ALTERNATIVE CONCEPTIONS OF CHEMICAL BONDING

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Chemical bonding is a topic which many secondary students find difficult to understand. The concepts in chemical bonding are abstract, so there is great potential for the formation of alternative conceptions as students try to derive meaning from what is said by the teacher or what is written in the textbooks. Thus teachers need to be aware of students' conceptions of chemical bonding in order to develop teaching strategies to enable their own students to construct ideas of chemical bonding which are compatible with the scientific concepts.

INTRODUCTION

Chemical bonding is an abstract topic, something far removed from the daily experiences of secondary school students (Grade 9 to 12). As students cannot see an atom, its structure and how it interacts with other atoms, it is difficult for them to understand the concepts involved in the topic of chemical bonding and there is great potential for the formation of alternative conceptions (Boo, 1994; Butts & Smith, 1987; Goh & Chia, 1989; Taber, 1994, 1997, 1998; 1999; Tan & Treagust, 1999). Teachers need to be aware of students' conceptions of various ideas associated with chemical bonding in order to develop teaching strategies to enable their own students to construct ideas of chemical bonding which are compatible with the scientific concepts. This paper presents some of the research concerning students' conceptions of certain aspects of chemical bonding, in particular, ionic and covalent bonding, and suggests alternative ways of teaching ionic and covalent bonding to secondary school students.

BOND AND BONDING

Boo (2000) in interviews with trainee-teachers enrolled in a post-graduate diploma in education course found that many did not seem to know the subtle differences in meanings between the terms 'bond' and 'bonding.' To them, the term 'ionic bonding' is synonymous with the term 'ionic bond' and the term 'covalent bonding' is synonymous with the term 'covalent bond'. She maintains that the subtle difference between the terms 'bonding' and 'bond' needs to be pointed out, since chemistry/science is after all a subject which stresses on precision and accuracy. The term 'bonding' refers to the process of bond formation whereas the term 'bond' refers to the attractive force which holds ions or atoms or molecules together. More specifically, the term 'covalent bonding' refers to 'the sharing of electrons between atoms of non-metallic elements, generally resulting in a noble gas electronic structure in the valence shell of the atoms involved.' In contrast, the term 'covalent bond' refers to 'the electrostatic force of attraction between the positively charged nuclei involved and the shared electrons.' Similarly, the term 'ionic bonding' refers to 'the transfer of electrons from the metallic atom to the non-metallic atom generally resulting in a noble gas electronic structure in the valence shells of the ions formed,' while the term 'ionic bond' refers to 'the electrostatic force of attraction between the oppositely charged ions formed as a result of the process of electron transfer.'

STUDENT CONCEPTIONS OF IONIC AND COVALENT BONDING

Previous research has identified a range of students' difficulties in understanding of ionic bonding. Butts and Smith (1987) found that most Grade 12 chemistry students associated sodium chloride with ionic bonds and the transfer of electrons from sodium to chloride, but many did not understand the three dimensional nature of ionic bonding in solid sodium chloride. A few students thought sodium chloride existed as molecules and these molecules were held together in the solid by covalent bonds. Others thought that sodium and chlorine atoms were bonded covalently but that ionic bonds between these molecules produced the crystal lattice. A three dimensional ball-and-stick model of sodium chloride also caused confusion among the students as many interpreted the six wires attached to each ball (ion) as each representing a bond of some sort. Boo (1998) interviewed 48 Grade 12 students and found that some students thought that the attraction between oppositely charged ions in an ionic compound results in the neutralisation of the charges, leading to the formation of a lattice consisting of neutral molecules.

Taber (1994), in interviews involving Grade 12 students, found that many students adopted a molecular framework for ionic bonding. He found that many students believed that:

1. The atomic electronic configuration determines the number of ionic bonds formed. For example, a sodium atom can only donate one electron, so it can form only one bond.
2. Bonds are only formed between atoms that donate/accept electrons. For example, in sodium chloride, the chloride is bonded to the specific sodium atom that donated an electron to it.
3. Ions interact with the counterions around them, but for those not ionically bonded these interactions are just forces. For example, in sodium chloride, a chloride ion is bonded to one sodium ion and attracted to a further five sodium ions, but just by forces and not bonds.

These findings were supported further by the data obtained in a later study (Taber, 1997) involving the administration of a thirty 'true or false' item test, 'Truth About Ionic Bonding Diagnostic Instrument' to Grade 10 to 12 chemistry students.

Tan and Treagust (1999), using a two-tier multiple choice diagnostic test, found that only 16.7% of the 119 Grade 10 students in a Singapore school appreciated that sodium chloride forms an ionic lattice. A high percentage of them (80.4%) believed that sodium chloride existed as molecules, and 46.1% thought that one sodium ion and one chloride ion formed an 'ion pair molecule' (Taber, 1994). Many students (22.5%) also indicated that when atoms of metals and non-metals combine, they form covalent bonds instead of ionic bonds. It was also found that 10% of the students thought that in ionic bonding, the number of electrons transferred depends only on the number of electrons that the atoms of the non-metal need to achieve a stable octet. Many of the findings from this study conducted in Singapore were similar to the findings by Taber (1994; 1997) and Butts and Smith (1987).

This seemed to imply that students from different parts of the world have similar alternative conceptions. It was very likely that these alternative conceptions arose due to similar methods of teaching and/or presentation of content in textbooks (Taber, 1997; Tan & Treagust, 1999) as students only encounter ideas about bonding during formal instruction.

In the case of covalent bonding, Taber (1998) highlighted that Grade 11 and 12 students would commonly identify and distinguish which electron in a covalent bond belonged to each of the bonded atoms. The students also considered the sharing of electrons as the 'force' holding the atoms in a molecule together instead of electrostatic attraction between the shared electrons and the nuclei involved. This finding is corroborated by Boo (2000) who found that some of her students held the misconception that a covalent bond is a pair of shared electrons. This misconception probably arises out of exposure to statements often found in textbooks such as the following:

> "A covalent bond is the pair of shared electrons in a covalent molecule,"
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> "One pair of shared electrons constitute a single covalent,"
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> "Two pairs of shared electrons constitute a double bond," and
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> "Three pairs of shared electrons constitute a triple bond."

Boo suggests that it needs to be clearly pointed out to students that a chemical bond is a force, an attractive force, and a pair of electrons by themselves cannot constitute an attractive force.

Boo (2000) also reports that some of her interviewees held the misconception that an ionic bond is electrostatic in nature but not the covalent bond. It seems that these students were unaware that all chemical bonds (including metallic bonds, Van der Waals bonds and hydrogen bonds) are electrostatic in nature. This misconception may have risen because in discussing ionic bonding, textbooks tend to mention that ions formed as a result of electron transfer (between metallic atom and the non-metallic atom) are held by an electrostatic attraction between these positively charged and negatively charged ions. At the same time, in discussing covalent bonding, these textbooks either make no mention of what constitutes the covalent bond or they merely mention that the pair of shared electrons is the covalent bond (see the discussion in the previous paragraphs). Because there is no mention that the covalent bond is electrostatic in nature, many students infer for themselves that the covalent bond is not electrostatic in nature.

Two similar studies involving the administration of a two-tier multiple choice diagnostic test on covalent bonding and structure to Grade 11 and 12 students in Australia (Peterson & Treagust, 1989; Peterson, Treagust & Garnett, 1989) and Singapore (Goh, Khoo, & Chia, 1993) found that the students incorrectly identified intermolecular forces as the forces within a molecule, while others thought that strong intermolecular forces exist in a continuous covalent (network) solid. This showed that students were confused between covalent bonds and intermolecular forces.

Tan and Treagust (1999) found that the younger Grade 10 students in their study also had many similar difficulties understanding intermolecular forces and covalent bonds as well as continuous covalent and molecular lattices. For instance, 21% of the students believed that the strength of intermolecular forces was determined by the strength of the covalent bonds present in molecules and that covalent bonds were broken when a substance changed state. This finding is supported by Boo (2000) who reports that some of her interviewees believed that covalent bonds are weaker than ionic bonds because they had the notion that covalent substances generally have lower melting points and boiling points compared to ionic substances. This appears to be linked to the inadequate textbook treatment on the concepts of bonding and properties of covalent and ionic substances. In many textbooks, the discussion on bonding often does not include the explanation that ionic bonding results in the formation of a giant ionic lattice structure whereas covalent bonding usually results in the formation of simple or discrete molecular structures. The notion that melting (or boiling) an ionic substance involves breaking the ionic bonds while melting (or boiling) a covalent substance with simple molecular structures does not involve breaking the covalent bond within the molecule but only involves breaking the relatively weaker bonds between molecules is often not pointed out in textbooks.

DIAGNOSING STUDENTS' UNDERSTANDING AND ALTERNATIVE CONCEPTIONS

The above-mentioned studies highlighted several alternative conceptions of covalent and ionic bonding. Teachers need to know their students' alternative conceptions in order to help them see the limitations of these conceptions and the advantages that the accepted science concepts have over them. Some of the studies made use of the rather labour-intensive and time-consuming interview-about-events technique to diagnose students' conceptions. Others used paper and pencil tests, such as the 'Truth about Ionic Bonding Diagnostic Instrument' designed by Taber (1997) and the 'Chemical Bonding Diagnostic Instrument' developed by Tan and Treagust (1999), which are convenient ways for secondary chemistry teachers to assess students' understanding of chemical bonding as these tests are readily administered and scored.

TEACHING AND LEARNING IONIC AND COVALENT BONDING

Taber (1994) believes that the way teachers presented their lessons on ionic bonding might have encouraged their students to develop the alternative conceptions on ionic bonding. Students might think that sodium chloride existed as discrete units of NaCl when they encounter textbooks and/or teachers illustrating ionic bonding by drawing the transfer of an electron from a sodium atom to a chlorine atom to form a positive sodium ion and a negative chloride ion and saying that the pair of sodium and chloride ions are attracted together by strong electrostatic forces. Ionic lattices typically would be introduced only a few lessons later when the students learned about the structure of solids, so they might not make the links between the formation of ionic bonds and ionic lattices. Boo (2000) while attributing the source of misconceptions to some textbooks' treatment of the topic 'ionic bonding' (which involves illustration with small numbers of atoms or molecules, and which contains no mention of the formation of the crystal lattice) also suggests that the misconception is further reinforced by the existence of multiple choice questions such as the following which can be found in assessment books or school tests:

How does a magnesium atom form a bond with an oxygen atom?

A. By sharing one pair of electrons, both electrons provided by the magnesium atom.

B. By sharing two pairs of electrons, each atom donating one pair of electrons.

C. By giving one pair of electrons to the oxygen atom.

D. By giving two pairs of electrons to the oxygen atom.

In addition, Boo (1994) suggests that the lack of understanding of what chemical formulae represented also contributed to the formation of alternative conceptions. For example, ionic sodium chloride is represented as NaCl which is very similar to covalent hydrogen chloride, HCl, so students might have the idea that one particle of sodium is bonded to one particle of chlorine just as one atom of hydrogen is bonded to one atom of chlorine. If covalent bonding, with its emphasis on valency and molecules, was taught before ionic bonding, it also could influence the adoption of molecular ideas in ionic bonding. Thus, Harrison and Treagust (2000) believe that teachers need to actively negotiate the analog-target mapping of all important metaphors, analogies and models that they use, as well as regularly check their students' visualisation of these models, metaphors or analogies. This means that teachers need to clarify the formation of ionic bonds within the context of a three dimensional ionic lattice - the use of ball-and-stick models (teachers need to be careful to explain the sticks do not represent a bond of sorts) or computer animations would be useful here.

Taber (1994) suggests the following strategies (p.102) to introduce ionic bonding to minimise alternative conceptions:

1. Focus on the electrostatic lattice forces, rather than ion formation—teachers should emphasise the omni-directional nature of the ionic bonds.

2. Clearly distinguish between ion formation (electron transfer) and ionic bonding.

3. Do not restrict diagrams to ones showing molecular-like entities (pairs of atoms and pairs of ions), but include ensembles of ions.
4. If the 'reason' for ion formation is the stability of noble gas electronic configurations, then make sure that this is not also considered sufficient reason for the subsequent formation of bonds between ions —for example, a sodium ion does not only form one bond with the chloride ion it transferred its electron to.
5. Discuss the differences (as well as similarities) between lattices held together by ionic, covalent and intermolecular forces.
6. Include an example of an ionic material formed via precipitation, eg barium sulphate (VI), to emphasise that ionic bonds can form even if no electron transfer is involved.
7. If the term valency is used at all, then discuss explicitly the meaning of electrovalency in terms of ionic charge formed, and compare this with covalency; and make it clear that the number of ionic bonds formed is not determined by electrovalency.

An alternative teaching strategy is to concentrate on the effective nuclear charge of the atoms involved in bonding. Metal atoms have lower effective nuclear charges compared to non-metal atoms and thus are more likely to lose their outermost electrons than share or gain electrons when combined with non-metal atoms. Thus metal atoms generally do not form covalent bonds or become negatively charged. When a metal atom such as a sodium atom loses its outermost electrons, it can be compared to the earth attracting all objects in all directions. The effects of the positive charge is omni-directional and positively charged metal ion would attract negatively-charged ions such as chloride ions in all directions. This image may help students overcome their conceptions that one sodium ion will only bond to one chloride ion. Taber (1999) also suggests emphasising the role of electrostatic forces from the time bonding is discussed at the particulate level.

The diagnostic instruments mentioned earlier can also function as effective teaching tools. Students can be instructed to answer the questions on their own in the first instance and then discuss in groups what the answers should be. This will allow students to air their ideas, listen to the ideas of others which may not coincide with theirs, and discuss to resolve any differences in answers. This exercise will make learning more effective as they construct, listen, discuss and reconstruct meanings of the concepts taught. Driver (1995) believes that discussion with peers encourages knowledge construction as it provides “a forum in which previously implicit ideas can be made explicit and available for reflection and checking [as well as] an opportunity for individuals to build on each other’s ideas to reach a solution” (p. 394).

Peterson et al. (1989) believe that science teachers may be placing greater emphasis on the completion of the syllabus or the acquisition of facts to the possible detriment of students’ understanding of the content. They also point out that chemistry teachers need to emphasise the distinction in meaning between the same words used in everyday English speech and in a chemistry context. The meanings of the chemical terms are obvious to the teachers so they may not define or teach the terms explicitly, so students may have difficulties in understanding and distinguishing the specialised use of the terms in chemistry (Fensham, 1994; Treagust, Duit, & Nieswandt, 1999). For example, “to share” in everyday usage means to possess jointly, whereas “a shared pair of electrons” mean that an electron pair exist in same space between the atoms in a molecule. Similarly, Boo (1998) found that some students thought that covalent bonds resulted from the sharing of one electron between two atoms because if each atom contributes an electron, then there is no ‘real’ sharing as in sharing an apple between two people. On a similar note, Goh et al. (1993) suggested that there was a need to use simple language to express scientific concepts in secondary education. For example, students seemed to be confused with the terms “inter” and “intra” in “intermolecular forces” and “intramolecular forces.” Simpler words such as “forces between molecules” and “forces within molecules” could be used as substitutes.

CONCLUSION

Chemical bonding is a difficult topic for secondary school students to understand as it involves many abstract concepts. Thus alternative conceptions will arise as students strive to understand the topic. Teachers need to be aware of these alternative conceptions. Scott, Asoko, Driver and Emberton (1994), and Wittrock (1994) believe that identifying and understanding student conceptions will improve science teaching and argue that the central focus of planning lessons should be the comparison of student conceptions and the accepted views of science. Suggestions to improve the teaching of chemical bonding have been cited from the literature as well as given by the authors, and it appears that explicit teaching is crucial. Thus, the prepositions and linkage of the concepts as well as their applications in different situations should be illustrated clearly to the students in order to minimise students' alternative conceptions.

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