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## From Common Sense Concepts to Scientifically Conditioned Concepts of Chemical Bonding: An Historical and Textbook Approach Designed to Address Learning and Teaching Issues at the Secondary School Level

Michael Croft

Avondale College, [mcroft@tpg.com.au](mailto:mcroft@tpg.com.au)

Kevin C. de Berg

Avondale College of Higher Education, [kevin.deberg@avondale.edu.au](mailto:kevin.deberg@avondale.edu.au)

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# From Common Sense Concepts to Scientifically Conditioned Concepts of Chemical Bonding: An Historical and Textbook Approach designed to address learning and teaching issues at the Secondary School level

**Abstract:** This paper selects six key alternative conceptions identified in the literature on student understandings of chemical bonding and illustrates how a historical analysis and a textbook analysis can inform these conceptions and lead to recommendations for improving the teaching and learning of chemical bonding at the secondary school level. The historical analysis and the textbook analysis focus on the concepts of *charge*, *octet*, *electron pair*, *ionic*, *covalent* and *metallic* bonding. Finally, a table of recommendations is made for teacher and student in the light of four fundamental questions and the six alternative conceptions to enhance the quality of the curriculum resources available and the level of student engagement.

**Key Words:** Charge; Cubic atom; Octet; Electron pair; Ionic bonding; Covalent bonding; Metallic bonding; Models

## Introduction

The chemical bond is fundamental to chemistry as it is what holds atoms together in molecules according to chemists. Up until about the middle of the 19<sup>th</sup> century philosophers and scientists thought of the chemical bond from a common sense point of view. Early Greek philosophers such as Democritus [ca 460BC-ca 370BC] spoke of ‘links’ between atoms (Barnes 2005) and in the 17<sup>th</sup> century Descartes [1596-1650] envisaged atoms to be held together by tiny hooks and barbs (Descartes 1984 Translation). In the 19<sup>th</sup> century John Dalton [1766-1844] imagined atoms hooked together to create molecules and from the combining masses was able to determine a list of atomic weights (Myers 2003). The discovery of the electron in 1897 by J.J Thomson [1856-1940] precipitated a move amongst chemists and physicists to examine what role the electron might play in the formation of a chemical bond. Twentieth century probing into the nature of the electron using quantum mechanical tools has had an important impact on our understanding, but, it would seem, has deepened the mystery of what constitutes a chemical bond. According to some chemists a chemical bond is “not a real measurable object and it cannot be clearly defined” (Gillespie & Robinson 2006, p. 97). Charles Coulson, former theoretical chemist from the University of Oxford, concluded: “Sometimes it seems to me that a bond between two atoms has become so real, so tangible, so friendly, that I can almost see it. Then I awake with a little shock, for a chemical bond is not a real thing. It does not exist. No one has ever seen one. No one ever can. It is a figment of our own imagination” (Coulson 1953, pp. 20-21). Statements of this nature remind us just how complex the scientifically conditioned concept of the chemical bond has become.

It is understandable, then, why secondary school students begin their study of chemistry thinking of a chemical bond as a material connection (Pabuccu & Geban 2006; Taber & Coll 2002; Talanquer 2006). Even teachers might be forgiven for initially using this common sense view. This is re-emphasised when we draw a line between two atoms to represent a bond. In fact, this common sense view was sufficient to allow chemists like Dalton and Berzelius [1779-1848] to determine atomic weights provided the composition of the substance was known. For this reason, much of 19<sup>th</sup> century chemistry involved the determination of the elemental composition of compounds. When the details of atomic structure began to emerge at the end of the 19<sup>th</sup> century and the beginning of the 20<sup>th</sup> century

chemists were anxious to understand how and why chemical reactions occur and this led to a deeper probe into the nature of the chemical bond. So, in secondary school chemistry, it is not long before students are introduced to a submicroscopic view of the nature of the chemical bond in ionic compounds like NaCl, in covalent compounds like H<sub>2</sub>O, and in metals like Cu as shown in Figure 1.

Such models as shown in Figure 1 involve what one might call scientifically conditioned concepts like atom, molecule, ion, cation, anion, proton, neutron, electron, attraction of opposite charges, repulsion of like charges, metallic, ionic, covalent, polar and non-polar. Many of these concepts do not reside in our everyday experience but take up residence within our schooling of the scientific way of knowing. Consequently, it has been found that students tend to rely on rote learning when it comes to the chemical bonding topic (Levy Nahum, Mamlok-Naaman, Hofstein & Krajcik 2007). This can be detected in the responses students often give when queried about chemical bond phenomena. Responses that

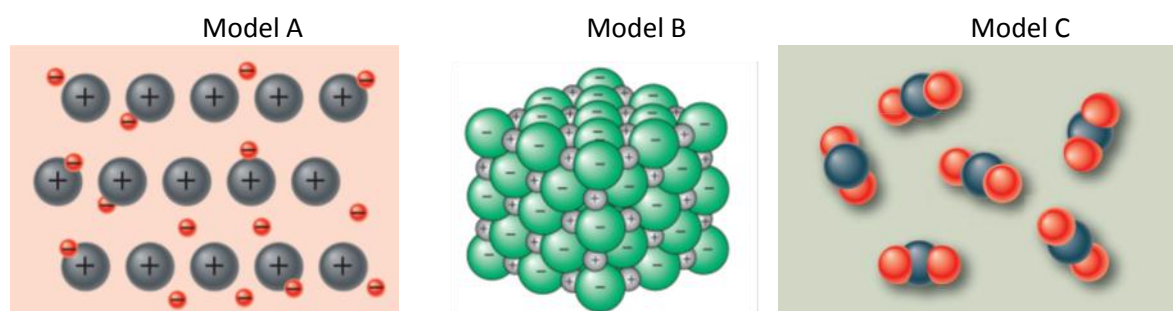


Figure 1. A typical school textbook representation of Metallic (Model A), Ionic (Model B), and Covalent (Model C) bonding (Lukins, Elvins, Lohmeyer, Ross, Sanders & Wilson 2006).

have drawn on a rote learning experience often use the scientifically conditioned concepts listed above in a random fashion so that the concept doesn't match the context. Some examples are: *Ionic bonds form molecules* (Tan & Treagust 1999); *Metals and non-metals form strong covalent bonds* (Unal, Calik, Ayas & Coll 2006); *Covalent bond formation involves the complete transfer of electrons* (Coll & Treagust 2002); *The negative charges in an ionic lattice are electrons*; *The positive charges in an ionic lattice are protons*; *The negative charges in a copper metallic lattice are copper ions* (Croft 2010). When a year 12 student was interviewed by Croft (2010) about the bonding in a hexane molecule, the student described the bonding as consisting of "a pair of electrons attracted to the positive nuclei of the carbon and hydrogen atoms", but then proceeded to call this type of bonding dipole-dipole bonding. When asked the same question about a water molecule, the student said that "a water molecule is held together by the attraction of a positive hydrogen atom to a negative oxygen atom" and identified covalent bonding as the attraction between non-metals. It is submitted that many such alternative conceptions identified in the literature arise from a rote learning experience rather than from what might be called a significant conceptual experience.

There are, however, some alternative conceptions which seem to arise from a significantly more permanent conception than those resulting from rote learning. In this paper it is planned to focus on six of these conceptions that relate to bonding within a molecule such as a covalent molecule or within a lattice such as a metallic or ionic lattice. The six conceptions are as follows.

1. Equal numbers of positive and negative charges cancel each other out to give a neutral molecule. Sometimes this conception is expressed as a positive charge using up a

negative charge. Some appropriate references are: Taber (1993); Coll & Treagust (2002); Croft (2010). This conception probably has a pedagogical source related to the teaching of formulae and chemical reactions. Taber (2001b) has called such conceptions ‘pedagogic learning impediments’. These impediments are also discussed by Levy Nahum, Mamlok-Naaman & Hofstein (2013). Such chemical processes as:  $H^+ + OH^- \rightarrow H_2O$ ; and  $H^+ + CO_3^{2-} \rightarrow HCO_3^-$ ; reinforce the view that a plus cancels out a minus. These chemical processes are distinctly different to that in a NaCl lattice for example. While the overall charge of the lattice is zero as indicated in the formula, NaCl, the lattice still consists of  $Na^+$  ions and  $Cl^-$  ions. This is an important distinction.

2. In an ionic lattice an ionic bond only exists where there has been an electron transfer between atoms to form oppositely charged ions. This view lacks an understanding of the omnidirectional character of the force field around a charge and is probably tied closely to the way the formation of an ionic bond is taught in terms of electron transfer. That is, an ionic bond is associated with the process of ion formation rather than with bond formation. This reflects the emphasis given in the school curriculum to the reaction between sodium metal and chlorine gas in the production of the sodium chloride lattice with sodium atoms losing one electron and chlorine atoms accepting it. However, the reaction between aqueous solutions of sodium hydroxide and hydrochloric acid with subsequent evaporation to give the solid salt does not involve electron transfer. An ionic bond is sometimes identified with the process of electron transfer rather than with the resulting attraction of opposite charges. Some appropriate references are: Taber (2002a); Kind (2004); Levy Nahum, Mamlok-Naaman, Hofstein & Taber (2010); Taber (2013).
3. Students tend not to regard attraction of opposite charges as a bond. They distinguish between ‘forces’ of attraction and ‘bond’ formation. While the models in Figure 1 are careful not to prescribe a bond as a material link, stick models like, Cl-Cl, for chlorine, for example, imply the existence of a ‘proper’ chemical bond whereas a lattice model for NaCl implies only forces of attraction rather than, as a student sees it, a chemical bond. Some appropriate references are: Taber & Coll (2002); Taber (2002a); Taber (2002b); Unal, Calik, Ayas & Coll (2006).
4. A sodium ion,  $Na^+$ , contains one more electron than the neutral sodium atom and a chloride ion,  $Cl^-$ , contains one less electron than the neutral chlorine atom. Since plus means more and minus means less it is understandable why a student might reason that a sodium cation with one plus attached must have one more electron than the neutral sodium atom and a chlorine anion with one minus attached must have one less electron than the neutral chlorine atom. However, the student has failed to relate the plus and minus signs to the proton and electron number present. Some appropriate references are: Croft (2010); Taber (1993).
5. Students have difficulty distinguishing between an element and its ions in properties and atomic structure. Since sodium is reactive then sodium chloride must also be reactive according to some students. Difficulty is experienced in determining the proton, neutron and electron number in the neutral atom and its corresponding ion. Some appropriate references are: Croft (2010); Taber (1993).

6. Students are preoccupied with the ‘octet rule’ when discussing chemical bonding even though there are many exceptions to it. This sometimes manifests itself as ‘shell-filling to achieve a noble gas electron configuration’. Thus sodium atoms lose one electron and chlorine atoms gain one electron so both atoms can achieve eight electrons in their outermost occupied shells. Carbon covalently bonds to four hydrogen atoms to achieve an octet of electrons in its valence shell. Some appropriate references are: Croft (2010); Taber (2001a); Taber (2002a); Taber & Coll (2002); Taber (2013); Tan et al (2008).

Five of the six conceptions listed here relate intimately to the concept of *charge* developed predominantly over a period of some three hundred years from the time of Benjamin Franklin [1706-1790] in the 18<sup>th</sup> century to the emergence of the field of quantum electrodynamics in the 20<sup>th</sup> century. However, the concept of *charge* is rarely considered to be important enough for discussion when considering new frameworks for teaching the topic of chemical bonding. This is in spite of the fact that: Models A and B in Figure 1 directly demand of the student an interaction with the concept of *charge*; scientific experts with an interest in chemistry education regard all chemical bonds as formed by the electrostatic attraction of opposite charges (Gillespie 1997); and it is acknowledged that chemical bonding is electrical in nature (Gillespie & Robinson 2006). The bottom-up approach to teaching the topic of chemical bonding (Levy Nahum, Mamlok-Naaman & Hofstein 2008) shows some promise in dealing with the so-called pedagogic learning impediments associated with chemical bonding but it is contended here that this approach will have limited success if restricted to a bottom rung of ‘atom’ and does not include an elaboration of the meaning of *charge*.

What do we mean when we say that something is charged, and in particular, why do we call one charge “positive” and another charge “negative”? And what is it that causes like charges to repel and unlike charges to attract? The fundamental significance of these kinds of questions was recently recognized in a popular science book written by Natalie Angier on what she called the “Beautiful Basics of Science” (Angier 2007). In introducing the readers to the basic components of the atom, the positively charged protons, the negatively charged electrons, and the neutral neutrons, she makes the following comment: “Well, that sounds breezy enough: a plus sign, a minus sign, and free with purchase. But what in the name of Mr Rogers’ last cardigan are we really talking about? What does it mean to say that a particle has “charge”, and how does this subatomic “charge” of the light brigade relate to more familiar, real-world displays of electric “charge”?” (Angier 2007, p. 16). What is interesting is that Angier (2007) identifies the asking of such basic but profound questions as consonant with the directions of current science education. This is evident in the following statement: “Moreover, in choosing to ask many little questions about a few big items, I was adopting a philosophy that lately has won fans among science educators-that the best way to teach science to non-scientists is to go for depth over breadth” (Angier 2007, pp. 16-17). Like so many science concepts, we have often taken for granted the idea that opposite charges attract and like charges repel without seriously challenging or searching for the foundation of this idea (de Berg 2011).

The sixth conception (related to the octet) relates very much to the work of Gilbert Lewis [1875-1946] on chemical bonding at the beginning of the twentieth century. Lewis was responsible for the notion of the electron pair or, as he called it, the law of two, in explaining covalent bonding in molecules like that depicted in Model C in Figure 1. It is contended that a great deal of the dissatisfaction experienced by the teaching community regarding the chemical bonding topic (Levy Nahum, Mamlok-Naaman & Hofstein 2013) rests in a lack of background support and consequently a lack of rapport with the fundamental ideas of *charge*,

the *law of eight*, and the *law of two*. Associated questions, along the lines proposed by Angier (2007), designed to enlighten these three fundamental ideas might include the following:

1. What do we mean when we say an electron is negatively charged and a proton is positively charged?
2. Why do opposite charges attract and like charges repel?
3. How did chemists arrive at the *law of eight* or *octet* and how significant does it remain?
4. How can an electron pair form a chemical bond when the electrons are both negatively charged?

If a popular scientist like Angier can ask such questions of the general public, how much more significant is it for informed teachers of science to be able to interact with such questions.

The purpose of this paper is to address these questions and consequently the six significant conceptions listed here and to provide some insights which might help reduce the tendency for rote learning. It is intended to do this by exploring the history behind the key concepts of *charge*, *electron pair*, and the *octet* electron structure. In a chapter entitled ‘A Role for History’, Thomas Kuhn was of the belief that, “History, if viewed as a repository for more than anecdote or chronology, could produce a decisive transformation in the image of science by which we are now possessed” (Kuhn 2012, p. 1). It follows that the content of science portrayed in our textbooks can also undergo a transformation when viewed through the lens of history. Tsaparlis (1997, p. 924) claims that, “We will understand the subject of atomic and molecular structure much better if its historic traces are followed”. What follows here is an account of the origin of the concepts of *charge*, *octet*, and *electron pair*, all of the details of which would not necessarily occupy space in a secondary school curriculum. The details are designed to alert the scholar, curriculum developer, and teacher to the opportunities for conceptual clarity but also to the framework of complexity that modern science presents. It is contended that if our teachers receive such support, the opportunity to create pedagogical learning impediments will be reduced and the teaching of the topic will become more enjoyable. To understand how the key ideas of *charge*, *octet*, and *electron pair* are presented in the secondary school curriculum, a textbook analysis is presented and some suggestions relating to improving the level of engagement of students with the topic of chemical bonding content for the secondary school curriculum made in the light of the historical discussion and the conceptions we know students bring to the learning-teaching situation which have been outlined here.

## The Concept of Charge<sup>1</sup>

A study of static electricity predated any serious study of atomic structure and bonding. Holton and Brush (2001, p. 352) trace the origin of the science of static electricity to ancient Greece where the property of rubbed amber attracting small objects was well-known. Priestley (1767, p. 2) identifies Thales of Miletus and Theophrastus in this respect and shows that the Greek name for “amber”, ηλεκτρον, can be translated “electron” from which the word “electricity” is obtained. In a series of simple but elegant experiments with rubbed objects and gold leaf, Du Fay proposed the existence of two kinds of electricity in 1733 (Assis 2010, pp. 106-110). One kind of electricity appeared to be resident in rubbed solid glassy substances and so was called *vitreous electricity*. The other kind of electricity was

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<sup>1</sup> A major portion of this section was first presented at the 11<sup>th</sup> meeting of the IHPST group in 2011 at Thessalonica in Greece by one of the authors.

resident in objects of a waxy, bituminous kind and was called by Du Fay, *resinous electricity*. When a rubbed glass rod was brought near a piece of gold leaf which had been previously touched by a rubbed glass rod, the gold leaf was repelled away. Exactly the same behaviour was observed with a rubbed amber rod. However, if a rubbed amber rod was brought near a gold leaf which had been previously touched by a rubbed glass rod, the gold leaf was attracted towards the amber rod. Du Fay interpreted these findings as suggesting the existence of two types of electricity as suggested above.

In a letter to Peter Collinson in 1747, Benjamin Franklin (1747) used the term *electrical fire* for what was gained by an object after being rubbed by certain materials. According to Franklin all animate and inanimate objects contain a certain amount of *electrical fire* by nature. In an experiment Franklin had two people stand on a sheet of wax with one person occasionally rubbing a glass rod and then touching the other person's knuckle with the glass rod. Franklin then observed that a person standing on the floor beyond the wax could receive a spark from either person on the wax when approaching the person with their knuckle. This spark was not as big, however, as the spark which occurred when the two people on the sheet of wax approached each other with their knuckles after electrification. To explain these observations Franklin suggested that rubbing did not create electric charge (electrical fire) but simply transferred it from the person rubbing the object to the object itself. When the electrical fire was transferred from the glass rod to the knuckle belonging to the other person on the sheet of wax, Franklin now regarded this person to possess an excess of electrical fire (positive fire), that is, more electrical fire than what was normally the case. The person who had transferred the electrical fire now possessed less electrical fire (negative fire) than what was normally the case. Of course, according to this fluid model of electrical charge, all objects could potentially possess positive or negative electrical fire or electrical charge. Eventually the terms positive and negative were used to classify the two types of electricity or electrical fire or electrical charge observed by Du Fay. By convention vitreous electricity was classified as positively charged and resinous electricity as negatively charged. The use of this convention enabled a more general classification system as not all objects that were positively charged were vitreous and not all objects that were negatively charged were resinous (Jensen 2005).

Now, one doesn't need to rub electrons to charge them because they possess charge by their very nature. We say they are negatively charged only because they are attracted to a positively charged electrode and 'positive' and 'negative' are nomenclature conventions based on the work of Franklin described above. Physicists remind us that, "It is best to realize right here that there is no easy picture for visualizing the concept 'charge'" (Holton & Brush 2001, p. 354). Bertrand Russell (1923) describes the enigma of the electron as follows: "When I say that an electron has a certain amount of electricity, I mean merely that it behaves in a certain way. Electricity is not like red paint, a substance which can be put on to the electron and taken off again; it is merely a convenient name for certain physical laws" (Holton & Brush 2001, p. 354). The concept of the electron is used in modern science to describe what happens when a glass rod, for example, is rubbed with a silk cloth. The rubbing is understood to have removed electrons from the glass rod and to have transferred them to the silk cloth so that the rod is positively charged and the cloth is negatively charged (Glynn 1999, p. 97). Angier describes how a hair comb can be used, after vigorously combing dry hair, to pick up pieces of paper by, "stripping off millions of electrons from the outermost shells of the atoms of your coiffure" (Angier 2007, pp. 99-100), thus becoming negatively charged and repelling electrons from the immediate edge of the paper. This leads to a negatively charged comb adjacent to the positive end of the paper providing for the attraction. But all this description still begs the question: "Why do like charges repel and unlike charges attract even without touching"?

This was a question of interest to Michael Faraday [1791-1867] who lived at a time when the dominant view of heat, electricity, and magnetism was that they were imponderable fluids. Another view, adopted by Humphry Davy [1778-1829], was that these phenomena were fundamental forces in nature. Michael Faraday learnt his trade, as it were, from Humphry Davy and was deeply influenced by his ideas including the pre-eminence of forces in nature. Faraday, however, was always reluctant to take a dogmatic stand on topics like the nature of matter and electricity, always deferring to the importance of discovering the laws of nature. Pearce Williams considers that, “Faraday’s suspension of judgement on the nature of electricity, magnetism, and matter itself well illustrates a distinguishing characteristic of his mind. He was always willing to refrain from leaping to a conclusion until the evidence appeared to justify it” (Pearce Williams 1965, p. 89). One of the topics that occupied his mind was how electrical and magnetic forces could operate over a distance without physical contact.

Why can objects exert a force on each other without touching? This was the ‘action at a distance’ problem addressed by Michael Faraday and a problem that Isaac Newton [1642-1727] identified with gravitation. In terms of electricity and magnetism Faraday introduced the *field* concept to explain interaction at a distance. The field was the means by which one object could exert a force on another object not in physical contact. A *field* as Giancoli describes, “is not a kind of matter. It is rather a concept-and a very useful one. Whether the electric field is ‘real’, and really exists, is a philosophical, even metaphysical, question. In physics it is a very useful idea, in fact a great invention of the human mind” (Giancoli 1989, p. 511). The *field* concept helps us describe how an electrically charged object can exert a force on another electrically charged object. If the emerging electric fields from both objects act in the same direction between the objects one gets attraction. If the emerging electric fields from both objects act in different directions between the objects one gets repulsion. The *field* concept was later mathematically elaborated by Clerk Maxwell [1831-1879] to produce his famous field equations. While the field concept was a useful idea to mediate the force between two charged objects in space, and how opposite charges could be shown to lead to attraction and like charges to repulsion, it did not explain why opposite charges lead fundamentally to attraction and like charges to repulsion. An answer to this question had to await the relatively recent development in particle physics of the field of Quantum Electrodynamics (QED).

Richard Feynman [1918-1988], regarded as the father of QED, shows how a photon acts as a carrier of electromagnetic force between two electrons using what has become known as a Feynman diagram shown below in Figure 2 (Sutton 1984, p. 65). A Feynman diagram is a qualitative one with time on the vertical axis and space on the horizontal axis. One of the negatively charged electrons emits a photon and recoils somewhat as a result. The second electron absorbs the photon and since the photon transfers energy and momentum the electron is diverted sideways giving the impression that like charges repel. This photon is different to that normally associated with electromagnetic radiation in that it is not measurable with our instruments and because of this, is called a virtual photon. The arrows in Figure 2 do not indicate direction of movement but rather that of the flow of negative electric charge (Veltman 2003, p. 246). Virtual photons can violate the law of conservation of energy by  $\Delta E$  for a very short time,  $\Delta t$ , provided that  $\Delta E \cdot \Delta t \approx h/4\pi$ , according to the Heisenberg Uncertainty Principle. Such quantum violations of energy conservation take place only in the short term and the system energy is conserved in the long run when the photon is reabsorbed by the other electron.

An attractive force between two particles such as an electron and a proton can be experienced if the particle on the left emits a virtual photon directed towards the left with the other particle absorbing the photon from the right-hand side. This is illustrated by the



Feynman diagram in Figure 3. This kind of exchange is possible because, as Ohanian (1987, p. 444) reminds us, a photon is not described by a localized orbit, but by a wave extended over a wide region.

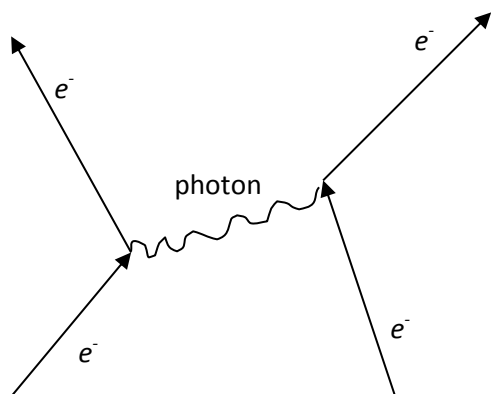


Figure 2. The interaction of two electrons pictured in a Feynman diagram showing the exchange of a virtual photon.

What is the status of QED in the scientific community? Scientists variously describe it as, “the most accurate theory in all of physics” (Ohanian 1987, p. 444); “probably the most accurate physical theory known” (Walecka 2008, p. 199); and “science’s most successful quantitative theory” (Davies & Gribbin 1991, p. 233). One of the reasons for such a glowing report of the theory is given by John Maddox as follows: “The success of QED has surprised even its practitioners. Many of the calculations of how electrically charged particles behave in interaction with photons and with the vacuum turn out to agree with experiment to within a few parts in 100 million” (Maddox 1998, p. 78). But how does such a successful theory feature in the literature read by the students of science? At this moment in our discussion it is not intended to restrict ourselves to the secondary school curriculum. That will come in the final section of the paper.

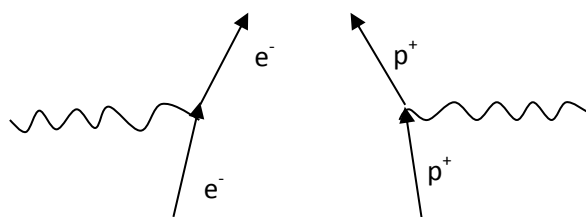


Figure 3. Feynman diagram for the attraction between an electron and a proton.

Electric charge is featured more in physics textbooks than in chemistry textbooks so what follows is drawn from physics textbooks at the tertiary level as it is at the tertiary level where QED is more likely to be discussed and where teachers may find the appropriate background material. Table 1 summarizes the information on electric charge portrayed in a selection of general physics textbooks used at the tertiary level. These textbooks were chosen because of their availability and because they are typical of textbooks used in first-year

physics where the concept of *charge* is usually considered. The historical information on *charge* gleaned in this paper is generally not featured in chemistry textbooks in spite of the fact that electric charge lies at the heart of atomic and molecular structure.

Table 1. Electric Charge features in general physics textbooks.

Text	Origin of +/-	Repulsion Explained by photon exchange	Attraction Explained by photon exchange	Analogies Used
Giancoli 1989	Yes	Yes	No	Yes
Ohanian 1987	No	Yes	Yes	No
Serway, Moses & Moyer 2005	No	Yes	No	No
Serway & Jewett 2004	Yes	Yes	No	No
Halliday, Resnick & Krane 1992	Yes	No	No	No
Fleming 1978	Yes	Yes	Yes	No
Walecka 2008	No	No	No	No
Thornton, & Rex 2000	No	No	No	No
Cassidy, Holton & Rutherford 2002	Yes	No	No	No

It should be kept in mind, when examining the data in Table 1, that textbooks often set about their task with different purposes. For example, the text by Cassidy, Holton, and Rutherford (2002) is designed for college undergraduates who are not planning a career in science or engineering and the text by Serway, Moses, and Moyer (2005) is based on the premise that students have already studied an introductory calculus-based physics course. This most likely explains why the former text does not include material on virtual photons and particle physics and the latter does not include fundamental material on electric charge. While five of the texts explain electron-electron repulsion through virtual photon exchange, only two explain attraction through virtual photon exchange. The reason for this is probably due to the fact that repulsion by photon exchange appeals more to common sense than does attraction. One has to appeal to the non-localised nature of the photon to illustrate attraction.

Giancoli (1989, p. 1023) is the only textbook in Table 1 that uses analogies to explain the difference between repulsion and attraction. The analogy involves two people exchanging pillows while on skates. Throwing the pillows to each other represents repulsion since reaction to the throwing action causes the two individuals to separate; and grabbing the pillows from each other represents attraction since the grabbing action causes the two individuals to move towards each other. This is consonant with QED theory since it is the way virtual photons are exchanged that determines whether attraction or repulsion occurs. In a non-physics textbook source, Morris (1997) likens repulsion to two skaters throwing a ball to each other and attraction is likened to two skaters throwing a boomerang to each other when back-to-back. In spite of the fact that these analogies are very helpful in picturing how repulsion and attraction might work, there is still an underlying mystery as to how a particle knows how to exchange a virtual photon. Is there something about having opposite charges near each other that tells one of the particles to exchange photons by the “grabbing” mechanism as opposed to the “throwing” mechanism? It is not certain whether we have an answer to this question.

It is interesting to observe that Ohanian does not discuss the origin of the concepts of positive and negative charge in the context of the history of frictional electricity, although he

does discuss virtual photons and particle physics. It is worthwhile quoting his reason as follows:

We will begin our study of electricity with the fundamental electric force between charged particles rather than following the historical route, because the origin of frictional electricity remains rather mysterious. Even now, physicists have no precise understanding of the detailed mechanism that generates electric charge on rubbed bodies or why some rubbed bodies acquire positive charge and some negative charge (Ohanian 1987, p. 572).

Holton and Brush describe the current understanding of charging by friction in terms of removal or addition of electrons but add; “we are not prepared to say just how or why some materials, like glass, joyfully allow their electrons to be carried off, whereas others, like amber and resin, tend to grab extra electrons very copiously” (Holton & Brush 2001, p. 354). So there remains a mystery at almost every level of the discussion right through from charging by rubbing to exchanging virtual photons.

So, how might one approach the question posed in the introduction: Why do opposite charges attract and like charges repel? Should one simply regard the question as an exercise in classification, that is to say, we know there are two types of charge and these happen, for historical reasons, to be classified as ‘positive’ and ‘negative’, although they could equally have been called ‘alpha’ and ‘beta’. As to why they interact as repulsion or attraction, the best we can currently do is to refer to quantum field theory and describe the interaction as a particular mode of virtual photon transfer. As to why some materials become positive through friction as opposed to negative through friction, there does not appear to be an adequate answer. Is it the case that conceptual reduction both clarifies and mystifies a scientific question? What is certain is that what began as an innocuous question has resulted in a fascinating journey in epistemology. There is a real sense in which unanswered questions become a catalyst for excitement in learning as long as progress is made in clarifying ideas along the way. It would appear that the question of electric charges is one such example. The second and third of our key ideas, that of the octet and electron pair, will now be discussed.

### **The Concepts of the Octet and the Electron Pair**

In 1904 J.J. Thomson [1856-1940] made the first attempt to explain the chemical bond in terms of electrons (Hudson 1992). He proposed that corpuscles (electrons) would be transferred from one atom to another as compounds form. Thomson further explained that as a result of the transfer of electrons, the electronegative atom would become negatively charged, the electropositive atom would become positively charged, and the oppositely charged atoms would be attracted to each other forming a compound (Shaik 2007). This theory of the ionic bond would become the dominant theory for the next two decades. It was used to explain bonding in every type of substance. It was widely accepted that all bonds were formed by transferring an electron. Even non-polar molecules were considered to have formed this way. For example, the hydrogen molecule was considered to be ionic, even though its lack of polar properties caused explanatory problems for chemists.

This problem became an intense topic of research in the early 20<sup>th</sup> century and in 1916 G.N. Lewis [1875-1946] presented the first satisfactory model of the covalent bond in terms of a shared electron pair (Niaz 2009) to explain the bonding in non-polar diatomic molecules like hydrogen (H<sub>2</sub>). Such a model was controversial in that many chemists questioned the stability of two negatively charged entities co-existing in a region between neighbouring nuclei while not experiencing significant repulsion. Lewis made an important observation that the vast majority of stable molecules contain an even number of electrons, which led him

to suggest that electrons are usually present in pairs but the notion of the electron pair proved controversial. It is quite instructive to contemplate the status of the theories of chemical bonding at the turn of the 20<sup>th</sup> century as detailed by Kohler as follows:

When it was first proposed, Lewis' theory was completely out of tune with established belief. For nearly 20 years it had been almost universally believed that all bonds were formed by the complete transfer of one electron from one atom to another. The paradigm was the ionic bond of sodium chloride,  $\text{Na}^+\text{Cl}^-$ , and even the bonds in compounds such as methane or hydrogen were believed to be polar, despite their lack of polar properties. From the standpoint of the polar theory the idea that two negative electrons could attract each other or that two atoms could share electrons was absurd (Kohler 1971, p. 344).

Lewis' thinking began with a model of the atom called the 'cubical atom' (Lewis 1916). The cubical atom consisted of an outer shell of electrons which were arranged symmetrically at the eight corners of a cube as shown in Figure 4. At the centre of the atom, the cube, was an essential kernel of positive charge. In terms of the model in Figure 4 neon would have a full shell with an electron at each corner of the cube.

Lewis provides a clue as to how he arrived at the cubical atom model as early as 1902:

A number of years ago, to account for the striking fact which has become known as Abegg's law of valence and countervalence, and according to which the total difference between the maximum negative and positive valences or polar numbers of an element is frequently eight and is in no case more than eight, I designed what may be called the theory of the cubical atom (Lewis 1916, p. 767).

Lewis thinks of valence and countervalence very much like we now consider oxidation number; with H nearly always being +1 and O being -2. Applying Abegg's law then to  $\text{H}_2\text{S}$  and  $\text{H}_2\text{SO}_4$ , sulphur has a negative valency of -2 and a positive valency of +6 respectively and the sum of the absolute value of the negative valency and the positive valency is 8. Chlorine's negative valency in HCl is -1 and its maximum positive valency in  $\text{ClO}_4^-$  is +7 with the sum of the absolute values of the valencies being again 8. A cubical atom, then, with eight equivalent spaces available for electrons in the valence shell or cube could accommodate, in the case of chlorine for example, the loss of seven electrons or the gain of one electron.

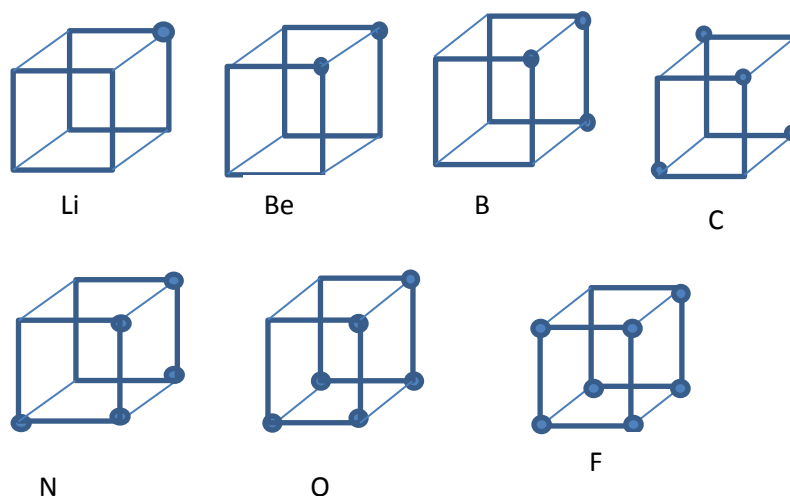


Figure 4. Lewis' cubical atom from lithium to fluorine showing the valence electrons positioned at the corners of a cube.

Another clue given by Lewis into the origin of the model of the cubical atom comes from his statement in 1923 as follows:

In the year 1902 (while I was attempting to explain to an elementary class in chemistry some of the ideas involved in the periodic law) becoming interested in the new theory of the electron (Thomson's discovery of the electron in 1897), and combining this idea with those which are implied in the periodic classification, I formed an idea of the inner structure of the atom (model of the cubic atom) which, although it contained crudities, I have ever since regarded as representing essentially the arrangement of the electrons in the atom (Lewis 1923, pp. 29-30).

What ideas in the periodic law may have assisted Lewis in his progress to a cubic atom? Periodic Tables in 1902 typically listed compositional formulae for the oxides and hydrogen compounds of the elements under each group (Quam & Quam 1934). For example, in Brauner's table of 1902, the oxides for element R from Groups I to VII are listed as:  $R_2O$ ,  $RO$ ,  $R_2O_3$ ,  $RO_2$ ,  $R_2O_5$ ,  $RO_3$ , and  $R_2O_7$ . The hydrogen compounds of R are listed from Groups IV to VII as:  $RH_4$ ,  $RH_3$ ,  $RH_2$ , and  $RH$ . The oxides give the positive valency of R and the hydrogen compounds give the negative valency of R. These formulae confirm that the sum of the absolute valencies is 8 as previously illustrated. For example, comparing the valencies of R in  $R_2O_5$  (+5) and R in  $RH_3$  (-3) gives a sum of 8 taking the absolute value of the negative valency. It is interesting to note that Mendeleev also used these oxides and hydrogen compounds in his short table of 1872. The presence of eight elements in the second row of the periodic table was also suggestive of a cubic atom if lithium possessed one electron in its outer shell and each subsequent element added one additional electron so that the last element in the row, neon, had eight electrons in its outer shell. At one stage Lewis thought there must be six unknown elements between hydrogen and helium but this eventually had to be abandoned when it was confirmed that helium had only two electrons.

In Lewis' cubical model chemical bonds formed when cubes joined together. According to Lewis, a single bond is formed when two cubic atoms share an edge and a double bond is formed when two cubic atoms share a face as shown in Figure 5. The single bond was constituted of two electrons (an electron pair) belonging to both cubes and the double bond was constituted of four electrons (or two electron pairs) belonging to both cubes. That is, the bonding electrons could be thought of as existing as pairs of electrons. Lewis emphasized that the single most important mechanism of chemical bonding was electron pairing (Shaik 2007) but he had no clear idea why electrons should be found in pairs in molecules, apart from the fact that the number of valence electrons in molecules was nearly always an even number. The cubical atom was simply a convenient model for representing the outer electrons of the eight second row elements of the Periodic Table and for representing chemical bonding using electron pairs. Lewis was keenly aware of the objections being raised to the electron pair model based on the repulsion properties of like charges. At one stage he proposed that Coulomb's law of repulsion might not operate at the submicroscopic level and on another occasion he was impressed with the suggestion made by Alfred Parson that "the force responsible for chemical bonding was not electrical but magnetic" (Kohler 1971, p. 364). In any case the cubical model seemed to be the best available at the time.

The triple bond could not be represented, however, using the model of the cubic atom as shown in Figures 4 and 5. Lewis suggested that a better representation of the electron structure, at least for small atoms, might be that derived from the known tetrahedral characteristics of the carbon atom. This could be obtained by moving the eight electrons of the cubic atom closer to each other along the edges of the cube as shown in Figure 6. Lewis understood that this could be argued to increase the repulsion experienced by these electrons.

However, Lewis surmised that the “electrons have a tendency to be drawn together, perhaps by magnetic force if the magneton theory is correct, or perhaps by other forces which become appreciable at small distances...” (Lewis 1916, p. 780). The fact that the two electrons in the electron pair had opposite spins according to the Pauli Principle with a consequent zero magnetic moment which enabled them to coexist in a region between nuclei was to await the development of quantum mechanics some years later. The tetrahedral atom now made it possible to represent single, double, and triple bonds. “Two tetrahedra, attached by one, two or three corners of each, represent respectively the single, the double and the triple bond” (Lewis 1916, p. 780).

Resident within the cubic atom model is the observation that having eight electrons in the outer shell provided the most stable conditions for the atom. Lewis called this observation the *rule of eight* and Langmuir [1881-1957] renamed it the *octet rule* (Coffey 2008; Gillespie & Robinson 2006). Lewis was aware of the exceptions to his *rule of eight* and was uncomfortable with the way the rule and its application became more universal than he ever intended. For example, he was aware that “in the row of the periodic table comprising hydrogen and helium we have in place of the rule of eight the rule of two” (Lewis 1916, p. 774). For a large number of students the octet rule is still seen as the most important bonding principle that they know. The transfer and sharing of electrons are both considered to be driven by the need for the atom to achieve an octet of electrons in the outer shell (Taber 2002a; Taber & Coll 2002). When students commence their chemistry education the octet rule can be of value in identifying stable species as Robinson (1998) suggests, but its perceived role in controlling bond formation and chemical reactions is rather misplaced. As Gillespie and Robinson (2006, p. 91) say, “Indeed the octet rule applies strictly only to the period 2 elements, C, N, O and F. For the atoms of all other elements the rule may or may not be obeyed”.

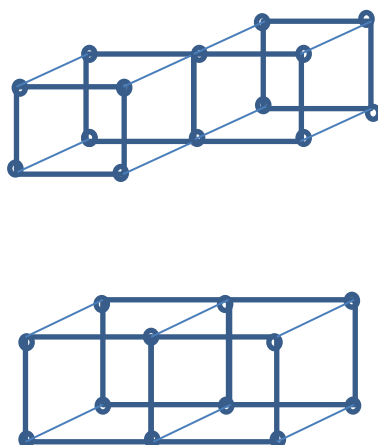


Figure 5. Lewis’ model for a single (upper diagram) and double bond (lower diagram).

According to Jensen (2009) early work on metallic bonding was done by Drude [1863-1906] and Lorentz [1853-1928] who proposed that metals must contain weakly bound electrons in order to conduct electricity. In 1913 Lewis argued for the existence of three types of chemical bonding - polar (ionic), non-polar (covalent), and metallic. The electrons were regarded as occupying fixed positions within the atom in ionic bonding; as moving freely from atom to atom within the molecule in covalent bonding; and as free to move outside any molecular structure in metallic bonding. Lewis (1913) suggested that all molecules would fall into at least one of these three categories. Stark [1874-1957] (1915)

made the first attempt to visualize the three bonding situations as shown in Figure 7. Fernelius and Robey (1935) illustrated the three bonding types by placing them at the corners of a triangle with linkages between them representing intermediate bonding possibilities as shown in Figure 8.

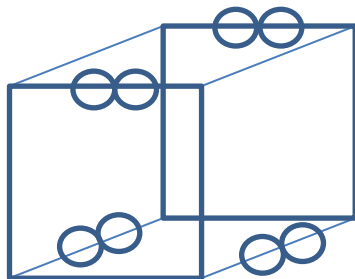


Figure 6. Lewis' tetrahedral atom/molecule derived from the cubic atom/molecule.

Science educators have argued for the importance of these intermediate bonding positions and view the “different traditional categories of chemical bonding as extreme cases of various continuum scales.....One of the key goals of the proposed framework is to stress that a continuum scale exists between extreme cases of qualitatively different bonding scenarios” (Levy Nahum, Mamlok-Naaman, & Hofstein 2008, pp. 1682-1683). To this end these authors are critical of many textbooks that introduce the metallic bond as:

metal ions floating in a sea of electrons. This analogy is problematic because it presents the metallic bond as a bonding entity that is entirely different from the covalent one, whereas a more modern description views both types of bonding as involving *electron sharing*. The difference is again explained in terms of a continuum scale, this time involving the degree of electron delocalization” (Levy Nahum et al 2008, p. 1861).

That is, as one moves from an ionic bond to a covalent bond to a metallic bond the extent of electron delocalization increases.

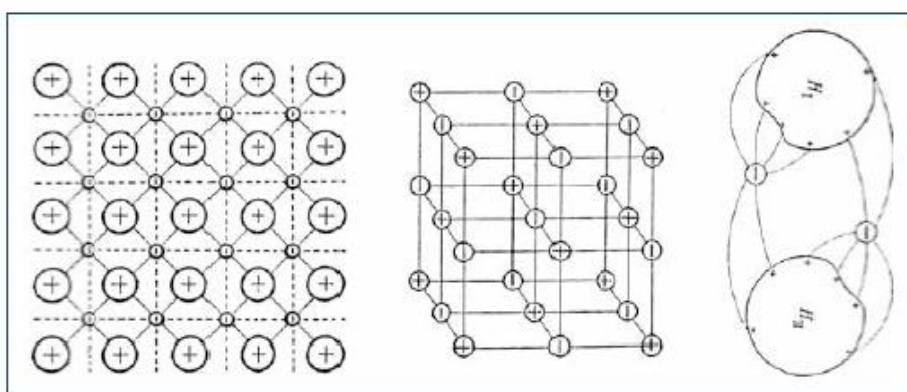


Figure 7. The first attempt to visualise all three types of bonding situations – metallic, ionic and covalent (from left to right) by J. Stark (1915).

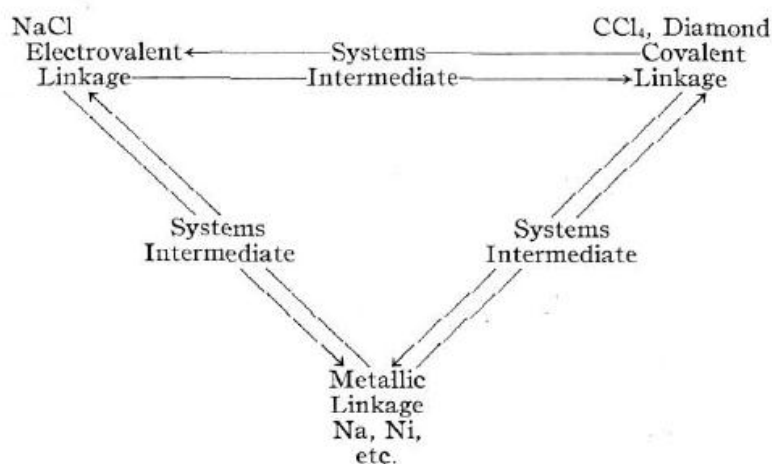


FIGURE 17.—TYPES OF VALENCE LINKAGES

Figure 8. The Bonding-Type triangle that explicitly outlines the three types of primary bonds and the intermediate types of bonds between the extremes suggested by Fernelius & Robey (1935)(Fig. 17).

### Chemical Bonding and the Secondary School Curriculum

If one was to summarise the recommendations for teaching and learning practice in secondary science classrooms arising from science education research over the past thirty years, it would be the need to focus on the active role of the learner in acquiring a knowledge foundation upon which new concepts could be built in meaningful ways. Taber (2006, p. 173) has called this the “cornerstone” for teaching and learning practice. Two questions arise from this focus: What constitutes a knowledge foundation for chemical bonding? And, What is involved in the learning process? In some ways these two questions are linked as far as the orientation of this paper is concerned. The paper has drawn attention to some significant alternative conceptions that have arisen from the learning process inside and outside the classroom and has chosen to provide an historical background to key ideas that appear in the alternative conceptions. The key ideas selected were *charge*, *octet*, and *electron pair*. So one way of providing a knowledge foundation is through an historical analysis. The question, then, is: What part of the historical analysis is relevant and should be applied to the secondary school curriculum? In addition: What part of the analysis might be pertinent for teachers and what part pertinent for students?

Other ways of producing a knowledge foundation for chemical bonding have been through the development of a concept map (Tan & Treagust 1999) and through locating the ‘elements of knowledge’ for concepts placed in a matrix for chemical bonding (Yayon, Mamlok-Naaman & Fortus 2012). Yayon, Mamlok-Naaman and Fortus (2012) also select *charge* as an important concept and describe the elements of knowledge pertinent to *charge* as shown in Table 2. *Charge* does not feature strongly on the concept map produced by Tan and Treagust (1999) as the emphasis given is to connecting bonding models to physical and chemical properties. It will be interesting to compare the content in Table 2 with that portrayed in the historical analysis given in this paper and the way *charge* is portrayed in textbooks at the secondary level. However, before comparisons are made and previous questions in this section are answered it will be useful to detail a little more of what we know about the learning process with a particular focus on chemistry education.



Table 2. The Elements of Knowledge for ‘Charged particles’ (Yayon et al 2012).

<p><b>Charged particles</b> Protons are positive Electrons are negative The charge of the nucleus equals the number of protons In a neutral atom the number of electrons equals the number of protons If the number of electrons differs from the number of protons, the atom is charged If the number of electrons differs from the number of protons, the atom is called an ion Positive ions have fewer electrons than protons; negative ions have more electrons than protons Momentary partial charges (<math>\delta -</math> and <math>\delta +</math>) in the electron cloud occur because electrons do not have fixed positions</p>
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According to Johnstone’s (2006) information processing model, new information is attached to some point in the student’s long term memory. During the process of learning new information, the student will need to recall information from the long term memory in order to make sense of the new information. Then the new information is stored alongside existing knowledge and understanding. If the learner thinks the new information is valuable, but cannot link it to existing information, the information enters the long term memory as rote learning. Such information is hard to recall. If the learner attaches the new information to some other knowledge in a faulty way, an alternative conception is created. This faulty attachment is very hard to undo because the alternative conception makes sense to the student. Some alternative conceptions in the area of chemical bonding have already been noted in the historical introduction and a selection of others can be found in the Footnote <sup>2</sup>. If the learner attaches new information to some other knowledge in a coherent and meaningful way, deep progressive learning occurs and begins to occupy a useful and easily locatable place in the long term memory. The elements of knowledge in Table 2 do not originate in common sense everyday experience but are scientifically conditioned. This means that initially rote learning may be a significant part of the learning experience. However, it should not be too long before a transition is made from rote learning to deep meaningful learning. The question is how might this transition be made.

The historical account of the concept of *charge* and the account of the alternative conceptions in the introduction relating to *charge* offer some hope in this regard. As far as the historical account goes it is clear that *positive charge* and *negative charge* are arbitrary classifications signifying two different behaviours of rubbed materials towards a gold leaf electroscope. The two different behaviours could have been classified as ‘*a*’ and ‘*b*’, or ‘*x*’ and ‘*y*’, or ‘ $\alpha$ ’ and ‘ $\beta$ ’. The terms ‘positive’ and ‘negative’ were conventions used by Benjamin Franklin. The elements of knowledge for *charge* listed in Table 2 commence with the startling revelation that ‘protons are positive’ and ‘electrons are negative’ with no indication that these are arbitrary classifications telling us that protons and electrons behave differently in an environment called an electric field. *Charge* hasn’t been rubbed or painted on to the proton or electron. The charge is a classification of behaviour in an electric field. A small narrative recapping Du Fay’s experiments with a vitreous and amber rod could well reduce the impact of rote learning at this point. It is interesting at this stage to review how school science textbooks treat the concept of *charge*.

Twelve textbooks covering the junior and senior secondary science curriculum for some Australian states and some overseas jurisdictions have been reviewed for the concept of *charge* as shown in Table 3. It is interesting to note that ‘charge as an arbitrary label for

<sup>2</sup> Coll and Taylor 2001; Coll and Treagust 2001, 2003a, 2003b; Peterson, Treagust and Garnett 1989; Taber 1994, 2003; Tan and Treagust 1999

behaviour' hardly features at all across junior and senior textbooks. Some senior physics textbooks do use the 'convention' terminology and speak about *charge* being a basic or fundamental property of matter but no indication of what this means is given. Most junior science and senior physics textbooks define *charge* in electron terms after having accepted the proposition that, as Table 2 indicates, protons are positive and electrons are negative. Thus when one's hair is vigorously combed electrons are stripped from the hair making the hair positive and the comb negative. There is no indication as to why the electrons are considered negative in the first place.

The significance of the last four columns in Table 3 should be evident on review of some of the alternative conceptions relating to *charge* outlined in the introduction. As indicated some students think that when the proton and electron numbers are equal then the charges actually cancel out making the atom neutral. The same reasoning is used when observing an ionic compound where the number of positive charges equals the number of negative charges. This reasoning is reinforced when we write Na for the sodium atom and NaCl for the salt sodium chloride. It is interesting to learn that some of the textbooks in Table 3 reinforce the 'cancellation of charges' point of view in such statements as: "the negative and positive charges neutralise each other" (Lofts et al 2004, p. 78) and with respect to the neutral lithium atom, "The three electrons surrounding the nucleus cancel out the positive charge of the three protons resulting in a neutral atom" (Spence et al 2004, p. 223). Two of the textbooks in Table 3 *imply* that charge balance does not mean charge cancellation in the formula NaCl but there is no indication at all in the other textbooks. With respect to the formula NaCl, one text says, "but it is important to remember that the ionic bond still exists" (Thickett 2000, p. 61), and the other text says, "Note that the ionic charges are not indicated in the formula" (Joesten & Hogg 2011, p. 75). These two statements however do not actually specify that the charges are still present. Georgiadou and Tsaparlis (2000) suggest including the charges in ionic formulae such as,  $(\text{Na}^+\text{Cl}^-)$ ;  $[(\text{NH}_4)^+(\text{NO}_3)^-]$ , to remind students that the charges are still present. However, this becomes rather difficult in cases where the compound has appreciable ionic and covalent character.

The last two columns specify that the *charge* on an ion depends on the *numbers* of protons and electrons present not on the *arithmetic* significance of (+) and (-) with respect to valence electrons. Thus  $\text{Na}^+$  does not indicate that the sodium ion has one more electron than the sodium atom, Na, because of the plus sign; and  $\text{Cl}^-$  does not indicate that the chloride ion has one fewer electrons than the chlorine atom, Cl, because of the negative sign. Most textbooks in Table 3 do present the *number* significance of the charges on an ion but despite this, the alternative conception based on the *arithmetic* significance of (+) and (-) persists for some students as indicated in the introduction to this paper. We suggest that this is an element of knowledge that should be specifically addressed in science/chemistry curricula as also the arbitrary nature of the assignment of 'positive' and 'negative' to *charge* derived from the historical account.

According to Johnstone's information processing model for learning, information is processed in the *working memory*. Since chemical information is commonly presented across the three levels of representation; *macroscopic*, *submicroscopic*, and *symbolic* (Johnstone 1991, 2006); the *working memory* can easily become overloaded. Students are often simultaneously introduced to new substances (the macro level); are required to describe these new substances in terms of sub-microscopic particles (the sub-micro level); and then to represent new substances using special symbols and chemical formulae (the symbolic or representational level). Ultimately, however, it is desirable for the student to gain confidence in thinking of matter across these three levels. The focus of the textbook bonding representations in Figure 1 is at the sub-micro level but students are expected to eventually function at all three levels.

**Table 3.** A Review of 12 secondary textbooks for the concept of *charge* (√ = present; x = absent).

Textbook	Charge as arbitrary label for behaviour	Charge defined in electron terms	Neutral atom as equality of proton and electron number	Charge balance as not equivalent to charge cancellation	Cation definition in terms of p>e	Anion definition in terms of e>p
Science Tracks 2000 (Junior) (Thickett et al)	x	√	√	x	√	√
Science World 2001 (Junior) (Stannard et al)	x	x	√	x	√	√
Science Focus 2009 (Junior) (Rickard et al)	x	√	√	x	√	√
Science Alive 2005 (Junior) (Nardelli)	x	√	√	x	√	√
Chemistry Contexts 1 2001 (Senior) (Irwin et al)	Convention- no behaviour mentioned	x	√	x	√	√
Chemistry Pathways 2000 (Senior) (Thickett)	x	x	√	√ but only implied	√	√
CHEM in your world 2011 (Senior) (Joesten et al)	x	x	√	√ but only implied	√	√
Chemistry- A contextual approach 2004 (Senior) (Spence et al)	x	x	√	x	√	√
Jacaranda Physics 2004 (Senior) (Lofts et al)	x but basic property of matter	x	x	x	√	√
Physics contexts 1 2002 (Senior) (Heffernan et al)	x but (+) and (-) attributed to Franklin	√	√	x	√	√
New Century Senior Physics 2004 (Senior) (Walding et al)	x but fundamental property of matter	√	√	x	√	√
Physics A-level 2002 (Senior) (Dobson et al)	x	√	√	x	x	x

Modelling chemical phenomena using the three levels of representation has become the principal way of thinking in chemistry. Justi and Gilbert (2002) explain that learning chemistry involves coming to understand the major models such as the textbook chemical bonding ones in Figure 1, their scope, limitations and roles. According to Gilbert:

A model can, at a given level, be expressed in ‘external representations’ - those versions physically available to others - and in ‘internal representations’ - those versions available mentally to an individual person. The making of meaning for any such representation is ‘visualization’ ..... Visualization is thus, in the first instance, concerned with the formation of an internal representation from an external representation such that the nature and temporal/spatial relationships between the entities of which it is composed are retained (Gilbert 2008, pp. 3-4).

Thus modelling and visualization are intimately involved in the learning process. The extent to which students engage with the models in Figure 1, process the elements of knowledge, and internalize the representations and their components determines the outcome of the learning process. Part of the engagement is related to the chemical language used to describe what the models in Figure 1 represent. In Table 4 is recorded the incidence of the language used to describe ionic, metallic, and covalent bonding in the same textbooks as outlined in Table 3. In addition the use of the octet model is also recorded.

First of all one notes from Table 4 that senior physics textbooks, while dealing with the concept of *charge* as shown in Table 3, do not, perhaps understandably, deal with chemical bonding. On the other hand, senior chemistry textbooks deal with chemical bonding as seen in Table 4 but do not deal with the concept of *charge* as shown in Table 3, even though it is submitted that an understanding of charge is central to an understanding of chemical bonding. This is a situation that needs to be rectified, perhaps by including a section on the concept of charge in senior chemistry textbooks even though some junior science textbooks visit the concept of charge at an elementary level.

Features A, B, and C in Table 4 deal with the idea of the *octet* and the idea of a stable electron configuration represented by the noble gases. The ideas of the *octet* and a stable electron configuration are used by students in explaining chemical behaviour often without realising that there are exceptions to the rule of eight which Lewis was well aware of when he proposed his cubic atom model. This fixation on the *octet* was detailed as an important alternative conception in the introduction to this paper. The information in Table 4 suggests that the term *octet* appears in senior chemistry textbooks and not junior science textbooks although the idea of a stable noble gas electron configuration of 2 (like helium) and 8 (like the other noble gases) is present in the junior textbooks. It is important for students to realise that scientific models have strengths and weaknesses and focusing on both can reduce the tendency to regard the octet notion as a fixed law or rule.

Lewis' cubic atom model provides a good setting for addressing the strengths and weaknesses of the octet idea. The fact that there were eight elements in the second row of the periodic table suggested that each element had one more valence electron than the previous element until one finally reached neon for which there were no known chemical reactions, suggesting a valence shell of eight electrons was stable. The cubic atom model was successful in demonstrating this provided no new elements with atomic weight between that of lithium and neon were discovered. Remember that, initially, Lewis thought there should also be eight elements in the first row of the periodic table between hydrogen and helium until it was discovered that helium had only two electrons. Lewis was aware that his cubic atom model was limited in this respect. The cubic atom model was able to explain successfully how single and double bonds could form leading to eight electrons around each cubic atom but was powerless to explain the triple bond. When it comes to the third row of the periodic table there are examples where the octet idea works as in  $\text{SiCl}_4$  and other examples where more than eight electrons in the valence shell applies as in  $\text{PCl}_5$ . Students need to be aware of such strengths and weaknesses in the models used in science. None of the textbooks examined made use of the cubic atom model in association with the octet idea. It may be too much to expect the cubic atom model to be included in a student's textbook given what is often described as an already overcrowded curriculum. But it could be a useful addition to a teacher's handbook as the strengths and weaknesses of the model are easily accessed with some knowledge of the periodic table. Gillespie and Robinson (2006) have shown that for a free atom or ion with a valence shell octet of electrons the most probable relative arrangement of eight electrons is with alternating spins at the corners of a cube. This was the arrangement suggested by Lewis but without the notion of spin. So the cubic arrangement of electrons still has some relevance in modern chemistry. According to

**Table 4.** The use of chemical bonding models across the twelve secondary textbooks ( $\checkmark$  = present; x = absent).

Textbook	A	B	C	D	E	F	G	H	I
Science Tracks 2000 (Junior) (Thickett et al)	x	x	$\checkmark$	$\checkmark$	$\checkmark$	x	$\checkmark$	x	$\checkmark$
Science World 2001 (Junior) (Stannard et al)	x	x	x	$\checkmark$	x	x	$\checkmark$	x	$\checkmark$
Science Focus 2009 (Junior) (Rickard et al)	x	x	$\checkmark$	$\checkmark$	x	$\checkmark$	$\checkmark$	x	x
Science Alive 2005 (Junior) (Nardelli)	x	x	$\checkmark$	x	x	$\checkmark$	$\checkmark$	x	x
Chemistry Contexts 1 2001 (Senior) (Irwin et al)	x	$\checkmark$	$\checkmark$	$\checkmark$	x	$\checkmark$	$\checkmark$	x	$\checkmark$ only in inter. forces
Chemistry Pathways 2000 (Senior) (Thickett)	x	$\checkmark$	$\checkmark$	$\checkmark$	$\checkmark$	$\checkmark$	$\checkmark$	x	$\checkmark$
CHEM in your world 2011 (Senior) (Joesten et al)	$\checkmark$	$\checkmark$	$\checkmark$	$\checkmark$	$\checkmark$	x	$\checkmark$	x	x
Chemistry- A contextual approach 2004 (Senior) (Spence et al)	x	$\checkmark$	$\checkmark$	$\checkmark$	x	$\checkmark$	$\checkmark$	x	x
Jacaranda Physics 2004 (Senior) (Lofts et al)	x	x	x	x	x	x	x	x	x
Physics contexts 1 2002 (Senior) (Heffernan et al)	x	x	x	x	x	x	x	x	x
New Century Senior Physics 2004 (Senior) (Walding et al)	x	x	x	x	x	x	x	x	x
Physics A-level 2002 (Senior) (Dobson et al)	x	x	x	x	x	x	x	x	x

**Note:** **A**= Octet as law or rule

**B**= Octet as model

**C**= Stability of noble gas electron configuration (2) or (8)

**D**= Ionic Bond as attractive force between (+) and (-) ions

**E**= Ionic Bond in electron transfer terms

**F** = Metallic Bond as attractive force between (+) ions and delocalised electrons

**G** = Covalent Bond in electron sharing terms

**H** = Covalent Bond as simultaneous attraction of electron pair to neighbouring nuclei

**I** = Bond and force equated

Gillespie and Robinson (2006) the electrons in the cube begin to pair off once other nuclei come into the vicinity to bond. It is interesting to note that only 3 of 27 General Chemistry textbooks used at college or university level used Lewis' cubic atom model when introducing the covalent bond (Niaz 2001).

Features D and E in Table 4 refer to the ionic bond. Most textbooks treating the ionic bond describe it as the attractive force between positive and negative ions in the lattice but some often confuse the issue by including a statement to the effect that it is the transfer of electrons that constitutes the bond. This may explain why this alternative conception has been noted in the literature. For example, one text suggests that, “the transfer of electrons... produces an ionic bond” (Joesten & Hogg 2011, p. 72). The omnidirectional character of the electric field around an ion suggests to some students that an attractive force is not really a chemical bond. This has been featured in alternative conception three in the introduction. The results for feature I in Table 4 suggests that some attention could be given in textbooks and other curriculum materials to this issue.

Feature F describing the metallic bond as an attractive force between positive ions and delocalised electrons is generally well represented in chemistry textbooks. However, there is a major issue for features G and H relating to the covalent bond. All junior science and senior chemistry textbooks describe the covalent bond as the sharing of electrons between two atoms or nuclei. None of the textbooks describe the bond as the simultaneous attraction of a shared electron pair to neighbouring nuclei. That is, the essential electrical nature of a chemical bond arising from the presence of charges has been lost. One can see how textbooks can assist in the development of pedagogical learning impediments in cases like this. Croft (2010) observed that only 20% of a secondary school cohort of 172 students were able to identify the covalent bond as the simultaneous attraction of a shared electron pair to neighbouring nuclei. It is so important that chemical models and the language used to describe the bonding represented are closely linked in the teaching and learning of chemistry.

## Conclusion

A reasonably comprehensive historical account of chemical bonding as it pertains to the secondary school curriculum has been given. Because of the emphasis on the secondary curriculum, valence bond theory and molecular orbital theory have not been discussed as they pertain more appropriately to the tertiary level. An introductory account of the role of virtual photons in explaining the repulsive and attractive effects of charge has been given particularly for the benefit of teachers. One would not expect the topic of virtual photons to appear in a secondary curriculum but it has been presented here to alert teachers to the complexity of the topic of chemical bonding and to recognize the importance of helping students understand that the frontiers of scientific knowledge are never closed. It is in this sense that scientific models are never complete but have demonstrated strengths and weaknesses. This paper has been written for the benefit of teachers and students and it will be useful to summarise the significance of our discussion of *charge*, *octet*, *electron pair*, and the *six alternative conceptions*, for teachers and students at the secondary level. The summary is given in Table 5. No attempt is made to specify the grade level appropriate to each category. Each section includes a question or activity for engaging students beyond the level of rote learning. Where possible, an attempt is made to show how the historical account, the textbook analysis, and the alternative conceptions or key questions are interrelated.

**Table 5.** A summary of the significance of the key ideas of *charge, octet, electron pair* and *six alternative conceptions* for teacher and student. (*SE* = Student engagement)

Key Idea/Question	Significance for Teacher	Significance for Student
What do we mean when we say an electron is negatively charged and a proton is positively charged?	Same as expressed for the student but in addition: At the time of Benjamin Franklin charge was thought of as an electrical fluid which all objects possessed to a greater or less extent. An object that experienced an increase in electrical fluid was said to be positively charged and an object that experienced a loss of electrical fluid was said to be negatively charged. Now positive and negative charges are labels decided by convention. A different convention could have assigned the electron a positive charge and the proton a negative charge. This conventional aspect is not well represented in textbooks as shown in Table 3.	An object is said to be charged if it sparks or spreads apart the two gold leaves of an electroscope. By convention a rubbed glass rod is said to be positively charged. Any object that further spreads apart the gold leaves of an electroscope is also said to be positively charged. Any object that collapses the gold leaves of an electroscope previously treated with a rubbed glass rod is said to be negatively charged. Positive and negative charge are labels given to the proton and electron respectively by convention because of their different electrical behaviour. Charge is not something coated on the electron and proton. Sometimes charge is expressed as a quantity of electrons: 1 mole of electrons is equivalent to 96500 coulombs: 1 coulomb is equivalent to $6.24 \times 10^{18}$ electrons: 1 electron is equivalent to $1.6 \times 10^{-19}$ coulombs. Experimentally it is found that the proton and electron have the same charge magnitude but of opposite sign. <b>SE:</b> How would you classify an object that did not affect the gold leaves of an electroscope?
Why do like charges repel and unlike charges attract?	This is not a theoretical proposition but can be established by experiment. The use of the field model can assist the student in illustrating the phenomenon. Currently, the phenomenon is best explained by Quantum Electrodynamics (QED). QED proposes that charges can communicate by the exchange of virtual photons and the mechanism of exchange determines whether the charges repel or attract. The analogy of ice skaters exchanging pillows suggests that tugging on a single pillow will lead to attraction and tossing a pillow to each other will lead to repulsion.	This is established by experiment. Advanced level physics is required to access the best answer that science can offer at this time. However, a knowledge of the experimental fact is sufficient for the student to be able to understand many electrical phenomena met at the secondary level. The use of the field concept can help in illustrating repulsion and attraction. <b>SE1:</b> Account for the fact that a comb passed through your hair a number of times can pull a stream of water flowing out of a burette towards itself. <b>SE2:</b> Use the field concept and appropriate drawings to illustrate a model of how a positive charge might attract a negative charge.
How did chemists arrive at the 'law of eight or octet' and how significant is the law today?	Same as expressed for the student but in addition: Lewis envisaged the valence electrons as positioned at the	If one considers the Periodic Table and places the 'd' and 'f' blocks on their own, the remainder of the Periodic Table consists of rows of

	<p>corners of a cube. A single bond involves a cube from each atom sharing an edge. A double bond involves a cube from each atom sharing a face. A triple bond was not able to be represented by this model. Modern chemistry envisages the eight valence electrons of a free atom or ion as positioned at the corners of a cube with neighbouring electrons of opposite spin. One can regard the cube as consisting of two intersecting tetrahedra with each tetrahedron containing electrons of the same spin but of opposite spin to those of the other tetrahedron. Be aware that senior chemistry textbooks, according to Table 4, tend to focus on the octet as a model rather than a law.</p>	<p>eight elements apart from the first row. Lewis even considered that there may have been six as yet undiscovered elements between H and He. The row of eight elements suggested a maximum of eight valence electrons in the outer shell of the atom. Also the maximum and minimum valence of the elements added up to eight without regard to sign suggesting the importance of eight spaces available for electron occupation. For example, Cl in HCl and Cl<sub>2</sub>O<sub>7</sub> has a valency of -1 and +7 respectively. The octet rule only strictly applies to C, N, O, and F. It may or may not apply to other elements. Lewis' 'law of eight' was named and promoted as the 'octet rule' by Langmuir. <b>SE:</b> Discuss how you can determine the maximum and minimum valencies for the element, nitrogen.</p>
<p>How can an electron pair form a chemical bond when they are both negatively charged and should repel each other?</p>	<p>Same as expressed for the student but in addition: The eight valence electrons positioned at the corners of a cube are drawn together as pairs of opposite spin in a tetrahedron when an additional nucleus approaches for chemical bonding. Be aware that textbooks, according to Table 4, tend to equate the covalent bond with electron pair sharing rather than simultaneous attraction of an electron pair to neighbouring nuclei.</p>	<p>The large number of molecules with an even number of valence electrons suggested that electron pairing seemed to be the only way to understand the origin of the formulae for molecules. The formula for methane, CH<sub>4</sub>, suggested that the eight valence electrons, four from C and four from the four H, must be paired off. The Pauli principle dictates that the electrons in each pair must have opposite spins but the electrons must be simultaneously attracted to neighbouring nuclei to remain paired. The combination of attraction to neighbouring nuclei and opposite spins counteracts the repulsion that two negative charges should experience. <b>SE:</b> If only a single electron, rather than an electron pair, was required to form a covalent bond between two elements, determine the formula for a species formed between carbon and hydrogen. Be prepared to argue your case.</p>
<p>Alternative conception 1: Students think positive and negative charges have cancelled each other out in a neutral ionic compound.</p>	<p>Same as expressed for the student but in addition: Electrical behaviour due to charge can be quantified using Coulomb's Law: <math>F = q_1 \cdot q_2 / (4\pi\epsilon r^2)</math> where <math>r</math> is the distance between the two charges (<math>q_1</math> and <math>q_2</math>) and <math>\epsilon</math> is the</p>	<p>It has been shown that the conventional allocation of the signs (+) and (-) to a charge represents different electrical behaviour. In the case of chemical species the (+) sign indicates a proton or an excess of protons and the (-) sign indicates</p>



	<p>dielectric constant of the surrounding medium. <math>F</math> is positive when the two charges have the same sign and <math>F</math> is negative if the two charges have opposite signs. The fact that charge balance is not equivalent to charge cancellation is not well represented in textbooks as shown in Table 3.</p>	<p>an electron or an excess of electrons. The fact is that a proton repels a test positive charge with the same magnitude of force as an electron attracts the test positive charge. So, when a test positive charge experiences a force (<math>+f</math>) in the presence of a proton it will experience a force (<math>-f</math>) in the presence of an electron. This means that a test positive charge experiences an overall zero force in the presence of an equal number of (+) and (-) ions. So an ionic lattice made up of equal numbers of <math>\text{Na}^+</math> and <math>\text{Cl}^-</math> ions is said to be neutral, not because the charges cancel out but because a test positive charge experiences zero force. Likewise, an atom made up of 11 protons and 11 electrons is neutral not because the 11 protons cancel the charge of 11 electrons, but because a test positive charge would experience zero net force in the presence of equal numbers of protons and electrons.</p> <p><b>SE:</b> A student argues that since the compound, NaF, is neutral, then the sodium and fluorine must not be charged. Present an argument that either agrees or disagrees with the student's position.</p>
<p>Alternative conception 2: An ionic bond only exists where electron transfer has taken place.</p>	<p>Same as expressed for the student but in addition: Try to restrict electron transfer processes to a discussion of redox chemistry and not chemical bonding. Keep in mind that historical remnants often impinge on the content of chemistry and some of these remnants can lead to 'pedagogical learning impediments'. According to Table 4, some chemistry textbooks speak of the ionic bond in electron transfer terms.</p>	<p>At one stage in the 19<sup>th</sup> century all chemical bonds were thought to involve electron transfer and this notion has continued to influence chemistry curricula as seen in the textbook analysis in Table 4. Think of all chemical bonds as involving the electrostatic attraction of opposite charges. Electron transfer has to do with oxidation and reduction processes and not with bond formation. Ion formation must be distinguished from bond formation.</p> <p><b>SE:</b> A student claims that only two ionic bonds exist in the compound, <math>\text{MgCl}_2</math>. Present an argument that either agrees or disagrees with the student's claim.</p>
<p>Alternative conception 3: Students do not equate 'forces of attraction' with bond formation.</p>	<p>Same as expressed for the student but in addition: Equating forces of attraction with bond formation could be a fact more strongly represented across science textbooks given the paucity shown in Table 4.</p>	<p>Typical covalent bonds are directional in character and can be represented as a 'line' in a formula such as Cl-Cl for <math>\text{Cl}_2</math>. Typical ionic and metallic bonds are omnidirectional in character and hence cannot be indicated by a line. All electrostatic forces of attraction represent a chemical bond of one</p>

		<p>type or another whether they can be represented by a line or not.</p> <p><b>SE:</b> Locate directional and omnidirectional bonding in <math>\text{NH}_4\text{NO}_3</math>. Use diagrams and formulae where appropriate.</p>
<p>Alternative conception 4: Students understand <math>\text{Na}^+</math> to have one more electron than Na and <math>\text{Cl}^-</math> to have one less electron than Cl because of the arithmetic significance of the (+) and (-) signs.</p>	<p>Same as expressed for the student.</p>	<p>Charges in chemical species always relate to relative proton (+) and electron (-) numbers so that <math>\text{Na}^+</math> has one more proton than electron compared to Na which has equal numbers of protons and electrons. Since proton numbers in <math>\text{Na}^+</math> and Na must be the same, this means that <math>\text{Na}^+</math> has one less electron than Na. <math>\text{Cl}^-</math> has one more electron than proton compared to Cl which has equal numbers of protons and electrons. Since proton numbers in <math>\text{Cl}^-</math> and Cl must be the same, this means that <math>\text{Cl}^-</math> has one more electron than Cl.</p> <p><b>SE:</b> Student A claims that <math>\text{K}^+</math> has 19 protons, 18 electrons, and 20 neutrons. Student B claims that <math>\text{K}^+</math> has 20 protons, 19 electrons, and 20 neutrons. Student C claims that <math>\text{K}^+</math> has 19 protons, 20 electrons, and 20 neutrons. Determine which student is correct by arguing your case.</p>
<p>Alternative conception 5: Students have difficulty distinguishing between an element and its ions in properties and atomic structure.</p>	<p>Same as expressed for the student.</p>	<p>A different electron structure is often associated with different chemical properties. The structure of the periodic table also illustrates this point. Since Na and <math>\text{Na}^+</math> have different electron numbers, their chemical properties should be different. Placing a piece of Na in water gives a much different reaction to placing salt in water. For a discussion of proton and electron numbers for elements and their ions see the comments above for Alternative conception 4.</p> <p><b>SE:</b> When a strip of magnesium is heated in air, it burns with a bright flame. A student heats a sample of magnesium sulphate expecting it to burn brightly also. Present an argument which either agrees or disagrees with the student's expectation.</p>
<p>Alternative conception 6: Students are preoccupied with the octet rule when discussing bonding and chemical properties.</p>	<p>Same as expressed for the student but in addition: The notion of the octet should not be used as a principle guiding the direction of a chemical reaction or the nature of its products. Focus on as many exceptions to the rule as</p>	<p>The law of eight or octet rule only strictly applies to C, N, O and F chemistry apart from a small number of free radical compounds. Think of the octet as a useful observation rather than a rule.</p> <p><b>SE:</b> There are two relatively</p>

	<p>adherents. Energy minimization (free energy) is a more useful guide than the octet. Since senior chemistry textbooks tend to focus on the octet as a model rather than a law (Table 4), a preoccupation with the octet rule or law may be a function of an adopted teaching style rather than a dependency on the textbook.</p>	<p>common chlorides of phosphorous. Write down their formulae and determine if there is an octet of electrons around the phosphorous atom in each case.</p>
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One of the major targets of this paper has been to draw attention to the centrality of the concept of *charge* to an understanding of the scientifically conditioned idea of the chemical bond. While *charge* is mentioned in the literature, the concept is very quickly glossed over in favour of a discussion of the types of bonding in substances. It has been demonstrated that the concept of *charge* looms high in the mind of the popular scientific press and underlies at least five student alternative conceptions. However, even the concept of the *octet*, related to the sixth alternative conception, draws upon the maximum and minimum valencies of the elements which are related to the theoretical charges an element might be assigned. At least some alternative conceptions can be referred to as pedagogic learning impediments which arise from the teaching/learning environment. Since teaching style and textbooks contribute to the teaching/learning environment, both textbooks and the needs of the teacher have been considered as well as the students. Historical and philosophical considerations of chemical bonding are particularly suited for resourcing the teacher and reducing the opportunities for the creation of pedagogic learning impediments.

Of special significance in this paper is the transition one must make from the generalised concept of *charge* and the conventional allocation of plus and minus discussed in the historical account to the concept of *charge* in chemical species in terms of protons and electrons detailed in Table 5. This transition is required to address the six alternative conceptions. Electrical behaviour, spoken of in the historical account, can be quantified by measuring the force experienced by a test positive charge placed near the charge in question. What is of great significance is that the magnitude of force experienced by the test positive charge when placed near a proton is exactly the same as when the test positive charge is placed at the same distance from an electron, but of opposite sign. It is this fact, and not the cancellation of charges, that leads to neutral species when the proton number and electron number are the same.

## References

- Angier, N. (2007). *The Canon-A Whirligig Tour of the Beautiful Basics of Science*. New York: Houghton Mifflin Co.
- Assis, A.K.T. (2010). *The Experimental and Historical Foundations of Electricity*. Montreal: Apeiron.
- Barnes, J. (2005). *The Presocratic Philosophers* (E- book edition). New York: Routledge.
- Cassidy, D., Holton, G., & Rutherford, J. (2002). *Understanding Physics*. New York: Springer.
- Coffey, P. (2008). *Cathedrals of Science: The personalities and rivalries that made modern chemistry*. USA: Oxford University Press.
- Coll, R.K. & Taylor, N. (2001). Alternative Conceptions of Chemical Bonding held by Upper Secondary and Tertiary Students. *Research in Science & Technological Education*, 19(2), 171-191.

- Coll, R.K. & Treagust, D.F. (2001). Learners' Mental Models of Chemical Bonding. *Research in Science Education*, 31, 357-382.
- Coll, R.K. & Treagust, D.F. (2002). Exploring Tertiary Students' Understanding of Covalent Bonding. *Research in Science and Technological Education*, 20(2), 241-267.
- Coll, R.K. & Treagust, D.F. (2003a). Investigation of Secondary School, Undergraduate, and Graduate Learners' Mental Models of Ionic Bonding. *Journal of Research in Science Teaching*, 40, 464-486.
- Coll, R.K. & Treagust, D.F. (2003b). Learners' mental models of metallic bonding: A cross-age study. *Science Education*, 87(5), 685-707.
- Coulson, C.A. (1953). *The Spirit of Applied Mathematics*. Oxford: Clarendon Press.
- Croft, M. (2010). *Towards the Development of a Pedagogical History for a Key Chemical Idea: Chemical Bonding*. Unpublished MEd (Hons) thesis, Avondale College of Higher Education.
- Davies, P. & Gribbin, J. (1991). *The Matter Myth*. London: Penguin.
- de Berg, K.C. (2011). But Miss Franklin: Why do opposite charges attract and like charges repel? In F. Seraglou, V. Kouountzos & A. Siatras (Eds.). *Science and Culture, Proceedings of the 11<sup>th</sup> IHPST meeting* (pp. 187-191). Thessaloniki, Greece: International History and Philosophy of Science and Science Teaching Group.
- Descartes, R. (1984). *Principles of Philosophy* (Trans. by V.R. Miller & R.P. Miller). The Netherlands: Kluwer.
- Dobson, K., Grace, D. & Lovett, D. (2002). *Physics*. London: Harper Collins.
- Fernelius, W.C. & Robey, R.F. (1935). The Nature of the Metallic State. *Journal of Chemical Education*, 12(2), 53-69.
- Fleming, P.J. (1978). *Physics*. Philippines: Addison-Wesley.
- Franklin, B. (1747). Benjamin Franklin Papers, Letter to Peter Collinson, May 25, Philadelphia.
- Georgiadou, A. & Tsaparlis, G. (2000). Chemistry Teaching in Lower Secondary School with methods based on: A) Psychological Theories; B) The Macro, Representational, and Submicro levels of Chemistry. *Chemistry Education Research and Practice*, 1(2), 217-226.
- Giancoli, D.C. (1989). *Physics for Scientists and Engineers with Modern Physics* (2<sup>nd</sup> ed.). New Jersey: Prentice Hall.
- Gilbert, J. (2008). Visualization: An Emergent Field of Practice and Enquiry In Science education. In J.K. Gilbert, M. Reiner & M. Nakhleh (Eds.). *Visualization: Theory and Practice in Science Education*. The Netherlands: Springer.
- Gillespie, R.J. (1997). The Great Ideas of Chemistry. *Journal of Chemical Education*, 74, 862-864.
- Gillespie, R.J. & Robinson, E.A. (2006). Gilbert N. Lewis and the Chemical Bond: The Electron Pair and the Octet Rule from 1916 to the Present Day. *Journal of Computational Chemistry*, 28(1), 87-97.
- Glynn, I. (1999). *The Anatomy of Thought-The Origin and Machinery of the Mind*. Oxford: Oxford University Press.
- Halliday, D., Resnick, R. & Krane, K.S. (1992). *Physics* (4<sup>th</sup> ed.). New York: John Wiley & Sons.
- Heffernan, D., Parker, A., Pinniger, G. & Harding, J. (2002). *Physics Contexts 1*. Australia: Pearson Education.
- Holton, G. & Brush, S.G. (2001). *Physics, The Human Adventure*. New Brunswick: Rutgers University Press.
- Hudson, J. (1992). *The History of Chemistry*. Hong Kong: The Macmillan Press.

- Irwin, D., Farrelly, R. & Garnett, P. (2001). *Chemistry Contexts 1*. South Melbourne: Pearson Education.
- Jensen, W.B. (2005). The Origins of Positive and Negative in Electricity. *Journal of Chemical Education*, 82(7), 988.
- Jensen, W.B. (2009). The Origin of the Metallic Bond. *Journal of Chemical Education*, 86(3), 278-279.
- Joesten, M. & Hogg, J. (2011). *CHEM In Your World*. Belmont, CA: Brooks/Cole Cengage Learning.
- Johnstone, A.H. (1991). Why is science difficult to learn? Things are seldom what they seem. *Journal of Computer Assisted Learning*, 7, 75-83.
- Johnstone, A.H. (2006). Chemical Education Research in Glasgow in perspective. *Chemistry Education Research and Practice*, 7(2), 49-63.
- Justi, R. & Gilbert, J. (2002). Models and Modelling in Chemical Education. In J.K. Gilbert, O. de Jong, R. Justi, D.F. Treagust & J.H.V. Driel (Eds.), *Chemical Education: Towards Research-Based Practice* (pp.47-68). The Netherlands: Kluwer.
- Kind, V. (2004). *Beyond Appearances: Students' misconceptions about basic chemical ideas*. Paper retrieved from:  
<http://www.rsc.org/education/teachers/learnnet/pdf/learnnet/rsc/miscon.pdf>
- Kohler, R.E. (1971). The origin of Lewis's theory of the shared pair bond. *Historical Studies in the Physical Sciences*, 3, 343-376.
- Kuhn, T. (2012). *The Structure of Scientific Revolutions* (50<sup>th</sup> anniversary edition). Chicago: University of Chicago Press.
- Levy Nahum, T., Mamlok-Naaman, R. & Hofstein, A. (2008). A New "Bottom-Up" Framework for Teaching Chemical Bonding. *Journal of Chemical Education*, 85(12), 1680-1685.
- Levy Nahum, T., Mamlok-Naaman, R. & Hofstein, A. (2013). Teaching and Learning of the Chemical Bonding Concept: Problems and Some Pedagogical Issues and Recommendations. In G.Tsaparlis & H.Sevian (Eds.), *Concepts of Matter in Science Education* (pp.373-390). Dordrecht: Springer.
- Levy Nahum, T., Mamlok-Naaman, R., Hofstein, A. & Krajcik, J. (2007). Developing a New Teaching Approach for the Chemical Bonding Concept Aligned with current Scientific and Pedagogical Knowledge. *Science Education*, 91(4), 579-603.
- Levy Nahum, T., Mamlok-Naaman, R., Hofstein, A. & Taber, K.S. (2010). Teaching and Learning the concept of Chemical Bonding. *Studies in Science Education*, 46(2), 179-207.
- Lewis, G.N. (1913). Valence and Tautomerism. *Journal of the American Chemical Society*, 35, 1448-1455.
- Lewis, G.N. (1916). The Atom and the Molecule. *Journal of the American Chemical Society*, 38, 762-785.
- Lewis, G.N. (1923). *Valence and the Structure of Atoms and Molecules*. New York: Chemical Catalog.
- Lofts, G., O'Keefe, D., Pentland, P., Phillips, R., Bass, G., Nardelli, D., Robertson, P., Tacou, J. & Pearce, J. (2004). *Jacaranda Physics 1* (2<sup>nd</sup> ed.). Milton: John Wiley & Sons.
- Lukins, N., Elvins, C., Lohmeyer, P., Ross, B., Sanders, R. & Wilson, G. (2006). *Heinemann Chemistry I* (4<sup>th</sup> edition). Melbourne: Harcourt Education.
- Maddox, J. (1998). *What Remains To Be Discovered*. London: Macmillan.
- Morris, R. (1997). *Achilles in the Quantum Universe-The Definitive History of Infinity*. London: Souvenir Press.
- Myers, R. (2003). *The Basics of Chemistry. Basics of the Hard Sciences*. USA: Greenwood Publishing Group.

- Nardelli, D. (2005). *Science Alive 3*. Milton: John Wiley & Sons.
- Niaz, M. (2001). A Rational Reconstruction of the origin of the Covalent Bond and its implications for General Chemistry Textbooks. *International Journal of Science Education*, 23(6), 623-641.
- Niaz, M. (2009). *Critical Appraisal of Physical Science as a Human Enterprise: Dynamics of Scientific Progress*. The Netherlands: Springer.
- Ohanian, H.C. (1987). *Modern Physics*. New Jersey: Prentice Hall.
- Pabuccu, A. & Geban, O. (2006). Remediating Misconceptions concerning chemical bonding through conceptual change text. *Hacettepe University Journal of Education*, 30, 184-193.
- Pearce Williams, L. (1965). *Michael Faraday*. London: Chapman and Hall.
- Peterson, R., Treagust, D. & Garnett, P. (1989). Development and Application of a Diagnostic Instrument to evaluate Grade-11 and -12 students' concepts of Covalent Bonding and Structure following a course of Instruction. *Journal of Research in Science Teaching*, 26(4), 301-314.
- Priestley, J. (1767). *The History and Present State of Electricity, with original Experiments*. London: J. Dodsley, J. Johnson & T. Cadell.
- Quam, G.N. & Quam, M.B. (1934). Types of Graphic Classifications of the Elements. *Journal of Chemical Education*, Jan., 27-32.
- Rickard, G., Burger, N., Clarke, W., Geelan, D., Jeffrey, F., Johnstone, K., Neville, C., Phillips, G., Roberson, P., Spirou, C. & Whalley, K. (2009). *Science Focus 3*. Melbourne: Pearson Heinemann.
- Robinson, W.R. (1998). An Alternative Framework for Chemical Bonding. *Journal of Chemical Education*, 75(10), 1074-1075.
- Russell, B. (1923). *ABC of Atoms*. New York: E.P. Dutton & Company.
- Serway, R.A. & Jewett, J.W. (2004). *Physics for Scientists and Engineers* (6<sup>th</sup> ed.). Australia: Thomson.
- Serway, R.A., Moses, C.J. & Moyer, C.A. (2005). *Modern Physics* (3<sup>rd</sup> ed). Australia: Thomson.
- Shaik, S. (2007). The Lewis Legacy: The Chemical Bond-A Territory and Heartland of Chemistry. *Journal of Computational Chemistry*, 28(1), 51-61.
- Spence, R., Bramley, L., Gemellarg, T., Wilson, D. & Wiseman, S. (2004). *Chemistry-A Contextual Approach*. Port Melbourne, Victoria: Harcourt Education.
- Stannard, P. & Williamson, K. (2001). *Science World* (Book 2). South Yarra, Melbourne: Macmillan Education.
- Stark, J. (1915). *Principien der Atomdynamik: Die Elektrizitat im chemischen Atom* (Teil 3). Leipzig: Hirzel.
- Sutton, C. (1984). *The Particle Connection*. London: Hutchinson.
- Taber, K.S. (1993). *Student conceptions of chemical bonding: Using interviews to follow the development of A-Level students' thinking*. Paper presented at the Facets of Education-Dimensions of Research conference, University of Surrey.  
<http://www.leeds.ac.uk/educol/documents/00001483.htm>
- Taber, K.S. (1994). Misunderstanding the ionic bond. *Education in Chemistry*, 31(4), 100-103.
- Taber, K.S. (2001a). Constructing chemical concepts in the classroom: Using research to inform practice. *Chemistry Education: Research and Practice in Europe*, 2(1), 43-51.
- Taber, K.S. (2001b). Shifting Sands: A case study of conceptual development as competition between alternative conceptions. *International Journal of Science Education*, 23, 731-753.

- Taber, K.S. (2002a). *Chemical misconceptions-prevention, diagnosis and cure-theoretical background* (Volume 1). London: Royal Society of Chemistry.
- Taber, K.S. (2002b). *Chemical misconceptions-prevention, diagnosis and cure-classroom resources* (Volume 2). London: Royal Society of Chemistry.
- Taber, K.S. (2003). Mediating mental models of metals: Acknowledging the priority of the learner's prior learning. *Science Education*, 87(5), 732-758.
- Taber, K.S. (2006). Beyond Constructivism: the Progressive Research Programme into Learning Science. *Studies in Science Education*, 42, 125-184.
- Taber, K.S. (2013). A Common Core to Chemical Conceptions: Learners' Conceptions of Chemical Stability, Change and Bonding. In G.Tsaparlis & H.Sevian (Eds.), *Concepts of Matter in Science Education* (pp.391-418). Dordrecht: Springer.
- Taber, K.S. & Coll, R.K. (2002). Bonding. In J.K. Gilbert, O. de Jong, R. Justi, D.F. Treagust & J.H.V. Driel (Eds.), *Chemical Education: Towards Research-Based Practice* (pp.213-234). The Netherlands: Kluwer.
- Talanquer, V. (2006). Commonsense Chemistry: A Model for Understanding Students' Alternative Conceptions. *Journal of Chemical Education*, 83(5), 811.
- Tan, K.D. & Treagust, D.F. (1999). Evaluating students' understanding of chemical bonding. *School Science Review*, 81(294), 75-83.
- Tan, K.C.D., Taber, K.S., Liu, X., Coll, R.K., Lorcuzo, M., Li, J., Goh, N.R. & Chia, L.S. (2008). Students' Conceptions of Ionization Energy: A Cross-cultural Study. *International Journal of Science Education*, 30(2), 263-283.
- Thickett, G. (2000). *Chemistry Pathways 1* (3<sup>rd</sup> ed.). South Yarra, Melbourne: Macmillan Education.
- Thickett, G., Stamell, J. & Thickett, L. (2000). *Science Tracks-Explorations for Australian Schools* (Year 9). South Yarra, Melbourne: Macmillan Education.
- Thornton, S.T. & Rex, A. (2000). *Modern Physics for Scientists and Engineers*. Orlando: Saunders College Publishing.
- Tsaparlis, G. (1997). Atomic and Molecular Structure in Chemical Education: A critical analysis from various perspectives of science education. *Journal of Chemical Education*, 74, 922-925.
- Unal, S., Calik, M., Ayas, A. & Coll, R.K. (2006). A review of chemical bonding studies: needs, aims, methods of exploring students' conceptions, general knowledge claims and students' alternative conceptions. *Research in Science and Technological Education*, 24(2), 141-172.
- Veltman, M. (2003). *Facts and Mysteries in Elementary Particle Physics*. New Jersey: World Scientific.
- Walding, R., Rapkins, G. & Rossiter, G. (2004). *New Century Senior Physics Concepts in Context* (2<sup>nd</sup> Ed.). South Melbourne: Oxford University Press.
- Walecka, J.D. (2008). *Introduction to Modern Physics-Theoretical Foundations*. New Jersey: World Scientific Publishing.
- Yayon, M., Mamlok-Naaman, R. & Fortus, D. (2012). Characterizing and representing student's conceptual knowledge of chemical bonding. *Chemistry Education Research and Practice*, 13(3), 248-267.