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ABSTRACT

This exercise is designed to correct the impression of many students that scientific facts are unchanging. Every scientific fact has a history and its history shapes the form which the fact takes in a textbook. Thus, the history of the octet rule is traced to show that scientific knowledge is not a cast iron set of facts but rather a fluid body of information shaped by the people who use it for specific purposes. The exercise consists of a reading and a set of discussion questions. The questions can be answered by students or could form the basis for a class discussion about the issues raised. It is suggested that the exercise be used after the lesson on the octet rule and Lewis dot structures. (JN)

CUBES, EIGHTS AND DOTS

A Student's Guide to the Octet Rule and its History

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"We do not dare to tell our students that Dalton's magnificent construction of the atomic theory was made possible by his deliberate adjustment of certain experimental results to fit his theory, since this would impugn the alledged empirical foundation of his atomic theory. ...In a debate of not too long ago, a scientific colleague argued that it was for this very reason that teachers of science would be well advised not to use accurate history. He also asked whether it would be an "edifying example" for our students to learn that Pauli discovered one of the most important principles of modern physics (the exclusion principle) while relaxing at a "girly show" in Copenhagen. The answer should be, of course, that we must decide whether we wish our students to become perfect models of what textbook writers envisage science and scientists to be, or whether we wish them to be Daltons or Paulis."

I. Bernard Cohen, "History and the Philosopher of Science" in <u>The Structure of Scientific Theories</u>, edited by Frederich Suppe (Univ. of Illinois Press, 1974) p.341.

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Knox College 1983

A NOTE TO TEACHERS

This work is designed to introduce students in an introductory chemistry class to a number of features about scientific knowledge that are seldom mentioned in a chemistry class. Our purpose is several fold. First of all we are seeking to correct the impression of many students that scientific facts are unchanging. Every scientific fact has a history and its history shapes the form which the fact takes in a textbook. In this case, the octet rule and the shared pair bond were born in association with a specific atomic model, the cubical atom, and the notation, still used today, connected with the theory had a specific meaning in terms of that model. Here we trace the evolution of the octet rule and the shared pair bond to its present textbook form. Secondly, this study shows how personalities can enter into the scientific process and change the course of scientific thinking. The role of models is demonstrated. Moreover, it can be seen that theoretical preconceptions can alter the way one thinks about scientific problems. Science does not always work from the data to the theory. After working through this study, we hope the student will have a clearer view of the way science works and a less naive attitude about the nature of scientific knowledge.

We envision that this work would be used after the lesson on the octet rule and Lewis dot structures. There are discussion questions which get at our major points and these could be answered by the students and could form the basis for a period of class discussion about the issues the study raises.

CUBES, EIGHTS AND DOTS

The purpose of this paper is in some ways simple and some ways complex. In a typical chemistry textbook you read about chemical facts a[illegible]em solving strategies; little is said about where th[illegible]s come from or how they evolved into the form which you [illegible] the textbook. The impression one gets is that the facts [illegible] no history and that they orginially took the form which you see in the textbook. This is seldom the case because most of the facts that you learn have a rich history, and they have ungone a long evolution to reach their present form. This form is generally easy to understand, remember, and simple for the beginning student to use. Such facts are the same in all textbooks. They have been reshaped by the chemists and teachers to become tools of instruction and calculation. Borrowing a phrase which is sometimes used by philosphers of science, we will call these kinds of facts <u>standardized facts</u>. Our purpose here is to take one such standard fact, the octet rule, and examine its history. We wish to demonstrate that scientific knowledge is not a cast iron set of facts but rather a fluid body of information shaped by the people who use it for specific purposes.

If you look up 'octet rule' in your textbook, you will probably find a short statement attributing the rule to the American chemist, G.N. Lewis, and then you will find a statement of the rule which goes something like this: "Atoms by sharing electrons to form an electron pair bond can acquire a stable noble-gas structure."(Masterton et.al.,Chemical Principles,Alternate Edition,2nd ed.,p.185). Remember that the stable noble-gas structure involves a valence configuration of eight electrons. Following the statement of the rule, you are introduced to the dot formula representation for the valence electrons in atoms and instructed in how to write dot structures for molecules, usually by following a set of directions to ensure that the correct formula is written. Next some examples are given in the use of the rules to predict bonding structure in molecules, and then you are invited to practice writing structures yourself. Some space is usually given over to exceptions to the rule which are taken as an indication of the lack of its validity in every case. In what follows, we will examine the history of the octet rule and trace its evolution in the work of G.N.Lewis, the chemist who is most often identified as its originator.

While Lewis first published the octet rule in 1916 (Lewis,1916), its origins go back much further. They are firmly rooted in Lewis's very original (and in our minds unusual) notion of the structure of the cubical atom. A page from his notebooks, dated 1902, is shown in Figure 1. Here he conceived of the atoms of the first three rows of the periodic chart as <u>cubes</u> with the electrons occupying places in the corners of the cube. Thus the structure of Li is shown with just one electron in the corner of the cube, the other positions being vacant; the innner electrons, as Lewis speculated, were thought to reside in smaller cubes inside the 'valence cube'.(Note, at the time Lewis thought that helium had a completed cube of eight electrons; the evidence showing He had only two electrons had not yet been discovered). The special role of eight in the

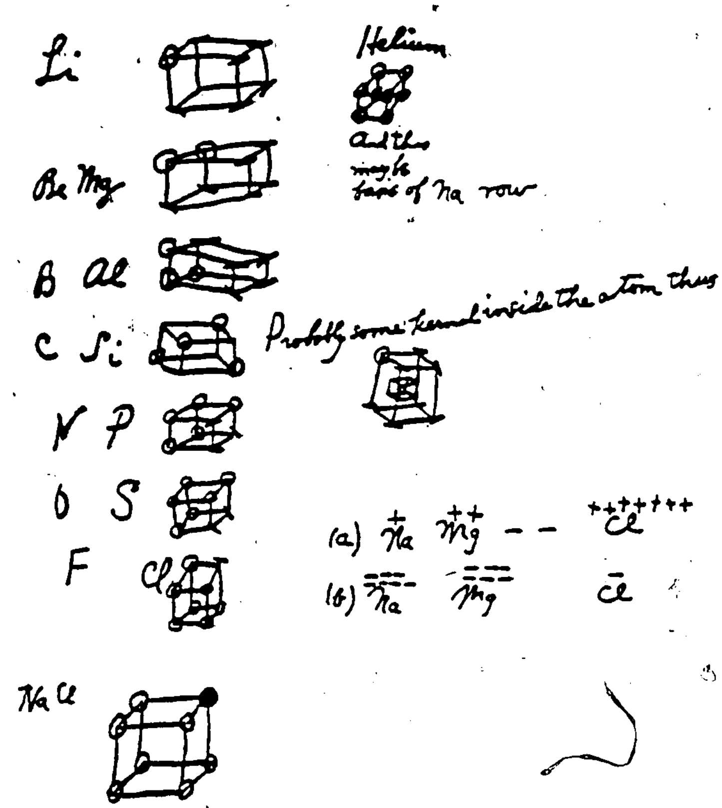


Figure 1. A page from Lewis's 1902 notebook showing his early cubical atoms.

theory of molecules was not one of Lewis's discoveries, but explaining it in terms of the electronic structure of atoms in cubical form was his novel idea. As the notebook shows, Lewis was trying to explain the bonding in ionic compounds. Thus, for NaCl sodium loses an electron to complete the octet of the chlorine atom (see Figure 1.). This leaves both the Na and Cl with completed octets, and hence in a stable configuration.

While the cubical atom may seem naive to you, you must remember that what you read in your textbook or hear in class is the result of the work of many chemists and physicists, much of which was done in the decades after 1902. The prevailing valence theory of the time was called the electrochemical theory. Basically, it assumed that chemical bonds were the result of complete transfer of electrons from one atom to another. The molecule was then held together by the attraction of unlike charges. This theory worked well for compounds that were ionic but ran into problems with covalent compounds like methane or carbon dioxide. This difficulty or anomaly, as they are now called, led some researchers to propose a dualist theory in which two distinct kinds of chemical union, one for ionic and one for covalent compounds, were said to exist. Lewis himself held a dualist view prior to his 1916 work. Dualisms in science are very unsatisfactory because rather than giving one solution to an anomaly, they make a virtue out of a vice by allowing two incompatible answers to be accepted as solutions.

Lewis did very little with his cubical atom until shortly before his paper in 1916. Others workers had been puzzling over the problem of explaining bonding in molecules. Many attempts were made to arrive at a theory that would use the recent discovery of the electron to explain the existence of ionic compounds as those compounds which transfer electrons from one atom to another to form completed octets. Also the theory had to explain the existence of what we now call covalent substances, like molecular hydrogen, where there is no apparent charge separation in the molecule. Lewis was certainly aware of the work being done by others in the field of valence and there can be little doubt that their ideas helped inspire Lewis to think on the same problems. One of Lewis's great achievements was to create a specific visual model of the atom as a cube or series of nested cubes which placed the electrons in specific places in space. Thus the notebook page of 1902 became transformed into the atoms depicted in his 1916 paper (See Figure 2).

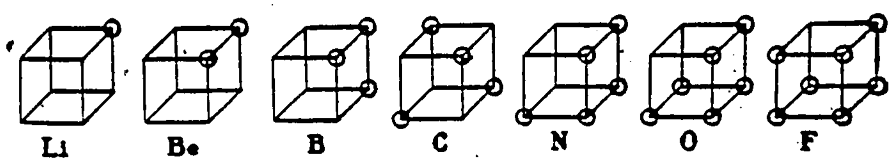


Figure 2.Cubical atoms of the first row of eight in the periodic chart (Lewis,1916).

The importance of visual models is that they condense a great deal of information into a form which someone can easily grasp and use. For example, consider how quickly you can grasp the realtionships of distance by merely looking at a road map or consider how easily you can plan many alternate ways to reach your destination. Science textbooks contain many diagrams, pictures and models because most people find them extremely useful in organizing and memorizing information. Moreover, most people feel that they have a deeper understanding of how something works when they can visualize or "see with the mind's eye" how processes interact. A famous example is picturing the ideal gas in terms of bouncing billard balls. Visual models, therefore, do three things; (1)they organize a quantity of information in a simple form, (2) they are useful tools for prediction or for planning further research and (3) they give us a deeper sense of comprehending what is going on.

The cubical atom was one of those powerful visual models which advanced chemical understanding. Ionic compounds like NaCl can be explained as a matter of course as has been noted above. The really important and novel insight that Lewis had came in the explanation of such molecules as molecular hydrogen, iodine and a host of other compounds, Figure 3 illustrates this insight. Here we see a picture of the formation of molecular iodine I_2. The electrons on edge can be shared between the two iodine atoms to simultaneously complete both octets. Because each iodine atom now has eight electrons around it, they satisfy what Lewis called the 'rule of eight' by forming an electron pair bond. The cubical atom theory allows us to picture and to understand Lewis's new idea of an electron pair bond.

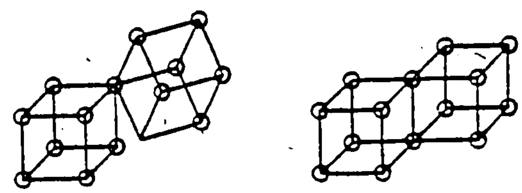


Figure 3 Formation of molecular iodine. (Lewis,1916)

With the cubical atom and the concept of the sharing of electrons to complete the octet, the structure of both ionic and covalent compounds could be understood in one comprehensive theory. In the case of NaCl, an ionic compound, the transfer of the electron to finish the octet was complete, but in a covalent molecule like ICl there was no complete transfer of electrons, merely a sharing which satisfies the rule of eight. The central idea could be taken further to explain the structure of molecular oxygen O_2. Here each O atom is deficient in 2 electrons to complete the octet; thus by sharing two electron pairs and forming two bonds (making a double bond), the stable eight configurations were preserved. This is illustrated in Figure 4. Again the power of the cubical atom was further shown in the case of ammonia NH_3. The N atom has five electrons in the valence shell, lacking three; thus three H's each with one electron to share can complete the octet. With three electron pair bonds taking 6 electrons, this leaves a pair that are not involved in bonding. A H+ with no electrons can attach to the unbonded pair to form the well known ammonium ion, NH_4^+.

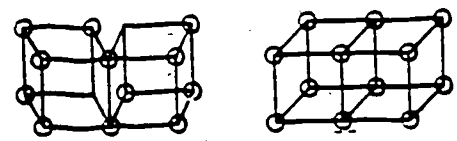


Figure 4. The double bond in molecular oxygen. (Lewis, 1916)

Today many people would say that the cubical atom is too simplistic or that it is too misleading because it makes us think of atoms as stable solid "building blocks". To judge the past in terms of the way we think today is the historical fallacy of only seeing the present in the past. To see the cubical atom in this way is to lose all sense of what an achievement it was and to fail to see how it promoted the growth of chemical knowledge. We can sometimes sense the fallacy of hindsight overtaking us when we start saying to ourselves "How silly to see things that way?", "Why did they go off in that direction?" or "Why couldn't they get the right answer?" If we begin thinking in that way, we cannot hope to understand the historical context of our present knowledge.

While the model of the cubical atom seemed to offer the key to explaining molecular structure, it was cumbersome to draw cubes tied to cubes in so many different ways. This must have been obvious to Lewis because he offered a notational system, a chemical shorthand, for representing the molecules shown in Figures 2 to 4. He proposed that the chemical symbol in boldface stand for the atom without the valence electrons (the charge on the atom would not be expressed but would be implicit in the notation). Hence the electronic kernel of lithium would be **Li** which would have a charge of +1 since Li has one valence electron. **Be** is the Be kernel with a +2 charge and so on. Lewis then proposed that compounds be expressed in the following way. Lithium fluoride(LiF) would be written as **LiFE**$_8$ where **Li** and **F** are the Li and F kernels and the E represents the valence electron count.

To make this more graphic and to show how atoms were joined together to make molecules Lewis further proposed another notation where the same kernel symbol above would be used, but the valence electrons would be represented by dots. Thus the H atom would be H. and the hydrogen molecule would then be H:H. The water molecule would be represented as H:$\ddot{\underset{..}{O}}$:H. More subtlety could be built into

this symbolism to express the fact that a compound like gaseous HCl had a shared electron pair which was more closely held to the chlorine atom than to the H atom. This unequal sharing could be expressed in the notation as H :Cl: or in the water molecule it would be H :O: H. Here we have the origin of the dot structures which you are learning to use in elementary chemistry. But it must be noted the dot structure for Lewis was only a convenient shorthand to depict molecules made up of cubical atoms. Some other examples of the use of the dot structure notation are given in Figure 5.

H:N:H + H = H:N:H (with H above and below the right N; H below the left N)

:O::O:

H:C::C:H (with H above each C)

Figure 5. Lewis dot structures for several compounds

Lewis's cubical atom worked very well for a large number of compounds but it failed for some examples in organic chemistry. For instance, the compound acetylene(HCCH) appeared to require the formation of three electron pair bonds between the two carbons in order to satisfy the rule of eight; however there is no way to join two cubes so that they can share three edges to make a triple bond. Thus as Lewis said in the case of very small atoms like carbon it is necessary to postulate a modification of the cubical atom where the eight electrons are pulled together in pairs along an edge to give a tetrahedral arrangement as

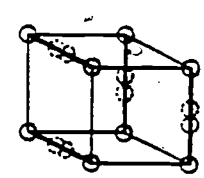


Figure 6. Tetrahedral atom from a modified cube. (Lewis,1916)

shown in Figure 6. Now it is possible for two tetrahedra to share one, two or three corners to give single, double or triple bonds as depicted in Figure 7.

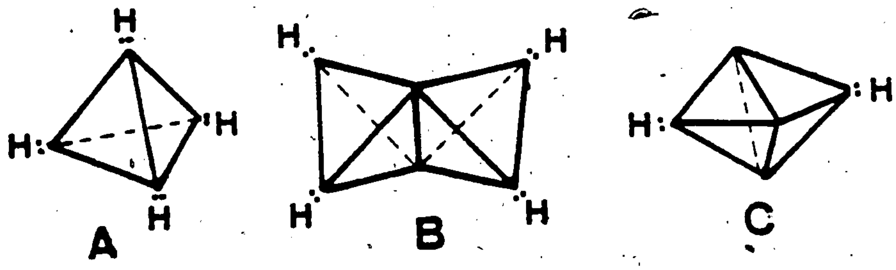


Figure 7. The tetrahedral carbon atom showing single, double and triple bonds for the compounds methane (A), ethylene (B) and acetylene(C).

The rule of eight also had what appeared to be some failures, and Lewis noted exceptions like NO_2 and ClO_2. It is not possible to complete the octet on either the N or the Cl(the dot structure for NO_2 is $:\ddot{\underset{..}{O}}:\dot{\underset{..}{N}}:\ddot{\underset{..}{O}}:$). However, Lewis regarded these exceptions, which he called odd molecules (expressing the fact they had an odd number of electrons), as the exceptions that proved the rule. First of all, odd molecules are fairly rare when compared to the number of molecules that have an even number of electrons and obey the octet rule. Secondly, these molecules are quite reactive and when they do react their products are even electron molecules that obey the octet rule. The instability and rareness of these substances demonstrated the universality of the octet rule for Lewis.

After the article's appearance in 1916, very little attention was given to Lewis's ideas. Lewis himself turned to other pursuits, some brought on by his involvement in work related to World War I. Other scientists were not impressed or did not understand the tremendous explanatory power of the octet rule and the shared pair bond. If this lack of recognition had continued, you probably would not be studying the dot notation now. It would have died unrecognized as so many chemical facts do. This was not to be the fate of Lewis's work, however, because Irving Langmuir, another famous American chemist, was puzzling over the theory of the structure of molecules. He found Lewis's ideas very powerful in explaining a large range of phenomena. Langmuir communicated his understanding of the octet rule and the shared pair bond in a long paper that appeared in <u>The Journal of the American Chemical Society</u> in 1919 (Langmuir,1919). He cited the paper of Lewis as being very important and proceeded to elaborate on Lewis's ideas and apply them to a wide range of substances.

It is interesting to see what Langmuir did with Lewis's work. Like Lewis, Langmuir took the cubical atom as thoroughly real. It was not a fiction. Atoms were cubical. Figure 8 shows some of Langmuir's constructions of the oxides of nitrogen. Secondly, Langmuir attempted a much more rigorous deductive presentation of the "octet rule" (a catchphrase which replaced what Lewis called the "rule of eight"). In contrast to the tersely written but more intuitive arguments of Lewis, Langmuir, in his 1919 paper, had eleven postulates which described the arrangement of the electrons in the atom and the stability of certain outer electron arrangements. His work followed the model of Euclidean geometry in presenting postulates, proofs and derivations.

He derived a formula

$$e = 8n - 2p \quad (1)$$

where *e* is total number of valence electrons in a molecule, *n* is the number of octets formed and *p* is the number of electron pair bonds in the molecule. He then proceeded to use the postulates and his equation (1) to determine the structure and properties of various molecules.

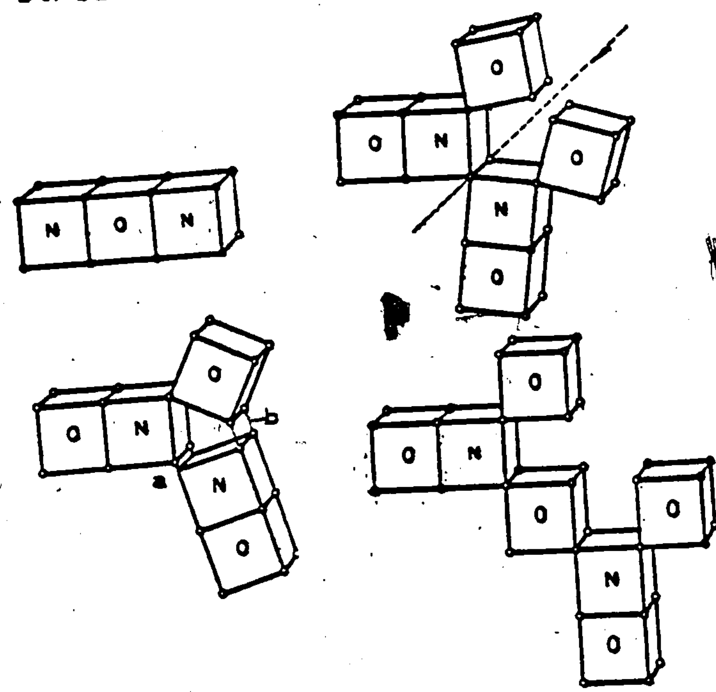


Figure 8. Languir's cubical atom depicting the molecular arrangements for various nitrogen oxides. (Langmuir 1919)

As just one example, take the nitrogen molecule N_2 which has two octets to complete in the two nitrogen atoms(n=2) and has 10 valence electrons(e=10). Langmuir's equation then yields p=3 or three electron pair bonds. Langmuir's conceptual attachment to the cubical atom and stability of a structure which had 8 electrons in the corners resulted in a rather curious picture for the nitrogen molecule. Nitrogen is exceptionally unreactive so Langmuir assumed that there must be an unusual cubical arrangement of 8 electrons to explain this fact. However, there were 10 valence electrons and a necessity for 3 shared pairs. Given these constraints, Langmuir postulated the structure shown in Figure 9. The black dots are the nitrogen kernels, the 6 shared electrons in pairs are inside the the outer completed octet which evidently was to give the molecule its stability.

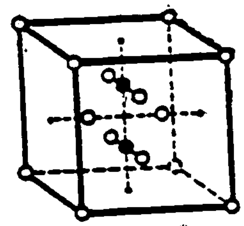


Figure 9. Langmuir's nitrogen molecule. (Langmuir, 1919)

Nowhere in Langmuir's paper does the Lewis dot structure appear; consequently, the result is a much more mathematical and logical presentation of the octet rule and the electron pair bond. Langmuir was also able to correlate electronic structure of molecules with their properties and noticed that molecules with similar bonding had similar chemical properties. For instance, both CO and N_2 have triple bonds. Langmuir noted that "the physical properties of these two gases are identical almost within the probable limits of experimental error"(Langmuir, 1919) and proved so by a table of properties. All in all, the paper was an impressive display of the validity of the concepts that Lewis had published in 1916.

One paper such as Langmuir's would not be enough to insure that the octet rule would live on. Rather what made the difference was Langmuir's personal enthusiasm for what he had rediscovered in Lewis's earlier work. In three short years Langmuir published 12 papers detailing his applications.

of the octet rule and he gave numerous popular lectures both in the United States and in England. The volume of material and the wide audience Langmuir generated, along with the fact that Langmuir was held in high esteem by the chemical community and was a dynamic and popular lecturer, resulted in immediate acceptance of the octet rule. It is quite clear that it was the force of Langmuir's personality and reputation rather than a real crisis in chemical bonding theory that resulted in the octet rule and the shared electron pair becoming the predominent valence theory.

In fact, Langmuir's presentation was so forceful that the theory began to be known as the the Lewis-Langmuir theory or sometimes just 'Langmuir's theory' despite the fact that Langmuir added nothing new to the concepts that Lewis had presented in 1916. Langmuir's success and acclaim was ultimately a source of irritation to Lewis. Lewis wrote later to W.A.Noyes(Kohler,1974)

> Perhaps I am inclined to be too caustic in this matter, but I really do feel that while people were justified in being carried away a bit by Langmuir's personal charm and enthusiasm some years ago, to persist, as they do in England, in speaking of the Langmuir theory of valence is inexcusable.

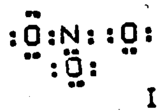
Langmuir's success was probably the prod for Lewis to once again return to valence theory by publishing a monograph entitled <u>Valence and the Structure of Atoms and Molecules</u> in 1923 (Lewis,1923). This work was widely read and frequently used as a textbook.

By 1923, the work of the Danish physicist Neils Bohr and others had made the concept of a static cubical atom untenable. Their work showed that the electrons in an atom must be in motion and could not occupy fixed positions in space. Fixed static atoms in the shape of building blocks could not be "real". Thus, in his monograph Lewis mentions the cubical atom only as an historical fact. He no longer claimed that atoms were really cubes. Instead, he attempted to rationalize the apparent fixed positions of electrons in chemcial bonds as the average position of the dynamical electron. Although, Lewis tended toward an atomic structure that placed the electron pairs at the corners of a tetrahedron, this concept was mentioned only briefly and did not play an important role in his exposition of the octet rule.

With so many changes in his original idea, you may wonder why Lewis's book was widely praised and used. Clearly the cubical atom was out, but something more important was permanently in the minds of chemists; namely the dot notation and the octet rule. In his book, Lewis leaned heavily on his dot structure notation to explain the bonding in molecules. He stated the octet rule in a form familiar to you (Lewis,1923).

> Two atoms may conform to the rule of eight, or octet rule, not only by the transfer of electrons from one atom to another, but also by sharing one or more pairs of electrons. These electrons which are held in common by two atoms may be considered to belong to the outer shells of both atoms.

Lewis's book showed others how to apply the rule to solving problems, just as you have done. Thus by applying the rule to the nitrate ion NO_3^- he was able to show that it had a double bond between one of the nitrogen's and an oxygen



It is interesting to compare Lewis's treatment of the nitrogen molecule with that of Langmuir. Lewis no longer tried to depict the atoms as cubical and instead relied on his dot notation. Thus nitrogen now became :N:::N:. Many of the examples in his book are the same ones used in his 1916 paper, although there are new ones introduced. Some of the structures from Lewis's book appear in Figure 10. It should be noted too that Lewis did not use Langmuir's formula (Equation (1)) in his book, preferring instead to rely on the visually effective dot notation to complete the electron counts about the atoms in a molecule.

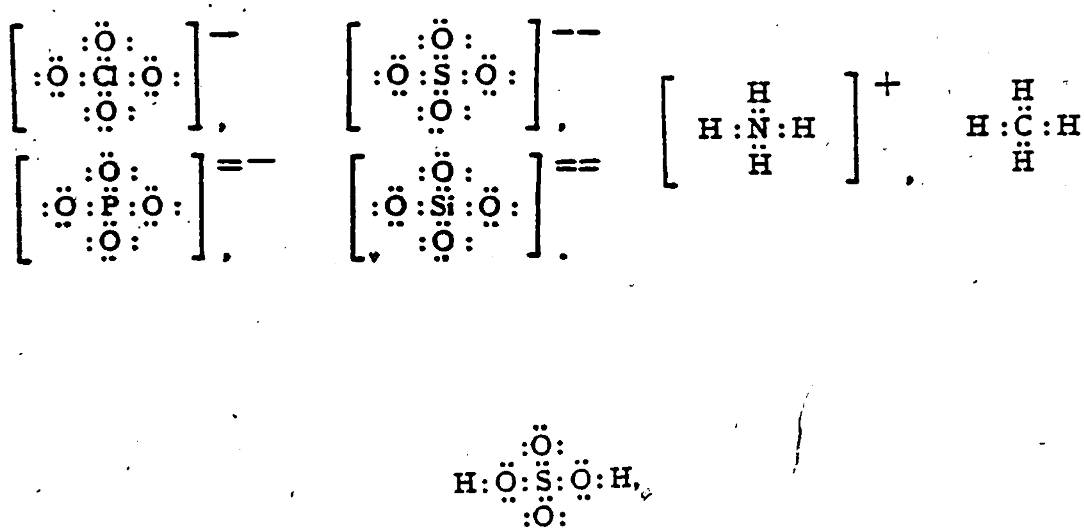


Figure 10. Various compounds whose electronic structure satisfies the octet rule. (Lewis 1923)

By the time Lewis's book had appeared, the chemcial community, through the popularizing figure of Langmuir, was ready to accept the valence theory that Lewis had created. It became a widely read text not only for chemists but for students of chemistry as well. Indeed it was the definitive text on valence until the work of Linus Pauling appeared in the 1930's.

To summarize, Lewis took the dot notation and the octet rule, which were formerly a part of an abstract and controversial theory about the nature of atoms and transformed them into tools for solving problems. He created a standardized fact by detaching the octet rule and the dot notation from their theoretical context and placing them in a more useful and practical role. To understand the importance of this, you merely need to reflect upon how it is possible for you or any beginning student to learn science. You obviously can't begin by studying the most complicated new theories. One must work their way up to that. Every student of science begins by mastering concepts which can be applied in problems with known solutions. You are learning what is sometimes called "normal science" which consists of the standardized facts organized in textbook form. These facts carry along only hints or shadows of their original beginnings. Perhaps you may have asked yourself why we use the dot notation. Working backwards from standadized facts you will find the discarded theories. And if you were to work forward from standardized facts, what would you find? Theories again and problems and experiments to which no answers are certain. Science on its creative edge is potentially revolutionary science. It involves free use of the imagination in much the same way that Lewis imagined the cubical atom and, of course, it requires a mastery of standardized facts. Science as it appears in textbooks and science as it is practiced by researchers into the unknown are two different things. Professor Gerald Holton of Harvard University once suggested that we call the former Science I and the latter Science II. Since most of us will never do Science II, we put our energies into learning Science I, but as we have seen Science II is historically behind and underneath the standardized facts we learn about in Science I.

In this case study you have encountered some characteristics which are generally true of the creative advance of science. First there is a high premium placed upon being first to come up with a solution to a tough problem. Research is competitive and this leads to personal disputes like the Lewis-Langmuir controversy. Many people find this hard to accept because they believe the myth that science is purely objective and hence all scientists must be in complete agreement. This myth results from confusing Science I with Science II. There is universal agreement about the standardized facts in Science I but in Science II the new ideas are being forged, and there is much disagreement about which new ideas are the right ones.

Secondly, we saw that Lewis was flexible in his thinking. He proposed the cubical atom, changed his mind in proposing the tetrahedral model and ultimately accepted a dynamic planetary model. Perhaps he saw that the important thing was the dot structure notation since it was compatible with many theories of the atom.

Thirdly, Science II depends heavily upon the power to visualize, picture or model unknown relationships. Creativity in Science II is much like creativity in other fields like art and literature because it calls for the free play of the imagination disciplined by familiarity with accepted facts.

Finally we have seen that in both Science I and Science II it is absolutely essential that scientific inquiry be carried on in a community of investigators which has free access to information and which shares common judgments about the adequacy, the quality and significance of new ideas. Both Langmuir and Lewis addressed the professional chemists of their day. These men were, perhaps, not as gifted or as creative but they knew that the Lewis-Langmuir work was a significant contribution because they could apply it to additional problems. Great ideas in science and elsewhere are always fruitful.

DISCUSSION QUESTIONS

1.J.J. Thomson, the great English physicist, had an early model for the atom where the electrons were embedded in a positive sea of charge(this is sometimes called the 'plum-pudding' model). What advantages do you see for the 1902 Lewis model for explaining the exsistence of say carbon dioxide CO_2?

2.Discuss the role that Lewis's model had in his formulation of the octet rule and electron pair bond. Is it likely that Lewis arrived at these observations from the data(i.e. the properties and structures of compounds) or did he use the cubical atom to fit the data? What advantages accrue from the fact that the model was geometric and pictorable? What role does the fact that Lewis was forced to modify the cubical atom to a tetrahedral one play in this discussion?

3 List the special features of the notation systems that Lewis developed. Remember he put forth two systems; one for the total valence electron count and the other using the dot notation.

4.What role did Langmuir play in the development of the octet rule?

5.How did Langmuir change the octet rule and the shared pair bond? List the differences in Langmuir's presentation and that done by Lewis in 1916. List the similarities? Can it be said that Langmuir's views are the same as Lewis's in 1916? Can it be said Langmuir's presentation of the octet rule and shared pair bond changed these concepts?

6.Langmuir's model for the nitrogen molecule is interesting. What things made him postulate the structure shown in Figure 9? Which of these things was the overriding feature that lead to the model? Did he primarily come from the data or were theoretical constraints the most important?

7.Why should Lewis have become irritated by Langmuir's work? In other words, what is there about scientific knowledge that one should be concerned with whose ideas are whose? Isn't the important fact that knowledge and new insights were gained?

8.Compare Lewis's 1923 presentation with his 1916 one. Focus particularlly on the dot notation as given in Figure 10 and in the text describing the 1916 work. What changes have taken place? Do the 1923 figures mean the same as the 1916 representations? What happened to the the cubical atom?

9.Compare the presentation of the octet rule and the dot notation given in your textbook or class notes with both the 1916 ideas of Lewis and the 1923 work. In what ways has the standardized textbook fact remained the same and in what ways does it differ from the 1916 and 1923 forms? Can it be said that the octet rule and the shared electron pair have undergone change?

10.How much of the history of the octet rule is given in your text? Is it important to know this history to use the octet rule and the dot notation? Why is this so?

11. If Lewis had not published his 1923 book, do you think your textbook's treatment of the octet rule would have been different? Explain your answer and in doing so think carefully about Langmuir's role in popularizing these concepts.

BIBLIOGRAPHY

To make the text readable, we have not footnoted the facts and assertions we have made about the story of the evolution of the octet rule. Much of what we have said has come from a study of the excellent work of Robert E. Kohler in the history of the octet rule. Additionally we have studied the original work of Lewis and Langmuir. Lewis's book, Valence, still remains a model of clarity and readability and it entirely possible to teach the octet rule by excerpts from it. The works cited in the text are

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In addition, we found the following sources extremely valuable in this work:

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