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BUILDING THE STRUCTURAL CONCEPTS OF CHEMISTRY: SOME CONSIDERATIONS FROM EDUCATIONAL RESEARCH

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ABSTRACT: This paper sets out to consider how educational research into the learning of structural aspects of chemistry might inform teaching practice. The paper is based around a review of research findings into learners' difficulties in developing the scientific models of chemical structures (atoms, molecules, lattices etc.) This forms the second of the four sections into which the paper is organised. The paper begins by considering how ideas about the learning process can inform our understanding of alternative conceptions and frameworks in chemistry, and - therefore - how we should view the research reviewed in the second section. This is a consideration of the findings of studies into difficulties learning about the molecular model; atomic structure; molecular structure; and lattices. This review is followed by a section identifying some key 'pedagogic impediments' - alternative aspects of learners' thinking that seem to *derive from the way the subject is taught*. In the final section some practical suggestions are made regarding how the teaching of chemistry may be revised to help learners construct the scientific models rather than develop the alternative conceptions. [*Chem. Educ. Res. Pract. Eur.*: 2001, 2, 123-158]

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INTRODUCTION: CONSTRUCTING CHEMICAL KNOWLEDGE

Given that we do not enter this world with an instinctive knowledge of chemistry, such a body of knowledge must be actively acquired by learners. Popular images of the learning process being the *transfer* of information from teacher (or book) to student are now widely discredited. Such an image suggested that learning should not be problematic, providing the information source is accurate and clear, the recipient is attentive, and the 'transmission line' has good signal-to-noise characteristics. In practice, learning of a complex subject such as chemistry certainly *is* problematic. Indeed, there is considerable evidence of both learners *failing to make sense* of teaching, and of students *developing alternative conceptions* of the material presented to them.

A constructivist perspective on learning science

The constructivist perspective on learning (Taber, 2000a) is now widely accepted - at least in general outline. Learning is an active process that takes place in the mind of the learner, and during which information from sources in the environment (including - but not

limited to - teachers, textbooks and peers) is re-interpreted in terms of existing knowledge and understanding. Whilst there are certainly differences in emphasis - for example, in the extent to which knowledge construction should be seen as an *inter*-personal rather than just an *intra*-personal activity - it is generally accepted that meaningful learning requires the student to make sense of new information in terms of existing 'cognitive structure'.

The overwhelming evidence for this type of approach is the vast literature on students' alternative conceptions in science (e.g., Driver et al., 1994), which demonstrates that '*understanding differently*' is as likely an outcome of teaching as *understanding-as-intended* or indeed *not understanding* (i.e., not *making any sense* of the presentation). This literature shows that the way that our brains work provides a strong imperative to *make some sense* of what we are told, and to use whatever resources are available (in terms of existing conceptions) in order to interpret information into something meaningful.

However, we can only interpret such information in terms of our existing conceptual schemes, which provide the contextual frameworks for converting the utterances of teachers and text from books into meaningful mental images. Evidence (Miller, 1968) suggests that the average person has a very limited mental working space for processing new information: a restricted span of immediate memory, or what Calvin (1997) refers to as a limited scratch pad. Therefore teachers cannot rely on the coherence and logical development of their arguments as being sufficient to convey new ideas of any complexity. No matter how skilled the exposition, the mental 'register' of the student is soon saturated. In effect *our cognitive apparatus biases human understanding towards interpretation of information in terms of previous learning, rather than in terms of its own semantic context.*

Perhaps something about the nature of our ancestors' information-environment gave those with such a bias a survival advantage over those with more labile interpretations of what was going on around them? Conversely, perhaps by the time complex spoken language developed the human brain had become channelled along certain evolutionary paths from which any increase in mental processing powers required too great an 'activation energy'? Whatever the reasons, and no matter how inconvenient for students and teachers alike, most of us can only process complex *novel* information slowly. We take in a great deal of information when it can be interpreted in terms of existing conceptual frameworks (thus the power of the narrative form), but the unfamiliar has to be made familiar (Taber, 2001c) in manageable chunks if we are to learn it effectively. New concepts have to be appreciated and integrated into our knowledge systems before they can be used as a secure foundation for developing super-ordinate concepts - no matter how logical and clear the teacher's exposition.

Teachers must therefore help learners gradually *re-construct*, for themselves, the conceptual structure of a subject like chemistry. Deep foundations must be firmly in place before any attempt is made to put together the higher concepts. Often the metaphor of 'scaffolding' is used to describe the teachers' role in helping the learner build the edifice of a subject (Scott, 1998). As we shall see in the next section, this is probably an especially apt metaphor when the subject to be learnt is chemistry!

What's special about learning chemical concepts?

This constructivist perspective on learning applies to all science disciplines (and beyond), and is certainly not limited to chemistry. However, as a subject to be studied, chemistry does have some specific complexities!

For one thing chemistry may be traditionally seen as a single discipline, but it is one of diverse approaches. It is hard to believe that most physical chemists are working within the

same Kuhnian ‘disciplinary matrix’ (or ‘paradigm’) as most organic chemists. Perhaps the same point could be made of other subjects: although physicists seem to share commonalities of approach, biology is certainly a ‘broad church’. (Indeed, perhaps there is no biology now, only biological or life sciences?)

This, however, brings me to a second point about chemistry: its relationship to physics. Although there is some truth in the simplistic notion of a hierarchy of sciences building upon each other (physics on maths., chemistry on physics, biology on chemistry, psychology on biology etc.), these relationships are not equivalent. Maths is clearly a tool used heavily in physics (and in other sciences), but there are large parts of biology that make little or no *direct* application of chemistry, and those parts of psychology that build heavily on biology are probably best conceptualised as parts of life science (whereas other fields within psychology are perhaps better classed as social science).

To my mind, the relationship between chemistry and physics is different in nature, in that *no matter which aspects of chemistry are considered, a fundamental reliance on aspects of physics is implicitly present*. A modern chemistry that was not heavily dependent on principles of electrostatics and quantum theory is difficult to envisage. The chemistry student is expected to appreciate the physical basis for the ideas used in chemistry - although this may often be a tacit assumption. However, in practice the student may well compartmentalise their chemistry learning separately from the physics, and this can be an impediment to effective learning (Taber, 1998a).

This is not to suggest that chemistry is ‘just’ an application of physics. The argument about the extent to which chemistry may be reduced to physics is an active one (e.g. Scerri, 1993, Van Hoeve-Brouwer 1996). A simplistic holistic/synergistic perspective would suggest that the higher level concepts of chemistry contain something more than the sum of their physical parts - that is *emergent* properties, not anything supernatural! Yet, even if it *were* possible to reduce chemistry to physics *in principle*, this would not be appropriate. When new, higher level, concepts are developed from more fundamental principles, they are then able to provide a ‘short-hand’ that stands in place of the basic concepts - akin to the mathematician who does not need to keep re-proving the established and accepted theorems used in new proofs.

Novel concepts introduced into the subject are then new mental tools that may be explored in their own right (Taber, 1995a). The limitations in human processing power (alluded to above) may well be part of the reason why we find it convenient to use the higher level concepts: perhaps the advantage that accrues from having a new mental entity to manipulate, formed part of the selection pressure against evolving a larger mental working capacity?

The ‘physical’ basis of many chemistry concepts is not the only potential difficulty for learners. Many ideas that students meet involve a form of ‘bootstrapping’ (a term borrowed from the paradoxical image of having to ‘pull oneself up by one’s own boot-laces’: intended to imply a ridiculous plan or fantastic achievement). In other words, although we may think of chemistry as being a *logical* subject, many chemical concepts can *not* be learnt in an entirely logical manner, at least not in terms of clearly following deductively from previously accepted ideas and/or interpretation of empirical evidence.

Indeed, to a great extent, the theoretical content of chemistry is *best seen* as a set of models. Models play a major role in all science disciplines (Gilbert, 1998), but again they seem to present a particularly problematic nature to the learner of chemistry.

In physics education there is a debate about the relationship between Newton’s and Einstein’s theories regarding motion. Newtonian physics is taught in school, although it has been superseded by Einstein’s ideas. Some views suggest that the equations deriving from

Newton's ideas are generally a good approximation to Einstein's theory: after all - it is argued - they give results that are accurate enough to put people on the moon! Conversely, because Einstein's world-view starts from a very different set of assumptions (i.e. the speed of light is the invariant, rather than distance and time being absolutes) some argue that Newton's ideas are strictly speaking *wrong*, and can not be seen as an approximation to the best model of reality available.

I do not wish to undermine the importance of such a debate, but in some ways this situation is very simple. Einstein's theories provide the best current model for understanding this aspect of reality. Newton's equations are based on false premises, but provide an important historical advance on earlier conceptualisations, and give accurate predictions in most practical applications. The debate is only about the status that should be accorded to the Newtonian model *when it is taught*.

How does this compare with the situation in chemistry? In my view chemistry teachers work within a much more fuzzy context. There are clearly many historical ideas in chemistry that are no longer that useful (phlogiston for one), but in many situations we have multiple models that each have some value (Taber, 1995a).

Consider the case of acids. The Lewis and Bronsted-Lowry definitions are just that - definitions - and the use of one does not automatically negate the other (although it may expand the set of substances which are *classed* as acids). Definitions of redox provide a similar case. Moving from *definitions* in terms of loss/gain of oxygen/hydrogen to transfer of electrons does not mean moving to a better model of what oxidation is *in nature*. *Oxidation is whatever we define it to be*: if chemists find it more useful to classify reactions they perceive as similar in terms of a definition of oxidation that no longer relates to oxygen, so be it. However, as Schmidt (1991) has highlighted (with relation to alternative conceptions of 'neutralisation', see below) labels can act as 'hidden persuaders', and to the learner the term *oxidation* may be suggestive.

Of course, we do not stop at the transfer of whole electrons, because we know that elements differ in electronegativity, so many reactions give *shifts in electron density* that may be *conceptualised as partial electron transfers*. As chemists *we decide* that it is useful to think about these shifts in this way, and to extend our definition of oxidation to include such partial shifts. We develop a formalism for this: oxidation states.

Now I am not suggesting that any of this is completely arbitrary, but it is not an *uncovering* of nature (something the naive chemistry student might expect us to be involved in), so much as *dress[ing] nature up* in a convenient way (c.f. Zavaleta, 1988). We can certainly *justify* this: we are developing a theoretical framework that helps us make sense of nature. But oxidation numbers *do not exist in nature* except in human minds. (I suggest Popper (1979) would class them as 'world 3' objects.)

When we come to teach oxidation numbers we may well try to contextualise the rationale (and logically relate the scheme to ideas about electronegativity, which in turn are based upon those physical principles some of our chemistry students find so difficult), but what we are actually teaching may certainly *seem* to learners about as arbitrary as the rules of chess. So oxygen is given the oxidation number of -2 in its compounds, unless the compound is classed as a peroxide or superoxide when... Given that limited processing capacity noted above (which will always provide a rate-limiting-step for learning), it is hardly surprising that students find the subject difficult.

The bootstrapping aspect may be seen in the previous example. To apply oxidation numbers to oxygen one not only has to distinguish between oxygen in the element (by definition: o.n. = 0), and in the ionic state (o.n. determined by the ionic charge), *and* in its compounds; but one further has to be able to distinguish between oxygen in compounds

where it is in oxidation state -2, -1 or -0.5. This is not difficult once one is an expert: but this presents a circular problem to the novice!

It was Thomas Kuhn (1977) who pointed out that definitions in science are not always as straightforward as we might be led to expect. This seems particularly the case in chemistry (Taber, 1995b). Definitions of the most fundamental concepts (element, compounds, molecule, ion etc.) are problematic. They are either bland and uninformative, technically dubious, or need to be long and detailed enough to fall outside our usual expectations for a 'definition'. (It is a useful exercise to look up the basic terms suggested above, either in general or technical dictionaries, and see how many definitions can be found which are both *accurate* - either in not being plainly incorrect, or that you cannot immediately think of exceptions - and also *brief* enough to be useful; and yet which also convey *useful* meaning to the novice.)

Consider what is meant by isomerism. Schmidt (1992) reports that learners often limit their notion of isomer to members of the same class of compounds (so for example $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$ and $\text{CH}_3\text{CH}_2\text{CH}(\text{OH})\text{CH}_3$ would be seen as isomers, but not $\text{CH}_3\text{CH}_2\text{O}\cdot\text{CH}_2\text{CH}_3$). This type of alternative conception seems to be best understood as a kind of categorisation error: the learner grasped some essence of what is meant by an isomer, but limits the definition to a narrower range of application than the chemist.

It is hardly surprising that as well as those chemical concepts considered 'mechanical' (based clearly on physical principle), there are those which are judged 'organicist' (where the holistic principle comes into play, and the notion 'takes on a life of its own'), and even those labelled as 'magical' (Benfey, 1982). Sadly, to many of our students, most of chemistry probably seems to fit into the last category!

In practice, although definitions may often be readily learned by a rote ('an element is a substance [sic] which cannot be broken down into something simpler [sic] by chemical [sic] means'), they only offer useful meaning *after* the concept itself is understood!

This *need* not be a major problem as long as, as teachers, we bear this in mind and teach accordingly. Thagard (1992) has studied episodes from the history of science, and suggests that 'paradigm-shifts' in the thinking of scientists are often preceded by long periods when the 'revolutionary scientist' is gradually exploring the new ways of thinking about the material. Similarly, there are many reports from contemporary scientists that show how breakthroughs follow periods of deep immersion in the problem, often followed by an interval (e.g. of sleep or relaxation) when some sort of subconscious processing has taken place. Evidence from detailed case studies suggests that when students learn science they may undergo similar processes: gradually building up several alternative frameworks for thinking about a concept area (Taber, 1995c, 2000b), and then undergoing 'sudden' insights that may just be the obvious signs of the gradually shifting sands of the learner's conceptual ecology (Taber, in press).

What's special about alternative conceptions in chemistry?

The study of learners' alternative conceptions, and alternative conceptual frameworks has been a busy field within science education for two decades now. Bibliographies (Carmichael et al. 1990-2; Pfundt & Duit, 1991) and reviews (e.g. Taber, 2000a; Barker 2000) of the findings have been produced, scholarly books have debated the nature and status of learners' ideas (e.g. Black & Lucas, 1993), and more recently the field has now been characterised as a spent force (Solomon, 1994), a disproportionate focus of research attention (Johnstone, 2000) and as potentially still a progressive research programme (Taber, 2000c).

In view of the vast amount that has been written, I will not attempt to add much here.

However, I do believe that it is important to note that *most alternative conceptions in chemistry do not derive from the learner's unschooled experience of the world.*

Common misconceptions in the other two major sciences can often be seen as 'naive' notions, or 'intuitive theories' deriving from interpreting early life experiences. The 'impetus' notion can probably be considered as the archetype of all alternative conceptions. Something like 85% of school pupils are likely to hold conceptions of force and motion alternative to those taught in school (Watts & Zylbersztajn, 1981). Basically, real life experience teaches us that if you push an object, you provide it with a limited amount of something (i.e. 'impetus'), which is soon used up, and so the object stops. Physics teaches us that when a force is applied to an object it will be accelerated to a new velocity at which it will then travel for ever (in the absence of any other forces). The two schemes are not inconsistent, but the context that physics takes as the starting point for analysis (no gravity, no friction...) is hardly that which suggests itself in everyday life.

Similarly, common experience of the world can understandably lead to a belief that most of the 'stuff' in a tree came from the soil. After all, most plants need soil, they often put down deep roots, and if essential minerals are lacking in the soil the plant dies. Yet, of course, science teaches that most of the mass derives from the carbon dioxide in the air. As with the impetus theory, the derivation of the alternative conception is no great mystery - it is simply the child as scientist (Driver, 1983) making sense of the world in terms of the data readily available (as simply as possible, without making too many auxiliary assumptions).

So in both physics and biology there is little surprise that pupils come to class with alternative ideas that do not match orthodox science. Indeed there may even be something more literally *intuitive* about some of these conceptions. Whereas some animals have their behaviour determined largely by instinct we tend to think of humans as being more evolved, and using logic to rationally decide how to act and what to think. Yet there are clearly many ways in which we are biased in the way we perceive our environment. For a simple example, human eyes have a particular spectral response to incident radiation. We do not see into (what is for us) the ultra-violet like some animals; we have receptors with three different types of frequency-sensitivity response, unlike many animals. We undoubtedly see the world differently to most other species we share it with.

In a similar way, our brains are channelled by evolution to process information in certain ways. Arguments about the degree to which mind can be considered 'modular' (e.g. Fodor, 1985, Karmiloff-Smith 1994) are not settled, but there is one school of thought (Mithen, 1998) which argues that the human brain has evolved certain *intuitive* responses to the physical world (that helped us develop tool use) and to the biological world (that helped in eating and not being eaten). Some evidence is found in the way that people tend to have similar 'instinctive' ways of classifying living things (Gelman & Markman, 1986) - much to the irritation of biology teachers. In 'folk biology' people and insects clearly should *not* be classed as animals alongside *real* animals like dogs and horses; and trees are clearly *bigger than* plants!

As well as modules for mechanics and nature study, there are good arguments for a kind of social psychology module, and it is now widely accepted that we all have an in-built Chomskyan 'language acquisition device' (Mithen, 1998). *No such inherent mental module has been suggested for chemistry.*

All this suggests that chemistry teachers should *not* have to cope with many alternative conceptions among learners. If there are no instinctive notions about molecules or oxidising agents, and little opportunity to develop naive notions about atomic structure or proton-donors, then the chemistry teacher *should* have a real advantage over her peers in the prep. room. Yet, of course, this is far from the case. There are a great many areas where

alternative chemical conceptions have been uncovered. It is therefore appropriate to ask how these derive?

Carr (1984) has argued that many problems that learners have in chemistry maybe best characterised as ‘model confusion’, and in view of the points made above about the nature of the subject, this seems a likely candidate. Where there are several models of acids, or oxidation, or of atomic structure, being presented to students there is clearly the scope for much confusion! This is particularly so when most learners have a very limited notion of the role of models in science (see Grosslight et al., 1991; Driver et al., 1996) Zoller (1990: 1063) refers to the,

“many *abstract, non-intuitive concepts* which are *not* based on, and/or derived from, and/or interrelated logically with one another, at least not in a simple and straightforward sense”.

Another possible source of problems are linguistic cues. So, for example, the notion of electron *shell* has been suggested to imply a protective outer coat to pupils (Harrison & Treagust, 1996), and Taber (1997a) found that one of the students he interviewed consistently used the term ‘shield’ for ‘shell’: so, for example, he suggested that a carbon atom ‘needs four electrons to make up a full outer *shield*’.

Schmidt (1991) has discussed the example of a common misconception about neutralisation: that the *neutralisation* of acid and base always gives a *neutral* product. Schmidt refers to the ‘*neutralisation*’ label as ‘a hidden persuader’: after all pupils are usually introduced to neutralisation reactions through examples where strong acids react with strong bases to give a neutral solution. It should not be a great surprise that many pupils infer that part of the meaning of ‘neutralisation’ is to produce something neutral!

Yet *some* alternative conceptions seems to require something more in the way of explanation. For example, Taber has described a common alternative conceptual framework from chemistry which leads to a range of misconceptions about aspects of such topics as bonding, ionisation energy, and reactions (Taber, 1998b, 1999a). This ‘octet’ framework has been well characterised (see below), and is judged to have several ‘causes’, including the over-generalisation of the octet rule to become a general-purpose heuristic for rationalising chemistry; the lack of appreciation of the electrostatic nature of interactions within and between molecules; and something Taber refers to as the ‘atomic ontology’: the mind-set that assumes we start thinking about chemistry from atoms.

Like other alternative conceptions - such as those from physics and biology - the *immediate* cause may be considered to be the way the learner makes sense of new information provided by the teacher in terms of what they already know. But whereas that pre-existing knowledge in *those* subjects may well be based on early life experience of pushing objects or ‘feeding’ plants, in chemistry the frameworks available for making sense of such abstract notions as molecular geometry, or lattice structures, derive from *the learners’ understanding of prior science teaching* (Taber, 1995d, 1999a). Such a distinction is of course, just to a ‘first approximation’; but, nevertheless, it may be significant.

Fleming interviewed 15 year old secondary pupils (1994a) and University students (1994b) in New Zealand to explore their ideas about chemical bonding. He found his work supported Taber’s (1993) suggestion that alternative conceptions about bonding should be considered to derive from instruction. Fleming elicited many alternative conceptions from the 15 years olds, who had not *formally* studied the topic in school science, and from the university students who told him that *their* ideas about the subject were largely what they had been taught.

Seen from this perspective, many failures to learn derive from the *mismatch* between the ideas the teacher believes the student has available, and those ideas they actually bring to mind in the context of instruction (Taber, accepted for publication). Taber's typology for such mismatches has its main distinction between cases where the learner can not find any suitable ideas on which to anchor the new knowledge (a 'null learning impediment' either because the learner lacks the prior knowledge - a 'deficiency impediment' - or does not spot its relevance - a 'fragmentation impediment'), and those where the learner relates the new information to existing ideas that are inappropriate. This second type of 'substantive' learning block may be due to the student bringing in aspects of their 'life-world' experience which have not been understood in scientific terms (such as the 'impetus' conception); *or* it may derive from prior classroom learning which does not adequately reflect the pre-requisite learning needed to make sense of the new topic.

The significance of such a distinction is that whereas the physics teacher is *always* likely to meet pupils who bring impetus-type conceptions from their life experience ('ontological' blocks to learning), the chemistry teacher's problems largely derive from science teaching that has not been adequately planned ('pedagogic' or 'epistemological' impediments): where the sequence of ideas, the pace of presenting new models, the level of abstractions, or the size of 'knowledge chunks' presented do not match the needs of the learner.

This should not be taken to mean that 'we only have ourselves to blame' (after all: chemistry is complex, teaching is difficult, and pedagogy is poorly developed), but rather that *to some extent we have the power to make things better* for learners. If the octet rule *is* a significant impediment to learning chemistry (Taber, 1995e), then it is one put in place by chemistry teaching!

This article should be read in the light of these comments. The sections that follow catalogue *flaws* in chemistry learning: but such flaws should be viewed as part of the data needed to plan a better approach.

LEARNERS' DIFFICULTIES IN LEARNING ABOUT STRUCTURE IN CHEMISTRY

Transferring between the molecular and the molar

Any meaningful study of chemistry requires learners (1) to have grasped the notion of substance; (2) to appreciate that substances maintain their identity through a change of state; (3) to recognise that during chemical changes (a) the products are different substances to the reactants, yet (b) there is a conservation of matter at a more fundamental level. There is a considerable literature which considers how learners respond to these challenges. The present paper is concerned with the subsequent learning about chemical structure which can take place once this prerequisite understanding is in place, but readers interested in these areas are directed to the insightful work of Philip Johnson (e.g. 1996, 1998a, 1998b, 2000).

When this basic appreciation of the chemical 'substance' concept is in place learners are ready to consider the molecular level explanations that are so commonly used to explain the properties of substances. The work of Renström, Andersson and Marton (1980) demonstrates that the acquisition of the scientific notion of molecules can be a slow process: with molecules seen variously by learners at different stages in developing their understanding of the model as: only part of the substance; made up from the substance; and surrounded by the substance. So, as Ault and co-workers propose, "the basic proposition that 'everything is made of molecules' needs the added emphasis 'and *nothing* else' " (Ault et al., 1984, p.459).

One general problem that has widely recognised in the learning of chemistry is the difficulty learners have with the relationship between the molecular and the macroscopic (Jensen 1995). The learner experiences chemistry (substances and their interactions) at a molar level. Yet, a great deal of the theoretical structure of chemistry relies on entities that are on a molecular scale (ions, electrons, orbitals, ...). Chemists tend to switch effortlessly between considering these two levels when discussing their subject. Yet this habit presents several problems for the novice learner.

For one thing, as has already been pointed out, learners have limited ‘working space’ in which to process information, and explanations which include both molar and macroscopic features may well exceed this capacity. And as Johnstone (1991) points out, there is often an additional layer of symbolic representation in terms of chemical formulae *also* being presented concurrently. So Tsapalis (1997) recommends teaching introductory chemistry through three cycles at the macro, representational and sub-micro levels.

Some observers see an even more complex teaching context. Jensen (1995) considers the electrical level as distinct from the molecular, and Van Hoeve-Brouwer (1996), in discussing how Van der Vet considers that understanding chemical structure requires a theoretical level of description (i.e. distinguishing element, compound etc.) in addition to the macroscopic and corpuscular, suggests that a *separate* fourth distinct level of quantum chemistry may *also* be needed.

Another problem is the *abstract* nature of many of the characters in chemical dramas: atoms, molecules, ions, electrons, lattices, shells, etc. are all conjectured hypotheticals for which there can be no direct evidence the learner can experience. No matter how well our molecules and electrons reflect reality (if such a notion means anything) at the nanometre and picometre scale, this is not a reality that is directly accessible to the learner. The molecules and electrons we talk about in chemistry lessons are entities that have been *created* by scientists and teachers, as theoretical *tools* to think about, and talk about, our subject.

In other words, the molecules and electrons represented in text books and discussed in class are models, and pupils generally have a very simplistic notion of the role of models in science (as discussed above). One of the main ways in which these models are used in chemistry is to provide theoretical explanations *at the molecular level* which make sense of phenomena observed *at the macroscopic level* (see Figure 1).

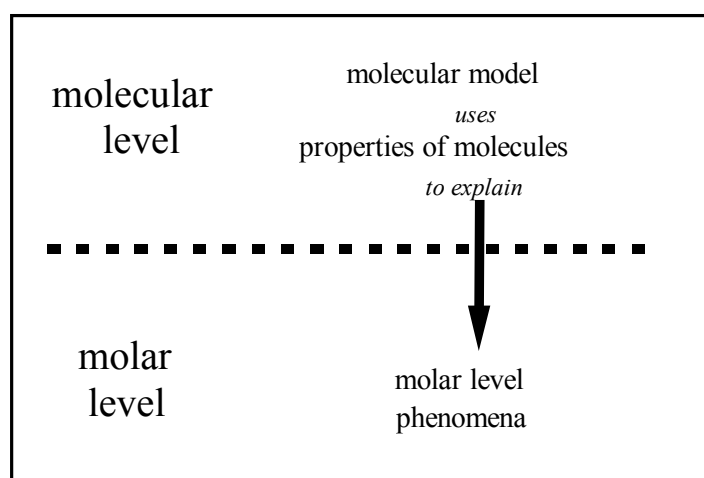


FIGURE 1: Two levels of analysis (from Taber 2000c).

Yet the research into learners' understanding of science suggests that learners commonly totally miss the key point that for these models to be useful the macroscopic behaviour of materials must make sense *in terms of the defined properties of the molecular level entities*. Often, when learners do not understand the role of the molecular level models, the properties to be explained at the molar scale are simply transferred to the molecular level (Figure 2). A simple example may make this clear. The melting of a solid (macroscopic phenomenon) may be explained in terms of the changes in the interactions between the particles (properties of the molecular level entities): but it is common for learners to make statements such as 'the substance melts because its molecules melt'.

There are many examples of this sort of thinking in the literature, such as 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 molecules (Ben-Zvi et al., 1986).

Even when molecular ideas are used to model explanations at the macroscopic level in a more sophisticated way, the learners may show naiveté in their conceptualisations - such as suggesting that light passes through glass but not metals because the molecules (sic) are closer together in metals than glass (de Posada, 1997).

As Gilbert and colleagues have discussed (Gilbert et al., 1998a, 1998b) using models to develop explanations is a key aspects of science. Yet this is not appreciated by learners, who generally have a very limited appreciation of the role of models in science (Grosslight et al, 1991). Learners also often fail to have a sophisticated appreciation of the nature of explanation (Watts & Taber, 1986; Taber & Watts, 2000), and so presumably view 'explanations' of the '*metals expand because the metal molecules expand*' variety as valid.

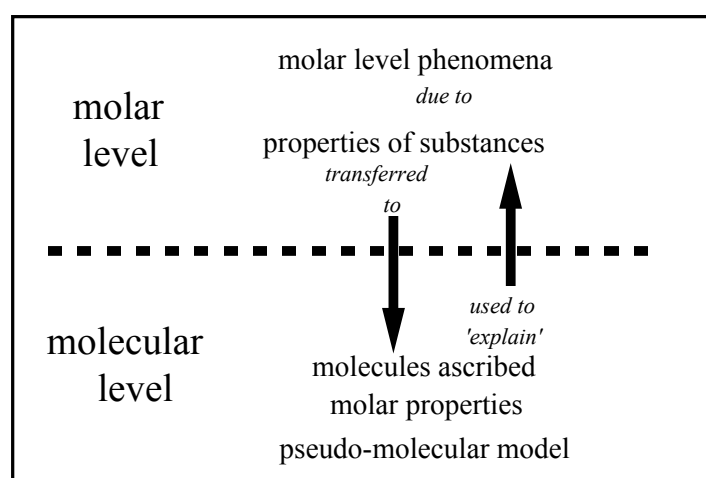


FIGURE 2: *Tautological molecular explanations (from Taber, 2000c).*

There is then a multiple barrier here: learners with limited mental working space (1), are asked to use abstract theoretical entities (2) at a level outside their direct experience (3), to explain apparently unrelated molar phenomena at another level; when they have limited appreciation of both the role of models (4) and the nature of explanation (5). Failures to learn chemistry should not surprise us.

It is quite difficult for teachers to help learners when we commonly have ingrained habits of using phrases such as 'the oxygen' to refer to the substance, its molecules (and sometimes its atoms) in the same explanation: after all, to the experienced chemist, the same phrase has different, but unambiguous, meanings *as the context shifts*. The teacher is blind to

the potential for confusion when she instinctively uses the same verbal label for quite distinct mental images.

In some cases there is not the vocabulary available to be as clear as we might like. For example, if we use the word 'react' to refer to what happens to substances at the molar level during chemical reactions, then we need a different word for what happens to the *molecular* level particles during this process. Interact is too general (as the molecules also 'interact' during the many elastic collisions that far outnumber the fruitful collisions where the [molecular equivalent of] reaction occurs! It has been suggested that a new term is needed: that the particles partact, or *quantact*, as the substances react (Taber, 2001a). In this way, we can reinforce the use of the molecular model as an explanatory scheme for molar events (Figure 3).

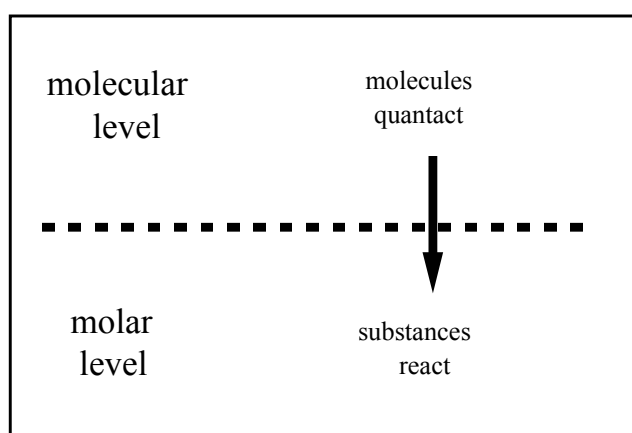


FIGURE 3: *Quantaction explaining reaction.*

Schmidt has posed the question of what is represented in the periodic table (Schmidt, 1998). Does the periodic table represent the elements (i.e. substances), or their atoms? Many periodic tables that pupils may see will include data that relates to both levels (e.g. melting temperature and electronic configuration). Again, the experienced chemist is not confused by this conflation - but there is perhaps room to consider how learners make sense of this aspect of periodic classification and trends.

Learning about atomic structure

There are several areas where learners may have difficulty learning about atomic structure. At a relatively trivial level, learners may confuse the various entities which are posited at the molecular level, leading to non-sensible comments. For example Taber (1996a) reports students setting out on an 'Advanced level' college course as describing atoms as "a very small molecule" and "the smallest part of the atom". Taber reports references to "elements all containing the same amount of atoms in the outer shell", and a covalent bond being "when atoms are shared together in an element". In the same class of students both 'molecule' and 'ion' were described as "part of an atom": and the ion was "a charged particle that orbits an atom". Whilst these errors should not be discounted, they are perhaps best explained as *confusing* the labels for concepts, or having *insufficient familiarity* with the molecular level to clearly differentiate between the different concepts.

Other problems may be more significant. For example, Schmidt (submitted) reports students believing that the nucleus of an atom must contain an equal number of neutrons as protons, as the neutrons had the role of neutralising the protons. (This would appear to be

related to the 'label as a hidden persuader' notion discussed above.) It might be thought that if learners can correctly identify the charge on the three main sub-atomic particles they would not make such an assumption: but Taber's work (1997a, 1998a) suggests that learners commonly fail to apply conventional electrostatic principles in the atom.

For example, some students believe that the orbiting electrons *push* on the protons in the nucleus to hold the nucleus together, overcoming the tendency of the positive protons to repel. This suggestion both ignores the difference in proton-proton and proton-electron distance *and* suggests the negative electrons repel positive protons. The nature of nuclear interactions is not usually discussed in any depth in chemistry, and some (but by no means all) students clearly feel the need to explain nuclear stability. (Schmidt's students' neutralising neutrons were actually much closer to the accepted explanation than Taber's students' pushing electrons.) Taber (1998a) reports that when it *is* recognised that the electrons pull on the nucleus, it is often believed that the force applied by electron on nucleus must be *smaller* than that applied by nucleus on electron: these students do not appear to distinguish between the size *of a force*, and the size *of the effect* it produces.

Taber (1997a, 1998a) also reports a related common alternative conception ('conservation of force') of how the nucleus attracts electrons. It is assumed by many students that the atomic nucleus gives rise to *a certain amount of attractive force* depending upon its charge, which is then *shared equally* among the electrons present. Thus the increase in successive ionisation enthalpies is explained because each electron removed means a bigger *share* of the attraction for each of the remaining electrons! Interestingly, Taber suggests that some students who can normally apply Coulombic principles to situations involving interacting charges, will switch to alternative conceptions if they construe the context (i.e. the configuration of charges) as being an atom. In at least some cases this seems to be related to (what is perceived as) chemistry and physics knowledge being stored separately. Atoms seem to be exempt from the laws of physics. It may be of relevance here that Harrison and Treagust (1996) report a pupil describing the atomic nucleus as the atom's control centre.

The problems outlined above apply to learning about a simple nucleus/electron shells model of atoms. Another area of difficulty is experienced when more sophisticated models of the atom are introduced. Cervellati and Perugini (1981) found some first year Italian university students *identified* orbitals with energy levels, and others thought they were electron trajectories. Mashhadi (1994, p.6) found a similar 'mechanistic' conception of the atom with "fast moving electrons in definite orbits, similar in some ways to the planetary model of the atom", among a quarter of sample of College level (A level physics) students in England. This was also "the dominant model of the atom" found among first year French undergraduates by Cros and colleagues (Cros et al., 1986, p.308). Harrison & Treagust (2000) report how learners may have difficulty differentiating between key aspects of different atomic models, for example confusing electron shells and clouds. Tsaparlis (1997b) provides a 'post-factum' analysis of undergraduate performance on an end-of-semester examination in a quantum chemistry course in Greece. This analysis suggests that a number of aspects of the course seemed to cause difficulty for students, including definitions of atomic and molecular orbitals, the nature of the mathematical descriptions and graphical representations of orbitals, and the approximate nature of the orbital models available in many-electron atoms.

Taber (1997a) reported that U.K. A level students he interviewed demonstrated a range of alternative conceptions about aspects of the quantum model of the atom. One student had confused the emission of an electron during thermionic emission, with the emission of light during electrical transitions within an atom. When learners were first taught about

orbitals some seemed to take this term as a synonym for *shells*, and for *orbits*: so all three terms tended to be used interchangeably. The introduction of *sub-shells* and *energy levels* then provided further distinct concepts to be confused with shells, orbits and orbitals. The designations given to atomic orbitals (e.g. 2s, 3d) were understandably puzzling to students who had not been able to keep these fundamental concepts distinct.

Diagrams representing orbital probability envelopes in the conventional way caused confusion, as the envelopes were considered to be orbital boundaries. The spin of an electron was, not surprisingly, understood in the same way that a macroscopic object might spin, and indeed one of the students explained that the electron's spin was caused by the way electrons repelled each other: they spun "because they're all going to be repelling each other and circling like that, always trying [sic] to get as far apart" (p.263).

Another aspect of atomic structure of concern is the way that learners have been reported to underestimate the stability of the electrically neutral atom in relation to ions with 'octet' structures. Although in the contexts of wider chemical systems, such ions are commonly found (unlike most atoms - a point that will be relevant later in this paper), few isolated atoms would *spontaneously* acquire such octet structures. (A number of atoms have exothermic electron affinities, but this only gives octets in the halogen atoms.) Yet Taber (1997a, 1999b) reports that college students who seem to understand about ionisation energy will often still consider the sodium atom is less stable than the sodium ion, and that ionisation will spontaneously occur. More surprising, perhaps, many students consider that the Na^{7-} ion is more stable than the atom - as despite being a highly charged *anion of a metal*, it does have an octet of outer electrons (Taber, 2000f). Taber also reports that some trainee science teachers seem to share these ideas (Taber, 2000g).

Learning about molecular structure

Pereira and Pestana (1991) studied the way that secondary pupils at a Portuguese Chemical Olympiad represented molecules of water, and found that the bond angle was shown variously from 30° to 180° (although it could be argued that in a two dimensional projection the apparent angle *could* take any value), and more significantly that the O-H bond length was shown as largest in the gaseous phase and smallest in the solid state. This would seem to be an example of learners retaining some confusion between the molecular model and the macroscopic properties (see above): with an increase in volume of the substance partly due to larger molecules. Griffiths and Preston (1992) also report that Canadian (grade 12) pupils thought that molecular shape and size depended upon phase (and that molecular shape reflected the shape of the container).

A chemical bond is "the linkage between atoms in molecules and between molecules and ions in crystals" (Penguin Dictionary of Chemistry, Sharp, 1983). Without chemical bonds there would be no condensed matter, and indeed most common gases - those that are molecular, such as oxygen, nitrogen, carbon dioxide, etc. - would not exist.

From the scientific world view there is no *a priori* reason to expect atoms to stick together, unless there is some form of force attracting the particles together. The scientific model of the atom, as containing positive and negative charges, however leads to an expectation that atoms will be attracted together, due to electrostatic forces. In general the expectation is that there will be an equilibrium distance between two atoms where attractions and repulsions balance: at lesser distances there will be a resultant repulsion; at greater distances a resultant attraction.

From the scientific viewpoint then

- atoms would not be expected to be linked unless there is some form of physical (i.e. in

terms of the laws of physics) bond;

- there are net electrical forces between atoms which at most separations would tend to attract them together;
- this electrical force is the physical basis of the chemical bond.

Yet Griffiths and Preston concluded from a Canadian study of grade-12 students (16-18 years) that, for many students, the concept of bonding might have little to do with forces of attraction, and that for some students molecules were not bound due to inherent interactions between atoms, but were held together by “something external to the molecules” (Griffiths & Preston, 1992, p.620). Cros and co-workers (1986) found that, among first year French undergraduates, the interactions between atoms in molecules were often unknown (38%) or poorly known (18%), and that some students were not even aware that such interactions existed.

Wightman undertook two case studies with English classes of 13-14 year old pupils being taught about particle theory (Wightman et al., 1986). She found that pupils’ ideas included notions of bonding that were based on mechanical analogies. One pupil referred to a force being like “*elastic* holding them together and it can stretch and contract to pull the particles back together again” (p.106), and another talked about the bonding being “like *string* between the atoms sort of holding it all together” (p.291). One pupil conjectured that as the bonding was “like *glue*” thermal expansion might occur because “the bondings get thicker” (p.305). Harrison and Treagust (2000) report that some Australian pupils also consider bonds as material connections.

These examples are instances of the situation, discussed above, where learners accept a particle model, but are unable to recognise that molecular level explanations are only useful if they have a distinct nature compared to molar level explanations. Wightman’s informants were simply *transferring* macroscopic notions (glue, string, elastic) to the molecular model. Even when electrical ideas were introduced, some form of mechanical linkage was still called upon,

“we were thinking that some [of the ‘atoms’] were positive and some were negative.. Say that all the positive ones had little holes in them all the way round - all the negative ones had things sticking out of them ... the things sticking out went in the holes...”
(Wightman et al., 1986, p.198.)

Taber’s (1997a) U.K. research, with college level students, revealed that his learners (several years on from the stage described by Wightman) had mastered a model of molecules as atoms bound by covalent bonds. Covalent bonds were seen as pairs of electrons ‘shared’ by two atoms to provide them with a preferred number of outer shell electrons.

Such a model of the bond is a significant advance upon the notions of string and elastic found by Wightman. For one thing, the image of an electron pair shared between two nuclei would seem to relate to both the electrical nature of the bond, and to the bonding orbital model that would be met by these students in their course. The use of the ‘octet rule’ (i.e. that stable chemical systems can usually be conceptualised as arrangements of atoms having noble gas electronic structures), enabled these learners to recognise the existence of discrete molecules, and to work out the appropriate molecular formulae in many cases.

However, Taber also found that student commonly had difficulties moving *beyond* the notion of the shared pair of electrons: which was commonly seen as something more than an image or metaphor. As Tsapalis (1984: 677) points out: “the mere statement that the chemical bond is due the sharing of electrons by the atoms is far from being satisfactory from the pedagogical point of view”. Yet for some of Taber’s interviewees the shared electron pair

was the bond (not an aspect of a stable configuration of atomic cores and electrons), and in covalent bonding the sharing was equal.

This view made learning about polar bonding difficult. 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 something between covalent and ionic). So Barker and Millar (2000) report that U.K. A level students believe that hydrogen fluoride exists as dissolved *molecules* in aqueous solution.

Taber also found that because the shared electrons were seen to still 'belong' to specific atoms, bond fission was often assumed to be homolytic - as each atom would *want* to get 'its own' electron back. Taber found such anthropomorphic language to be very common (Taber, 1997a, 1998b; Taber & Watts, 1996). Molecules were usually assumed to arise from discrete atoms (although it is unlikely the students had ever met any reactions where the reactants were in this form!) because the atoms *wanted* or *needed* to obtain 'full outer shells'.

Again, such expressions were not generally being used as some form of shorthand for processes that might happen due to physical forces. An explanatory principle that atoms strived to achieve full shells (based on the octet rule, but imbuing it with causative power) had replaced the younger pupils' ideas about string, glue and elastic. For some students a shared electron pair held atoms together *because* it enabled them to have octets of electrons. Barker and Millar (2000) report a similar finding when they suggest that the notion of *valency* is imbued with an explanatory or causative power by learners - leading to anthropomorphic explanations in terms of atoms wanting to form certain numbers of bonds - where this might better be seen as a re-description ('explanation by description') or pseudo-explanation (Taber & Watts, 2000).

As might be imagined, Taber found that the existence of bonding which did not lead to atoms having full electron shells was often something of a mystery to students: although the formula of an 'electron deficient' compound such as BeCl_2 or AlCl_3 might make sense in terms of the metal atom forming as many bonds as its valency allowed, this does not explain why sulphur would want to go beyond SCl_2 or SF_2 to give SCl_4 or SF_6 !

These students also had difficulty appreciating why the chlorine atom in AlCl_3 would 'want' to share an electron pair to form a dative bond, as it already had the electrons it 'needed'! For a student who *does accept* the existence of such dative bonds, the common assumption is that the electron pair is shared equally (despite significant differences in electronegativity) as a dative bond was considered to be a covalent bond: for other students it is not really a chemical bond at all, but just a force or interaction.

Molecular orbitals

For a learner operating with the notions of shells and electron *orbits* (see above) bonding electrons may seem to be *more* confined than other valence electrons, as they must remain in the area of overlap of two shells (Taber, 1997a).

Taber found that (U.K. A level) students became confused between the mathematical modelling of molecular orbital formation (i.e. LCAO), and the orbitals themselves: giving references to 'linear orbitals'. Taber also found that students confused molecular orbitals with atomic orbitals: suggesting electrons in bonds in molecules were in orbitals they designated as 's' or 'p', or confusing sets of rehybridised molecular orbitals (e.g., sp^3 hybrids) with

molecular orbitals. Taber reports that when a learner who had grasped the curriculum scheme for rehybridisation and orbital overlap (to give molecular orbitals) was asked about the orbitals present in a molecule, they would still suggest the atomic orbitals of the appropriate isolated atoms.

Presumably, learning the abstract scheme for atomic orbital occupation ($1s^1$; $1s^2$; $1s^2 1p^1$; $1s^2 1p^2$; $1s^2 1p^3$; ...) requires some considerable effort, and once the scheme is acquired it acts as an example of a pedagogic learning impediment (see above), tending to 'come to mind' more readily than ideas about molecular orbitals.

π -bonds seem to provide a particular difficult concept for students, presumably because the practice of representing orbitals with probability envelopes *suggests* two non-connected areas of electron density. One student explained σ -bonds as being a simple overlap of atomic orbitals, *whereas* the π -bonds were like a 'hamburger'! Another described the delocalised system in benzene as having two π -bonds, one *above* the plane of the ring, the other *below*. When students *did* understand the notion of rehybridising orbitals to provide atomic orbitals better suited to overlap, they might assume that rehybridisation was *always* required.

As might be expected, Taber found evidence of learners having difficulty conceptualising resonance. In particular students considered the resonance to mean an *alternation* between the canonical forms, rather than something intermediate to them. Tan and Treagust (1999) report an even more gross error: that learners may believe that it is *atoms* that are delocalised in such structures as graphite, rather than electrons. The difficulty in appreciating molecular bonding concepts has also been noted by Coll and Taylor (in press) who report one New Zealand student describing anti-bonding orbitals as 'silly things' that just stuck out!

Learning about lattice structures

This section considers aspects of learning about structure in different types of solid. In their study of students entering University to study science, Cros and co-workers found that crystals "remained a mystery for most" (Cros, et al., 1986, p.309):

"When asked about the interactions in a crystal 42% of the students did not reply, and 15% gave incorrect or completely inadequate information. Only 27% of the students referred to a clearly defined arrangement of atoms or ions."

After one year of University study the interactions within the crystal were only described as "somewhat less mysterious" to the students (Cros et al., 1988). Mention of electrostatic interactions increased: although only from 8% to 18% of the sample!

Macromolecular solids

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). Taber (1997a) found that one of his student interviewees considered solid carbon as comprising as 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, 2001b; c.f. Nelson, 1996).

Ionic bonding

Butts and Smith (1987) undertook an interview study in Australia to follow up a survey finding that *the difference in properties between ionic compounds and molecular compounds* had been rated as a difficult topic by students. They found that most of the year 12 students in their sample associated sodium chloride with ionic bonding, which is appropriate, but that the students often *also volunteered a description of how ions might be formed through an electron transfer event* (i.e. from sodium atom to chlorine atom) which could result in the formation of the bond.

It was also common for 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 and Smith reported that some of the students thought that this assumed molecular nature of sodium chloride explained why the solid did not conduct electricity: as ions were only formed from the molecules on dissolving (p.196).

When the Australian students were shown physical ball and stick models of the ionic lattice they did not always appreciate the network of ionic interactions: one student thought that the six wires represented *one* ionic bond, and *five* "physical" bonds (p.196). Another student expected each chlorine ion in the lattice should have seven wires attached "because chlorine has seven electron in its outer shell" (p.196).

Students in Taber's U.K. study often had quite similar ideas about ionic materials to their Australian counterparts (Taber, 1994, 1997a, 1998b). It was common for the English students to consider NaCl ion-pairs within the lattice *as if* they were molecules (with some actually using this term). Although 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, with similar results reported from other studies in the U.K. (Barker & Millar, 2000), Australia (Harrison & Treagust, 2000) and Singapore (Tan & Treagust, 1999).

Like the Australian interviewees, Taber's English students tended to present descriptions (both in words and diagrams) about electron transfer when asked about ionic bonding. Indeed, Taber found that ionic bonding was often *identified* with a conjectured (although chemically unlikely) electron transfer event between isolated atoms. To many students the ionic bond in sodium chloride is *defined* as an electron transfer from a sodium atom to a chlorine atom.

As in the Australian study, again, English students suggested that each ion in the sodium chloride lattice could only form one bond. Taber related this to the same 'full shells' explanatory principle that students used to explain covalent bonds (see above): a sodium atom was considered to 'need' to donate one electron (and therefore formed one bond) and a chlorine atom was seen to 'need' to gain one electron (forming one bond). The bond would therefore only exist between ions that had transferred electrons.

There is some suggestion that these common alternative conceptions may survive University teaching. Oversby (1996) found that some of his trainee (post-graduate) chemistry teachers considered the students' alternative conceptions of the ionic bond reported by Taber (1994) to be an acceptable model of ionic bonding - i.e. that ions could only form the number of bonds allowed by their valency, that a bond only exists where an electron has been transferred, and that ions that were *not bonded* by electron transfer were held together by *forces instead*.

Taber (1997b) summarised his findings in terms of a comparison between the

scientific ('electrostatic') model for ionic bonding, and a common alternative ('molecular') framework which learners commonly seemed to draw upon, as shown in Table 1. Survey data suggested that many learners seemed to draw upon both frameworks, and should perhaps be considered to be in transition between the two models. Further, the pattern of responses showed that some learners were selecting inconsistent and contradictory responses. Taber suggested that this could be interpreted in terms of learners operating with multiple frameworks (Pope & Denicolo, 1986). Although the evidence from the survey could not in itself be considered to justify this claim, case studies of individual learners did show that learners held manifold conceptions for particular concepts such as the ionic bond (Taber 1997a, 2000b; Taber & Watts, 1997).

TABLE 1: *An alternative framework for ionic bonding (from Taber 1997b).*

	molecular framework	electrostatic framework
<i>status</i>	<i>alternative framework</i>	<i>curricular science</i>
role of molecules	ion-pairs are implied to act as molecules of an ionic substance	ionic structures do not contain molecules - 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. (e.g.: 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 (e.g.: the co-ordination is 6:6 in NaCl)
history conjecture	bonds are only formed between atoms that donate / accept electrons. (e.g.: in sodium chloride a chloride ion is bonded to the specific sodium ion that donated an electron to that particular anion, and <i>vice versa</i> .)	electrostatic forces depend on charge magnitudes and separations, not prior configurations of the system (e.g.: 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. (e.g.: 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 (e.g. the forces between a chloride ion and each of the neighbouring sodium ions are equal.)

Metallic bonding

Taber (1997a) reported that U.K. students starting out on a college chemistry 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 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 by sharing (covalent bonding) or transferring (ionic bonding) electrons. As with ionic bonding, some students assume metals are molecular (something also found by Fleming, 1994; de Posada, 1997; Harrison & Treagust, 2000; Coll & Taylor, submitted). Students who were unable to make sense of metals in terms of either pattern (e.g. as one informant explained “*it’s not ionic, and it’s not covalent either, it’s like, it’s hard to explain this*”) might conclude that there was no bonding in pure metals, or that there was a ‘lesser’ form of bonding (also reported by Coll & Taylor) - that was 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”

(Taber 1997a, p.354)

Other students found 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 took turns in having full shells (either by gaining enough, *or* losing enough electrons); or the electrons were considered to have been *transferred* to the lattice so that the atoms could gain a full shell.

Taber found that some students would describe the bonding in metals in terms of the ‘sea’ of electrons: but that often they had learnt the term with little understanding of this model. Some student diagrams showed the ‘sea’ as a vast excess of electrons, and that 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.”

Intermolecular bonding

Three of the students interviewed by Butts and Smith (1987) did 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. These students thought that a grain of sugar was a single molecule, and had a giant structure like diamond. This way of thinking can lead to the common idea that covalent bonds must be relatively weak as covalent materials usually have low boiling temperatures (Barker & Miller, 2000), and the covalent bonds break on change of state (Tan & Treagust, 1999).

In their study of Canadian students, Griffiths and Preston reported that some thought that the molecules in ice were not bonded in any particular pattern (Griffiths & Preston, 1992).

Taber’s U.K. students generally had difficulty accepting anything that was not clearly

explicable in ‘octet’ terms as being a chemical bond. So whereas covalent bonds (seen as sharing electrons to get a full outer shell) and ionic bonds (seen as a transfer of electrons to get a full outer shell) clearly counted, forms of bonding such as hydrogen bonding, and van der Waals’ forces could not be understood within such a scheme. Taber (1997a, in press) describes students’ progress in developing a broader category of chemical bond, and the way in which the status of intermolecular bonding fluctuates between being ‘just a force’ and accepted as a genuine form of bonding.

The status of hydrogen bonding was sometimes found to follow a trajectory from chemical bond, to just a force and back. Taber (1997a) discovered that some students first heard about hydrogen bonding in biology classes, where they took the meaning to be ‘a [covalent] bond to hydrogen’ (such as that in methane). When they came to understand that this was not what a hydrogen bond was, it could lose its status as a chemical bond.

Barker and Millar (2000) also report that U.K. A level students did not class the hydrogen bond as a bond, and suggested that this could derive from the way their teachers referred to hydrogen bonds as attractions, and Fleming (1994a) reported 15 year old New Zealand pupils who did not think there was any bonding between the molecules in water. Pereira and Pestana (1991) found that Portuguese pupils’ representations of hydrogen bonding in water sometimes showed the bonds between hydrogen atoms in different molecules, and sometimes showed the hydrogen bond to be shorter than the *intramolecular* bond.

Taber reported that the ambiguity of the status of intermolecular bonding was demonstrated in the context of discussing Raoult’s law: where variation from ideal behaviour could be explained in terms of the *difference in bonding* in a mixture compared with the pure liquids: Yet the solvent-solute interactions and intermolecular bonds were not considered to be real bonds in other contexts.

SOME PEDAGOGIC OBSTACLES TO LEARNING ABOUT CHEMICAL STRUCTURE

This survey of learners’ attempts to master the structural concepts of chemistry leads to some suggestions of ‘pedagogic learning impediments’ - that is, blocks to learning that derive (at least in part) from the way in which the subject is taught.

In the preceding sections, reference has been made to the ‘octet framework’ (Taber, 1997a, 1998b), a common alternative conceptual framework that is drawn upon by many learners in chemistry (see Figure 4). Although not all learners are likely to develop the full range of ideas, certain key aspects (in particular the striving to obtain octets/full shells) were found to be both common and tenaciously held by learners.

As is suggested by Figure 4, the framework can provide an extensive and relatively coherent theoretical structure in which to ‘understand’ chemistry, albeit to understand chemistry *differently* to the teacher. Consequently, a student who understands chemistry in terms of the octet framework is able to interpret much new information in a manner consistent with existing ideas, and the framework therefore interferes strongly with some aspects of learning college chemistry.

Although it is not suggested here that teachers explicitly teach alternative notions (although some may well do!, c.f. Oversby, 1996; Barker & Millar, 2000) such as molecules of NaCl, or the spontaneous ionisation of metallic atoms, it is clear that such a common set of ideas must derive from somewhere. Perhaps part of the appeal of these ideas lies in aspects of brain structure (being biased towards certain types of explanations - either genetically, or

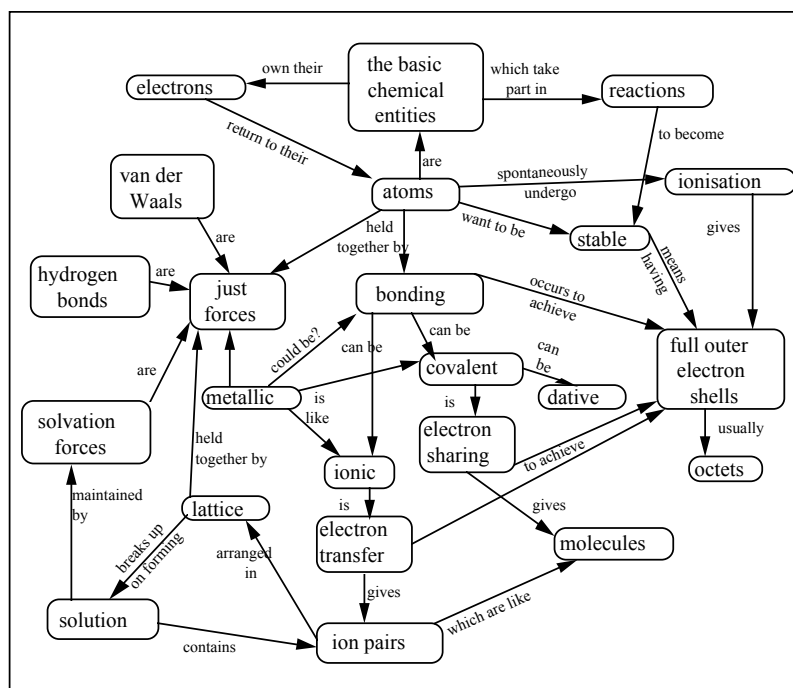


FIGURE 4: An alternative conceptual framework (from Taber, 2000c).

through early life experience), or of common metaphors etc. in our culture: *but clearly teaching must play a part.*

The acquisition of so many specific ideas about atoms and electrons is only likely to have occurred through science teaching, and therefore even if teachers are not actually presenting dubious ideas, something about *the manner of presentation of the subject* leads to learners developing such dubious ideas. One of Taber's students was quite explicit about the way that he felt some aspects of school chemistry ('G.C.S.E.': General Certificate of Education) blocked his subsequent college study.

"going through my course, 'if I hadn't have done' some aspects of chemistry G.C.S.E., I would have found this easier to understand maybe, because what they taught us at G.C.S.E. and what they teach us now contradicts itself, as it were, and it's harder for you to understand, 'cause they ham[mer] it into you that you have to learn this for this exam, and then you learn it and then you remember it, and then when I do *this* course ... I always think of that thing that I learnt for G.C.S.E. and it sort of like clashes, therefore like it's harder to remember sometimes." (Taber 1997a, p.628)

This type of 'pedagogic' learning impediment (Taber, accepted for publication) echoes the type of 'epistemological obstacle' that Bachelard (1968 {1940}) suggested inhibits scientists - where their current understanding of a concept area includes the vestiges of historical models that impede their appreciation of the most current thinking.

Taber (1995d, 1997a) has discussed pedagogic learning impediments in the context of learning about concept areas such as bonding, ionisation enthalpies and the driving force and mechanisms of reactions (see below). However, similar concerns may be found when considering learners' models of the atom. Taber's (1997a) work - discussed above - suggests that having learnt to think about atomic structure in terms of electron shells may impede learning about orbitals, and that learning the details of shapes and designations of atomic orbitals then acts as an impediment to thinking about molecular structure in terms of molecular orbitals.

Harrison and Treagust (2000) suggest that pupils' models should be considered to have a 'sell-by date' (i.e. that they should become obsolete as the learner studies more detailed models), and that - when they continue to be used beyond this - they may encourage the learners to rely on alternative conceptions. For example, Gillespie (1996) has suggested that the common use of orbital diagrams to illustrate rehybridisation encourages students to see this as a physical process (with redistribution of electron density), rather than as a formal mathematical operation.

Jones has argued that the failure to attempt to teach quantum mechanics in some form from early in the science curriculum means that by the time students are introduced to quantum theory they are already *so familiar* with classical mechanics that they develop "an uneasy hybrid" of the two perspectives (1991, p.93). Jones suggests that this leads to "half-baked and incorrect conceptual models *which stunt understanding* and the development of interest" (p. 93, present author's emphasis). Students' familiarity with classical mechanics, and the usual approach of introducing quantum theory through the models of the first two decades of the twentieth century (when the scientists themselves were still trying to move beyond their classical notions) acts as a pedagogic learning impediment. One of Mashhadi's sample of Advanced level students explained that they had been taught about electrons *as particles* from early in secondary school, and about light *as being a wave* from even earlier, and "you have a long time to think of one thing before it is even mentioned that it is possible that may not be completely true" (1991, p.8).

However, whilst some authors recommend an *earlier* introduction to ideas from quantum chemistry, there is also a view that orbital concepts should be *avoided* completely for longer (Tsapalis, 1997a; 1997b), with a conceptualisation of atomic and molecular structure in terms of electron pair domains (Gillespie, 1996) used as the preferred approach in senior secondary and introductory university levels.

Of course, as this shows, chemistry is a complex subject, and whilst it is easy to offer criticisms of common practice, it is always possible that substantial changes could lead to even greater problems. However, some of the suggestions offered here may be seen in terms of needing *a shift in emphasis*, or a rethink of sequencing, rather than a wholesale revolution in teaching of the subject.

An atomic ontology

It was reported above that research suggests that learning about the electronic structure of isolated atoms may act as an impediment to understanding the nature of molecular orbitals. Van Hoesve-Brouwer (1996) has gone so far as to reject "atomic structure as an appropriate starting point for teaching chemical bonding". She explains that,

"Atomic structure is an unsuitable basis for teaching chemical bonding in secondary schools because there is a gap between the models of free and bonded atoms that cannot be bridged."

(Van Hoesve-Brouwer, 1996, p.157)

Tsapalis (1997a, p. 924) has argued that the molecule should be understood as "the smallest bit of matter that retains certain (but not all) properties of the pure substance" *before* it is described as being comprised of atoms. Taber (1996b, 1997a, 2000h) goes further than this, and has argued that molecules are usually best conceptualised as systems of cores (positively charged spheres comprising of nuclei surrounded by symmetrical electron density) and valence electrons.

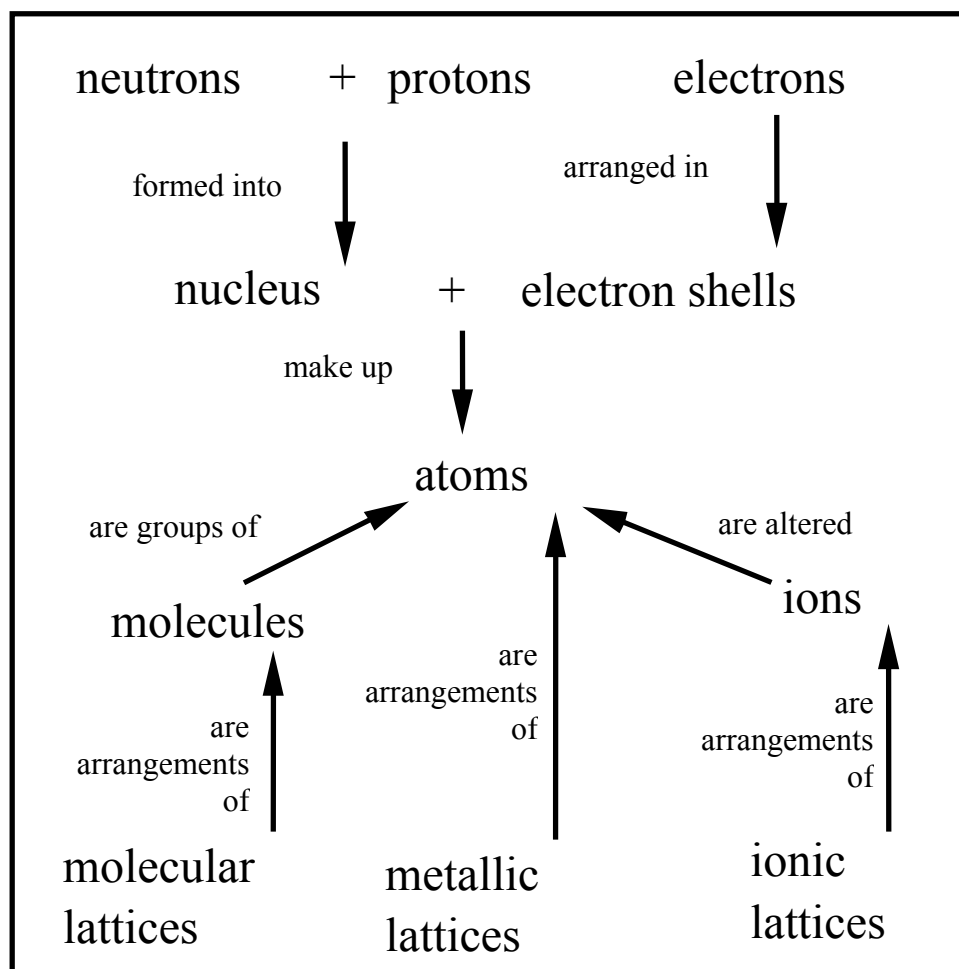


FIGURE 5: An atomic ontology (from Taber, 2000h).

Understanding molecular structure is then seen as understanding the configuration of cores plus valence electrons. The ‘quantaction’ (Figure 3) that (at the molecular level) explains chemical reactions (at the molar level) may be seen as interactions resulting in reconfigurations of the cores and valence electrons (Taber, 2000h, but see also: de Vos, 1990). Yet, in practice, molecules are almost always conceptualised by learners as atoms stuck together (Figure 5).

The concept of the atom is seen as central in chemistry, and, Taber has suggested that it may actually be *over-emphasised* in chemistry teaching to such an extent that it becomes unhelpful to the learner. So learners may make contradictory claims that atoms are the smallest possible entities, in the same sentence as describing their internal structure (Taber 1997a), or describe an atom as being the ‘smallest particle of matter’ in labelling a diagram showing the atom as an assemblage of sub-atomic particles (Harrison & Treagust, 2000).

Part of the alternative conceptual framework Taber (1997a, 1998b, 1999a) described reflecting common aspects of learners’ alternative ideas about chemical processes and chemical bonding - the octet framework (see Figure 4) - was labelled ‘the atomic ontology’. Put simply, this is the idea, taken *literally*, that everything is made of and from atoms.

Whilst this may appear to be a valid chemical principle, Taber argues that pupils learn to always start thinking from atoms. In particular, when considering chemical reactions, learners often make *an assumption of initial atomicity*. So, for example, a reaction between sodium and chlorine is assumed to involve sodium atoms and chlorine atoms; a reaction

between oxygen and hydrogen is assumed to involve oxygen atoms and hydrogen atoms, etc. (Teachers and texts encourage such a perspective when they present diagrams showing separate atoms representing reactants that are actually molecular, or found in the form of metallic lattices etc.).

The ‘assumption of initial atomicity’ means that learners may ignore the energy required to break bonds during a reaction (e.g. Barker & Millar, 2000), and that most chemical reactions come to be explained in terms of the full shells explanatory principle - that the reaction occurs *so that the atoms* present can get full shells. Of course, a more realistic representation of the reactants at the molecular scale would show that both reactants and products usually have stable electronic structures!

The atomic ontology is also considered to encourage learners to rely on anthropomorphic explanations. Apart from arguments about atoms ‘needing’ to ‘share’ or ‘donate’ electrons, it encourages the notion that an electron is a part of a (particular) atom. Learners often see the electrons ‘shared’ in a covalent bond as still *belonging* to specific atoms: with the consequence of expecting all bond fission to be homolytic - with each atom getting ‘its own’ electron back.

Some students (Taber, 1997a, in press) have been found to hold similar ideas in the case of ionic materials: that in a double decomposition reaction the ions from the reactants have to return electrons to their original (their own) atoms, before a new electron transfer can occur to form the precipitate! So it is argued that in a silver chloride precipitate, the chloride ions received their ‘extra electron’ from the silver when the ions formed - even though those same ions had an extra electron from a completely different species shortly before! One aspect of ‘the atomic ontology’ is that the *history* of particles (e.g., ‘which atom did this electron come from?’) is given undue importance.

The over-generalisation of the octet rule

In the absence of an appreciation of the role of electrical interactions in chemical structures, and chemical processes, students commonly adopt the octet rule as the basis of a general explanatory principle.

The octet rule - a consequence of the effect of electrical interactions in the quantized atom - is a useful ‘rule of thumb’ in elementary chemistry. Using the octet rule (that the electronic structures of stable species are usually such that each atom has the same number of surrounding electrons as a noble gas atom: 2, 2.8, 2.8.8 etc.) learners can identify the formulae of many stable ions and molecules. So, for example, the learner knows that O^{2-} and NH_3 are more likely to be found than O^{3-} or NH_6 !

However, pupils commonly seem to *over-generalise* the rule from being a way of identifying likely stable species, to become a general purpose *explanation* for why reactions occur. So hydrogen is said to react with fluorine because the hydrogen and fluorine atoms *need* an extra electron each to satisfy the octet rule to be stable - and so they share! Of course, this explanation only works when the assumption of initial atomicity is made - but research suggests that learners readily do make this assumption, sometimes even when the molecular nature of the reactants is made explicit (Taber, 2000d). It seems once this way of thinking about reactions has taken hold, it is applied with little thought given to the contrary evidence!

A dichotomous classification of bonding

Another of the themes deriving from this survey of research is the idea that learners often consider that there are two types of chemical bonding: covalent and ionic (Taber, 1997a):

<i>covalent</i>	<i>ionic</i>
electrons are shared between non-metal atoms	electrons are transferred from metal to non-metal atoms

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.

For one thing learners have difficulty making sense of bonds that are 'in-between' ionic and covalent, i.e. polar bonds: after all, what exactly does something *between* 'electron sharing' and 'electron transfer' mean? So pupils may be told that 'dative bonding is just like a covalent bond, except both electrons come from the same atom', and if they take this literally they may not realise that dative bonds tend to be polar.

Other types of bonding may prove even more problematic. Metallic bonds, hydrogen bonds, van der Waals' forces, solvent-solute interactions etc. are often either conceptualised as modified versions of covalent or ionic bonds, or are discounted as 'just forces' (Taber, 1997a).

Ionic bonding understood by analogy with covalent bonding

It seems common that learners are taught about covalent bonding *before* they learn about ionic bonding. As has been noted above, pupils often seem to define the ionic bond in terms of electron-transfer, as compared to electron-sharing in covalent bonding. If the covalent bond is taught first it is natural that it then becomes a model for understanding ionic bonding. In the covalent case the number of bonds formed is (usually) determined by the (co)valency of the atom: i.e. how many unpaired electrons are available to pair up.

In the case of ionic bonding the notion of electrovalency is important, as it determines the charge on the ion (which in turn is important for the stoichiometry of the compound, which is neutral overall). However, it appears that learners commonly associate the ionic charge with the *number* of bonds formed. So in sodium chloride, it is commonly believed that each ion has one ionic bond, as the valency, rather than the co-ordination number, is assumed to determine this.

In the covalent case the bond is (in a simple sense) the pair of electrons between two atomic cores (the 'shared' electrons). Again, *by analogy*, the electrons which are *conjectured* to have been transferred are seen to have a key role in the ionic bond, whereas from the scientific viewpoint it is the ionic charges that are significant. (I use 'conjectured' because this whole mental image assumes that an ionic substance has been made fresh from its elements. It is unlikely, in many cases, that the 'extra' electrons on the anions came from the cations in the lattice: and - of course - it is completely irrelevant to the properties of the bond whether they did or not.)

Pre-occupation with ion formation

This focus on the (non-existent) role of transferred electrons in ionic materials relates to a pre-occupation with ion formation. Learners will commonly respond to a request for a diagram of *ionic bonding* with a neat picture showing *the formation* of a single sodium ion, and a single chloride ion, from isolated atoms of sodium and chlorine. Such pictures are chemically unrealistic representations of a chemically unlikely reaction (there are easier ways to make sodium chloride than binary synthesis) which is irrelevant to the ionic bond within a crystal lattice. Yet such a characterisation is harsh on the learner because this type of diagram is common in text books, and even expected by some examination boards.

Social chemistry?

One aspect of learners' explanations in chemistry, which is of particular note, is the common use of anthropomorphic language (Taber, 1993, 1997a; Taber & Watts, 1996). Such language was found to be used particularly heavily when college students discussed bonding - something that has also found to be the case among University students (Fleming, 1994) - but was also found in other contexts. So electrostatic interactions might be explained in terms of electrically charged species *trying* to attract together, or *trying* to move apart, and electrons would be said to *try* to get into their ground states!

Taber found that the use of such language was particularly prevalent when students explained *why* they thought reactions occurred, and *why* bonds were formed. It was very common for students to explain that atoms *want*, or *need* to get stable electronic configurations. Atoms - according to students - *like* to be stable, *wish* to be stable, *prefer* to be stable and indeed can be *very eager* to be stable. They *try* to become stable by forming bonds, and will *think* they are stable once the bonds were formed. Atoms - it is claimed - *try* to achieve full orbitals, as that is when they are *happiest*.

Taber reported that most of the students he interviewed used this sort of language to describe chemical processes, and he probed the extent to which the learners were comfortable with explaining chemistry in this way. He reports examples of both a student who happily used language of this type but was *aware* that it was figurative, and of a classmate who denied that atoms had feelings *when asked directly* - but who was comfortable with the most blatant anthropomorphisms (atoms *enjoying* and *desiring* and being *jealous*). It is of interest to note that Griffiths and Preston (1992) report that over half of their sample of Canadian grade 12 pupils thought *atoms were alive*, although their research methodology did not allow this belief to be probed further. Harrison and Treagust (1996) reported that a fifth of a sample of secondary pupils in Australia thought that atoms were alive, and grew and divided.

It seems that when learners are not aware of, or do not understand (or perhaps, are not fully convinced by) the physical forces at work during chemical processes, they are quick to use 'social' language to talk about those processes. The talk of atoms behaving in certain ways to achieve personal goals provides a narrative which enables a rationale to be discussed. However, Taber's data also suggested that as learners become comfortable in describing and explaining chemistry in this way, and become more familiar with their narratives, there may well be no 'explanatory imperative' to search for a more physical 'story' (Taber & Watts, 2000).

The way of talking (and thinking?) can become so ingrained that it acts as a block on learning more scientific ways of thinking. Taber reported an example where he challenged one student's acceptance of explanations in terms of an atom that would *prefer* to have eight electrons; would *want* to have eight; and therefore *wanted* to get another atom's electron.

When the sentience of the atom was queried the student admitted that the atom did not know how many electrons it had: but later in the interview the student returned to explanations in terms of what atoms *wanted*.

Despite considering the habitual use of anthropomorphic language as an impediment to further learning, Taber was equivocal about whether such language should be completely discouraged by teachers. He distinguished between weak forms of anthropomorphism, where language is intended metaphorically, and stronger forms where the meaning is literal.

Teaching can be seen as making the unfamiliar familiar (Taber, submitted), and one way the very abstract world of electrons, orbitals and molecules can be made familiar is to tell the stories by analogy with more familiar social interactions. So, for example, Benfey (1982) refers to molecules having a life-history, being born and dying in an analogous way to living creatures, and Tsapalis (1984) suggests an atomic tug-of-war analogy for the chemical bond. At one level Taber's students were all aware that atoms did *not* have psychological needs and social lives, and some of his students were simply using language metaphorically to stand in the place of the more mysterious physical processes. However, in some of the students the metaphor had become 'invisible' to the user, and explanations in terms of what the atom wanted came to be used as if they were literally correct.

Taber's research implies that the use of anthropomorphic language may be a good thing in the 'familiarisation' stage, where learners are becoming aware of the descriptions of molecular level processes: however, once the narratives have made the processes familiar, the metaphorical language needs to be substituted by the more scientific language of physical forces, *before* it becomes habitual.

RESEARCH-BASED PRACTICE: SOME ADVICE TO CURRICULUM PLANNERS, AUTHORS AND TEACHERS

In this paper I have reviewed some of the research about learners ideas about structural aspects of chemistry, and have identified a number of common 'learning impediments' that may be considered to be pedagogical - in the sense that they must derive to some extent from the way chemistry is currently taught and learnt. Consideration of these learning impediments leads to a number of suggestions for the teaching of the subject:

(1) Build on physical principles

Firstly it should be remembered that, notwithstanding the emergence of genuinely chemical concepts at the level of analysis that chemists are concerned with, our subject is based around physical principles. The structures that we study in chemistry can be explained in terms of concepts such as electrical forces, and the principles of quantum mechanics.

Yet we find that many pupils and students either do not share this underlying physical framework, or - at least - do not apply it in chemical contexts. Many seem to explain chemistry in terms of 'magical' concepts such as the 'octets', and almost vitalist concepts such as the needs and desires of the individual atoms.

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 at work. In particular, at an introductory level, the electrical interactions within and between molecular level systems need to be emphasised at all times.

(2) Focus on molecules and ions not atoms

The atom has a privileged place in the teaching and learning of chemistry, yet discrete atoms are seldom featured in significant chemical processes. Most real chemistry involves molecules, or ions, or more extensive systems. Some significant chemistry does involve atomic radicals (such as $\bullet\text{Cl}$), but even in these reactions the atoms do not feature in the initiation step. Yet pupils and students often assume that chemical reactions occur between atoms - even when they 'know' that the reactants are molecular!

This - at first sight - strange finding is perhaps less so when the frequency with which pictures of separate atoms are shown involved in chemical processes in text-books and examinations questions is considered. Presumably many teachers *also* draw such diagrams when talking about processes such as ionic bond formation. As well as being unrealistic, such a scheme supports the development of an explanatory principle based around the notion of *atoms reacting [sic] to obtain full shells or octets*.

Clearly it is sometimes more time consuming to draw molecules than single atoms, and this may result in more complex diagrams being presented to learners. Yet schemes showing interactions between discrete atoms usually have no chemical validity or relevance, and may support the acquisition of inappropriate ideas.

In view of the mooting of the atomic ontology (Figure 5) as a serious learning impediment, it is suggested that it is pedagogically more sound for teachers to conceptualise chemical structures from an alternative perspective (Figure 6) which does not give atoms an undue emphasis in the molecular menagerie.

(3) 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. Although some aspects of chemical bonding are abstract and complex (such as the effects of spin-pairing electrons, or the molecular orbitals in benzene), the most important idea at an elementary level is that bonds can be understood as electrical interactions.

Bonds hold molecules together (both internally, and to one another), because to separate the parts of the system work must be done against electrical forces. Bonds form because of those electrical forces. Of course these processes can be understood in more detail, and a school pupil cannot 'see' that the configuration of water molecules is at a lower electrical potential than the mixture of separate oxygen and hydrogen molecules (and nor can the teacher, of course). Yet it seems appropriate to teach bonding in electrical terms, at some 'optimum level of simplification' (Taber, 2000e) 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. It seems perverse to focus on *ion formation* in terms of electron transfer to explain chemical bonding. In the reaction between methane and oxygen we do not begin by speculating on the genesis of the methane and oxygen molecules: we take the molecules of the reactants as a given *in that process*. In a similar way, 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 came into being! Yet many learners have been found not only to associate, but *to identify*, the ionic bond with an electron transfer between discrete atoms.

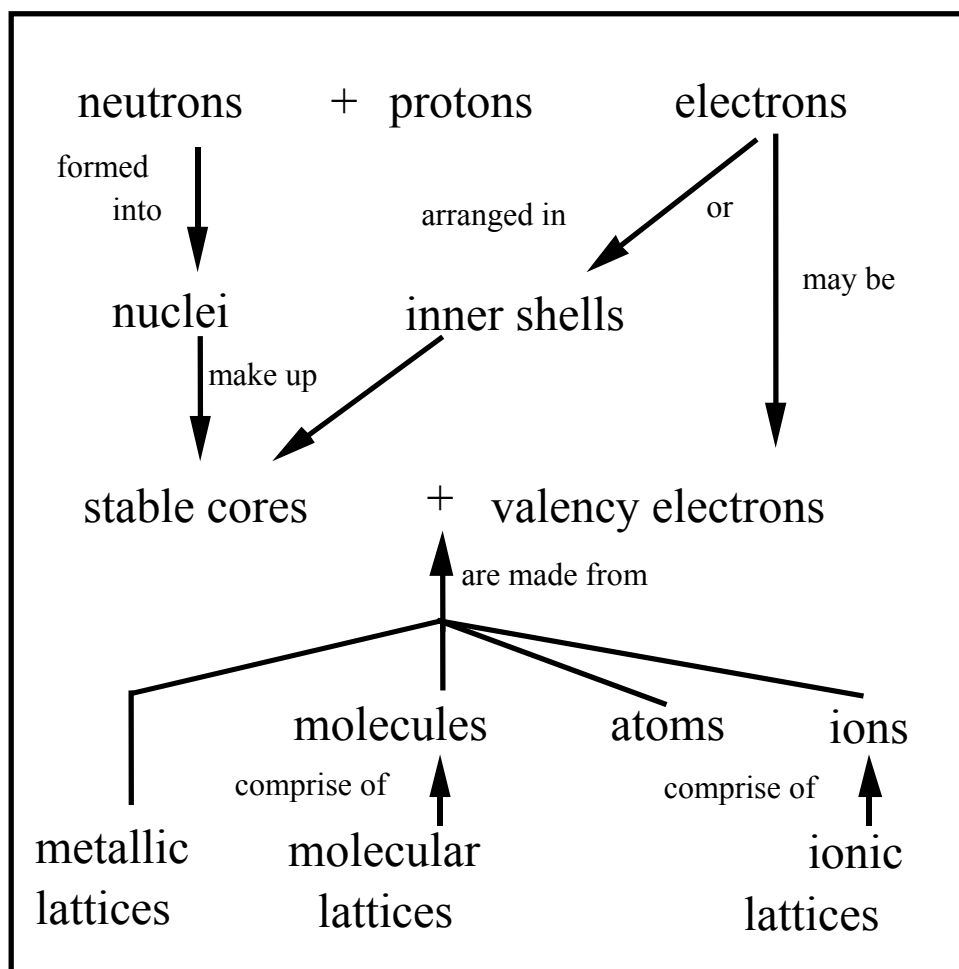


FIGURE 6: An alternative chemical ontology (from Taber, 2000h).

When bonding is seen as a primarily electrical phenomena there is no need to explain it in terms of the *needs* of atoms. If the ionic bond is understood in electrical terms learners have no need to find additional explanations for how an ion is 'also' attracted to other counter-ions other than the one in its own 'molecule' that it has exchanged an electron with!

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 bonds. 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 the electron originated in that atom. They can then keep an open mind on where electron will go on bond fission - so heterolytic fission becomes a conceivable process in certain circumstances.

(4) 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 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 to consist of discrete molecules of similar atoms, in a similar way to iodine or phosphorus. In the ionic case valency is seen 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 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. 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 (see Table 2).

TABLE 2: *A teaching order for solid structures.*

<i>type of structure</i>	<i>bonding</i>	<i>comments</i>
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.

(5) Take care with language

A final suggestion in teaching structural aspects of elementary chemistry is to take particular care in the use of language - bearing in mind both that learners may lack the assumed background, and that they may often have alternative conceptions such as, or similar to, those discussed in this review.

The need, when explaining chemistry, to shift between the molar and 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*.

We must be careful to ensure that substances, and not molecules, are said to evaporate, melt, expand, reflect, conduct and so forth. Samples of chemical substances *react* together: their molecules *quantact* (or whatever term we collectively agree to use!)

Bonding, a key structural concept area, has a peculiar position in this scheme: for we talk about ionic and covalent bonds *as if* they are macroscopic properties of materials. One could be pedantic, perhaps, and suggest that *bonds* exist at the molecular level, whereas the type of *bonding* is a property of the substance? However, I am not sure we would all agree this usage. We can define a substance as being a metal in terms of macroscopic properties (i.e., it conducts electricity, is sonorous, has lustre, is ductile etc.), but our molecular level models have become so significant in our thinking that a metal is possibly now considered primarily in terms of being a substance with metallic bonding.

In the area of discussing bonding then, we have to be particularly careful to ensure we are clear about the level of description we are employing (Selley, 1978). Perhaps all chemistry teachers should practice a mantra of key terms until they become habitual indicators of the transitions in our explanations:

“...and at the molecular level this is explained in terms of...”

“...which leads to the substance on the molar scale being...” etc.

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 displacement reactions can be *explained* in terms of what happens to a nerdy boy with a gorgeous girlfriend when a hunk come along, or that electrons tend to occupy distinct orbitals in a subshell for the same types of reasons that people getting on a bus usually avoid sitting next to a stranger. Yet these everyday scripts can provide familiar images that give pupils 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: and *not* the starting point of an extensive explanation of chemistry in terms of atoms and molecules as social agents with psychological drives. It is suggested that whenever anthropomorphic language is used by the teacher it is important to follow it with subsequent reiterations in more formal scientific terms.

This requires particular care where a dead metaphor is used: so for example, the analogical nature of describing a covalent bond as a *shared* pair of electrons has probably become invisible to most experienced teachers and chemists. We know in what ways the electron pair is ‘shared’ (i.e. we mean it is positioned between, and attracted to and by two atomic cores), and in what ways it is not (e.g. there is no formal social contract, and no notion of ownership). It takes a very careful teacher to suffix each use of the term with a proviso: ‘the bond is a shared pair of electrons, by which we mean that due to its position between the two atomic cores, it is attracted to and by both, and this the arrangement holds the two cores together in the structure’. It is easy to see why the ‘shorthand’ term is used.

Similarly, learners’ anthropomorphic comments can be accepted in class, but reflected back in more technical terms. Where the learner’s metaphor is not valid, it is also necessary to probe for a more scientific explanation: ‘...and what do you mean by saying the atom *needs* one more electron?’ 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. Indeed teachers might think of setting up ‘straw’ explanations that learners are invited to criticise and improve upon!

Concluding comments

The topic of teaching structural chemistry is a large and important one, and it has not been possible to do it full justice in this review. However, the present paper has considered some of the findings from research into learners' ideas and learning difficulties, and from the perspective of a general constructivist position, has attempted to suggest how we can teach this area of chemistry more effectively. Indications of 'pedagogic learning blocks' deriving from the way chemistry is currently learnt (and taught) have been highlighted - and suggestions have been made for how these impediments to learning might best be avoided. Only time will tell how easy some of these suggestions are to adopt, and how successful they might be, but at least research does provide us with some specific ideas for improving teaching practice.

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