaporate, melt, expand, reflect, conduct and so forth. Where we do not have distinct vocabulary to distinguish molecular from molar phenomena it may be necessary to coin suitable terms: for example, when

samples of chemical substances *react* together: their molecules should be said to do something else. (The term 'quantact' has been mooted, but has not yet been widely adopted.)

Bonding has a peculiar position in the molar-molecular scheme: for we often talk of bonds *as if* they are macroscopic properties of materials, so here we need to be particularly careful to ensure we are clear about the level of description we are employing.

The evidence about using anthropomorphic and other metaphorical language to teach chemistry is more equivocal. In the long term, we do not want our students to think that molecular interactions are like social interactions, and can be explained in terms of psychological drives. Yet these everyday scripts can provide familiar images that give students get a way into imagining and describing atomic and molecular level systems.

The professional pedagogic skill of the teacher is needed to make sure that such comparisons become the *anchors* for more scientific descriptions and explanations (Taber, 2002). Whenever anthropomorphic language is used it should be followed by subsequent reiterations in more formal scientific terms. If learners' anthropomorphic comments are indulged in class, they could still be reflected back in more technical terms. Learners can be challenged to translate their explanations into a form that does not call upon needs, wants, desires, preferences and what the atom might be thinking. In this way their 'conceptual ecologies' can be shifted (Taber, 2001c) to provide the fertile ground needed for the development of more scientific models of the chemical bond.

Conclusion

Research has revealed that students may have a wide range of alternative conceptions relating to aspects of chemical bonding - although many of the common ideas fit into a clear pattern: a coherent alternative framework. Early in secondary school, students may see bonds as substantive physical links, but as they develop a greater appreciation of the molecular realm they start to conceptualise bonding differently. However, rather than understanding bonds as the electrical forces due to the charges on electrons and nuclei, they are most likely to consider bonds as akin to social arrangements made between small groups (molecules) of atoms in order to obtain full outer

shells of electrons. By the end of compulsory secondary education (around 16 years), such a model seems to meet students' needs in terms of the limited explanations of covalent and ionic bonding that they are expected to reproduce in tests, in order to achieve scholastic success.

Yet such explanations are not consistent with the range of bonding phenomena which the student will meet if they elect to continue their study of chemistry. The octet framework leads to assumptions about why bonds form, what should count as a bond, and how bonds should be classified, which are inconsistent with the requirements of college and university level chemistry. Even when students are able to move beyond the limitations of 'octet thinking', they find it hard to fully transcend these habits of mind, which may resurface in the thinking of university students and graduates.

In this chapter, we have not only reported the findings from research into learners' ideas about chemical bonding, but we have also considered how and why such ideas develop. This has allowed us to make informed recommendations on what research based practice might look like when teaching about chemical bonding.

References

- Barker, V. and Millar, R. (2000) Students' reasoning about basic chemical thermodynamics and chemical bonding: what changes occur during a context-based post-16 chemistry course?, *International Journal of Science Education*, 22 (11), pp.1171-1200.
- Benfey, O. T. (1982) The concepts of chemistry - mechanical, organicist, magical or what?, *Journal of Chemical Education*, 59 (5), pp.395-398.
- Ben-Zvi, R., Bat-Sheva, E. and Silberstein, J. (1986) Is an Atom of Copper Malleable?, *Journal of Chemical Education*, 63 (1), pp.64-66.
- Birk, J. P. and Kurtz, M. J. (1999) Effect of experience on retention and elimination of misconceptions about molecular structure and bonding. *Journal of Chemical Education*, 76(1), 124-128.
- Boo, H. K. (1998) Students' understandings of chemical bonds and the energetics of chemical reactions. *Journal of Research in Science Teaching*, 35(5), 569-581.
- Butts, B. and Smith, R. (1987) HSC Chemistry Students' Understanding of the Structure and Properties of Molecular and Ionic Compounds, *Research in Science Education*, 17, 192-201.

- Coll, R. K. (1999, July) *Learners' mental models of chemical bonding: A cross age study*. Paper presented at the 30th Annual Conference of the Australasian Science Education Research Association. Rotorua, New Zealand.
- Coll, R. K. and Taylor, N. (2001a) Alternative conceptions of chemical bonding amongst senior secondary and tertiary students: Nature and origins. *Teaching and Learning*, 22(1).
- Coll, R. K. and Taylor, N. (2001b) Alternative conceptions of chemical bonding for upper secondary and tertiary students. *Research in Science and Technological Education*, 19 (2), 171-191.
- Coll, R. K. and Treagust, D. F. (2001) Learners' mental models of chemical bonding, *Research in Science Education*, 31(3), 357-382.
- Cros, D., Amouroux, R., Chastrette, M., Fayol, M., Leber, J. and Maurin, M. (1986) Conceptions of first year university students of the constitution of matter and the notions of acids and bases, *European Journal of Science Education*, 8 (3), pp.305-313.
- Cros, D., Chastrette, M. and Fayol, M. (1988) Conceptions of second year university students of some fundamental notions in chemistry, *International Journal of Science Education*, 10 (3), pp. 331-336.
- de Posada, J. M. (1997) Conceptions of High School Students concerning the internal structure of metals and their electronic conduction: structure and evolution, *Science Education*, 81 (4), pp. 445-467.
- Driver, R., Squires, A, Rushworth, P. and Wood-Robinson, V. (1994) *Making Sense of Secondary Science: research into children's ideas*, London: Routledge.
- Fensham, P. (1975) Concept formation, in Daniels, D. J. (Ed.), *New Movements in the Study and Teaching of Chemistry*, London: Temple Smith, pp.199-217.
- Fleming, K. (1994) Chemical bonding - what conceptual ideas do students really hold? What implications are there for teachers? Paper presented at CONASTA 43, Launceston, Tasmania, July 1994.
- Gilbert, John K. (1998) Explaining with models, in Ratcliffe M. (ed.) *ASE Guide to secondary science education* (London: Stanley Thornes), pp.159-166.
- Gilbert, J. K. and Zylbersztajn, A. (1985) A conceptual framework for science education: The case study of force and movement, *European Journal of Science Education*, 7 (2), 107-120.
- Goh, N. K., Chia, L. S. and Tan, D. (1994) Some analogies for teaching atomic structure at the high school level. *Journal of Chemical Education*, 71 (9), 733-788.
- Griffiths, A. K. and Preston, K. R. (1989). *An investigation of grade 12 students' misconceptions relating to fundamental characteristics of molecules and atoms*. Paper presented at the sixty-second conference of the National Association for Research in Science Teaching, San Francisco.
- Griffiths, A. K. and Preston, K. R. (1992) Grade-12 students' misconceptions relating to fundamental characteristics of atoms and molecules, *Journal of Research in Science Teaching*, 29 (6), pp. 611-628.
- Hapkiewicz, A. (1991) Clarifying chemical bonding, *The Science Teacher*, 58 (3), pp.24-27
- Harrison, A. G. and Treagust, D. F. (1996) Secondary students' mental models of atoms and molecules: implications for teaching chemistry, *Science Education*, 80 (5), pp.509-534.

- Harrison, A. G. and Treagust, D. F. (2000) Learning about atoms, molecules, and chemical bonds: a case study of multiple-model use in grade 11 chemistry, *Science Education*, 84, pp.352-381.
- Jensen, W. B. (1995) *Logic, History and the Teaching of Chemistry*, text of the Keynote Lectures, given at the 57th Annual Summer Conference of the New England Association of Chemistry Teachers, Sacred Heart University, Fairfield, Connecticut.
- Johnstone, A. H. (1991) Why is science difficult to learn? Things are seldom what they seem, *Journal of Computer Assisted Learning*, 7, pp.75-83.
- Lijnse, P. L., Licht, P., de Vos, W. and Waarlo, A. J. (editors) (1990) *Relating Macroscopic Phenomena to Microscopic Particles: a central problem in secondary science education*, Utrecht: Centre for Science and Mathematics Education, University of Utrecht: CD-fl Press.
- Nelson, P. G. (1996) To be a molecule, or not to be?, *Education in Chemistry*, 33 (5), pp.129-130.
- Ogilvie, J. F. (1990). The nature of the chemical bond, *Journal of Chemical Education*, 67(4), 280-289.
- Oversby, J. (1996) The ionic bond, *Education in Chemistry*, 33 (2), pp.37-38.
- Pereira, M. P. and Pestana, M. E. M. (1991) Pupils' representations of models of water, *International Journal of Science Education*, 13 (3), pp.313-319.
- Peterson, R. F. and Treagust, D. F. (1989) Grade-12 students' misconceptions of covalent bonding and structure. *Journal of Chemical Education*, 66(6), 459-460.
- Peterson, R. F., Treagust, D. F. and Garnett, P. (1989) Development and application of a diagnostic instrument to evaluate grade-11 and grade-12 students' concepts of covalent bonding and structure following a course of instruction. *Journal of Research in Science Teaching*, 26(4), 301-314.
- Pfundt, H. and Duit, R. (1991) *Bibliography: Students' alternative frameworks and science education*, IPN Reports in Brief, Germany: University of Kiel.
- Schmidt, H.-J. (1997) Students' misconceptions—looking for a pattern. *Science Education*, 81 (2), 123-135.
- Selley, N. J. (1978) The confusion of molecular particles with substances, *Education in Chemistry*, 15, pp.144-145.
- Solomon, J. (1993) The social construction of children's scientific knowledge, Chapter 5, in Black, P. J. and Lucas, A. M. (eds.), *Children's Informal Ideas in Science*, London: Routledge, 85-101.
- Taber, K. S. (1994) Misunderstanding the ionic bond, *Education in Chemistry*, 31 (4), pp.100-103.
- Taber, K. S. (1995) An analogy for discussing progression in learning chemistry, *School Science Review*, 76 (276), pp.91-95.
- Taber, K. S. (1997) Student understanding of ionic bonding: molecular versus electrostatic thinking?, *School Science Review*, 78 (285), pp.85-95.
- Taber, K. S. (1998) An alternative conceptual framework from chemistry education,.
- Taber, K. S. (2000) Finding the optimum level of simplification: the case of teaching about heat and temperature, *Physics Education*, 35 (5), pp.320-325.
- Taber, K. S. (2001a) Building the structural concepts of chemistry: some considerations from educational research, *Chemical Education: Research and Practice in Europe*, 2 (2), pp.123-158.

- Taber, K. S. (2001b) The mismatch between assumed prior knowledge and the learner's conceptions: a typology of learning impediments, accepted for publication in *Educational Studies*, 27 (2), pp. 159 - 171.
- Taber, K. S. (2001c) Shifting sands: a case study of conceptual development as competition between alternative conceptions, *International Journal of Science Education*.
- Taber, K. S. (2002), *Misconceptions in chemistry - prevention, diagnosis and cure*, London: Royal Society of Chemistry.
- Taber, K. S. and Watts, M. (1996) The secret life of the chemical bond: students' anthropomorphic and animistic references to bonding, *The International Journal of Science Education*, 18 (5), pp. 557-568.
- Taber, K. S. and Watts, M. (2000) Learners' explanations for chemical phenomena, *Chemical Education: Research and Practice in Europe*, 1 (3), pp.329-353, available at http://www.uoi.gr/conf_sem/cerapie/
- Tan, D. and Treagust D. F. (1999) Evaluating students' understanding of chemical bonding, *School Science Review*, 81 (294), pp.75-83.
- Taylor, N. and Lucas, K. (1997) The trial of an innovative science programme for preservice primary teachers in Fiji. *Asia-Pacific Journal of Teacher Education*, 25(3), 325-343.
- Tsaparlis, G. R. (1984) The chemical bond as an atomic tug-of-war, *Journal of Chemical Education*, 61 (8), pp.677.
- Tsaparlis, G. (1997) Atomic and molecular structure in chemical education, *Journal of Chemical Education*, 74 (8), August 1997, pp.922-925.
- Watts, D. M. & Gilbert, J. (1983) Enigmas in school science: students' conceptions for scientifically associated words, *Research in Science and Technological Education*, 1 (2), 161-171.
- Wightman, T., in collaboration with Green P. and Scott P. (1986) *The Construction of Meaning and Conceptual Change in Classroom Settings: Case Studies on the Particulate Nature of Matter*, Leeds: Centre for Studies in Science and Mathematics Education - Children's learning in science project, February 1986.