Inquiry-Based Methods of Instruction in Chemistry

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Abstract

Inquiry based methods of instruction are designed to make students active participants in constructing their own knowledge. Inquiry based methods may involve hands-on laboratory work, computer techniques, group activities or lectures. In this paper I will demonstrate how the important chemical concepts, valence and Lewis electron dot structures, can be taught using inquiry based lecture techniques.

Introduction

Research indicates that student achievement and motivation for the study of science improves dramatically if students are active participants in constructing their own knowledge and in learning to use that knowledge to analyze scientific processes. Inquiry based teaching methods are designed to make the students active participants. Such methods are now extensively used in physics education {1-5}but are in their infancy in chemical education.

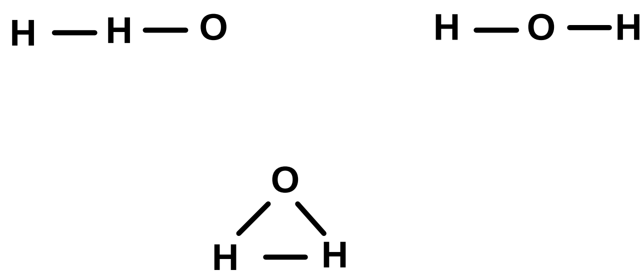
In this paper I will show how inquiry based lecture methods can be employed to teach students the important chemical concepts of valence and Lewis electron dot structures. The valence concept was devised in the middle of the nineteenth century and G. N. Lewis was the one of the first scientists to explain valence in terms of electronic structure of atoms. Lewis' papers on this subject were published in the early part of the twentieth century. High school and college chemistry texts usually present valence as a consequence of the Lewis concept in spite of the fact that the valence concept predates the Lewis concept by fifty years. With an inquiry based approach, valence is introduced first, and the students themselves create the concept using the same logical reasoning that the chemists who devised the concept used. The Lewis concept is then introduced, and the students use the same logical reasoning that Lewis used. In this way students have a much better understanding of both concepts than is the case when they are taught the concepts using traditional methods.

Valence

By the middle of the nineteenth century chemists had learned that atoms unite to form molecules and that the atoms in molecules are held together by strong bonds. They also knew how many bonds an atom of a given element can use to link itself to other atoms. Let us examine a few examples of molecules to see how they deduced this. An example of a simple molecule is HF. This molecule consists of two atoms bonded together. Let's represent this bond with a dash between the symbols of the elements as follows.

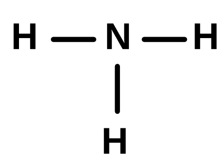
$$\mathbf{H — F}$$

Now let us consider a water molecule, H_2O. There are three possible ways we can put the three atoms in a water molecule together as shown below.



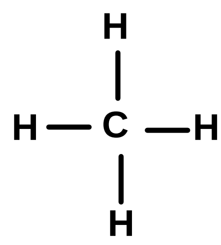
We can have the two hydrogen atoms bonded together with the oxygen atom bonded to one of them as indicated in the upper left, or we can have the two hydrogen atoms bonded to the oxygen atom but not to each other as in the upper right, or we have the three atoms bonged together in a cyclic fashion as in the lower structure. Which of these three structures makes the most sense? In the structure at the upper left, one of the hydrogen atoms is bonded to two atoms and the other one is only bonded to one. The two identical atoms are exhibiting different bonding characteristics. In the lower structure the two hydrogen atoms are bonded in the same fashion but they are exhibiting different bonding behavior than in the case of HF. In the other structure the two hydrogen atoms are exhibiting the same bonding characteristics, and moreover, they are behaving the same as in HF, that is, each hydrogen atom is bonded to only one atom. This structure makes the most sense; the differences in the structures of H_2O and HF are due to differences in the bonding behavior of oxygen and fluorine not hydrogen.

Now let us extend these ideas to an ammonia molecule, NH_3. There are several ways we could bond the four atoms in an ammonia molecule together, but the one that makes the most sense is one in which the three hydrogen atoms are bonded to the nitrogen atoms as follows.



In this structure the three hydrogen atoms are exhibiting the same bonding characteristics as in HF and H_2O, but nitrogen is behaving differently than fluorine or oxygen as we might expect.

If we extend these ideas to methane, CH_4, we come to the same conclusion, that the four hydrogen atoms are bonded to the carbon atom.

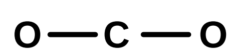


Chemists in the nineteenth century reached the same conclusions concerning the bonding in these molecules and represented the bonds in the same fashion we have. This practice is continued today.

The number of bonds an atom can form is called its **valence**. Hydrogen is said to have a valence of one, or hydrogen in univalent. Similarly, oxygen exhibits a valence of two or is divalent, nitrogen has a valence of three or is trivalent and carbon has a valence of four or is tetravalent.

The valence of hydrogen, oxygen, nitrogen and carbon have been deduced from the formulae of just a few molecules, HF, H_2O, NH_3 and CH_4. Now let make the assumption that hydrogen, oxygen, nitrogen and carbon exhibit the same valences in other compounds and attempt to explain bonding in light of this assumption.

First let us consider the carbon dioxide molecule, CO_2. If we treat CO_2 in a manner similar to that used for H_2O we might come to the conclusion that CO_2 has the following valence structure.



But in this structure the carbon atom is exhibiting a valence of two rather than four and the oxygen atoms are exhibiting a valence of one rather than two. We can fix both problems by assuming that each of the two oxygen atoms is linked to the carbon atom by two bonds as follows.

This is the same conclusion reached by nineteenth century chemists. Bonds of this type are often referred to as **double bonds.** The bonds in HF, H_2O, NH_3 and CH_4 are called **single bonds**.

Now consider the structure of a hydrogen cyanide molecule, HCN. We would like to put these three atoms together so that the hydrogen atom exhibits a valence of one, the carbon atom a valence of four and the nitrogen atom a valence of three. Which of the three atoms should be placed in the middle of the valence structure? Clearly the hydrogen atom won't work if it is to assume a valence of one and form only one bond. What about the nitrogen atom? If the nitrogen atom is to exhibit a valence of three, one of the bonds will have be used to bond to the hydrogen atom, leaving only two bonds to be used to bond the nitrogen atom to the carbon atom as follows. Although both hydrogen and nitrogen exhibit their expected valence, carbon does not.

$$H-N=C$$

What about placing the carbon atom in the center? If we bond the hydrogen atom to the carbon atom with a single bond we then have three bonds to bond the carbon atom to the nitrogen atom as follows.

$$H-C\equiv N$$

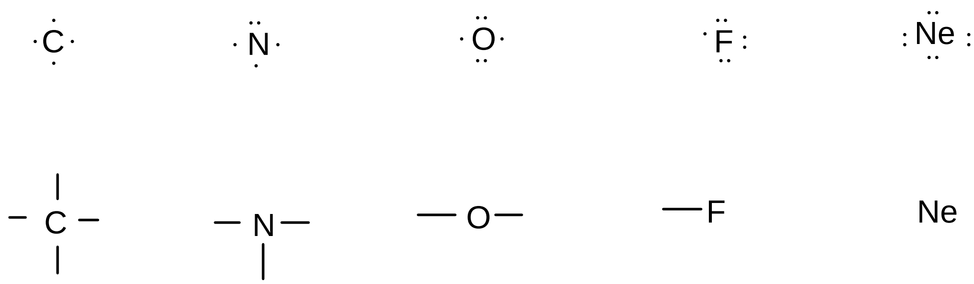
In this way the three atoms exhibit their expected valences. This is also the same conclusion reached by the chemists of the nineteenth century. Bonds like that between the carbon atom and the nitrogen atom are called **triple bonds**.

Chemists extended these ideas to the hundreds of compounds of carbon, hydrogen, oxygen , nitrogen and fluorine known at the time and found that with very few exceptions the structures were consistent with the notions that atoms of these elements exhibit the same valences illustrated in the few examples we have been considering and that single, double and triple bonds are possible.

Lewis Electron Dot Structures

In 1916 the American chemist G. N. Lewis showed how the concept of valence could be explained on the basis of electronic structures of the atoms. By this time the nuclear atom model had been established and the number of planetary electrons contained in the atoms of each element was known. Hydrogen atoms each contain a single electron and the atoms of carbon, nitrogen, oxygen , fluorine and neon atoms contain 6, 7, 8, 9 and 10, respectively. Lewis assumed that two of the electrons in atoms of each of these latter elements are in an inner core close to the nucleus and the remaining electrons are at the outer part of the atoms in each case. He was able to show that these outer electrons are responsible for the valence of the respective atoms, and referred to these electrons as valence electrons. He represented these valence electrons by dots around the symbols of the elements as shown below. This representation is in wide use today, and structures using this symbolism are called Lewis electron dot structures.

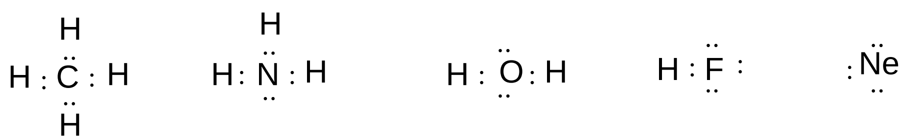
Let us see how Lewis accounted for valence. Below are Lewis electron dot structures along with the valence structures of these atoms. It is important to note that neon exhibits no valence, that is, its atoms do not unite with each other or with atoms of other elements.



Lewis dot structures 1916

valences 1865

Do we see a correlation between the valences and the Lewis electron dot structures? Yes, for each dot missing from the neon situation there is a bond. Hydrogen atoms bond to the atoms above, and each hydrogen atom has one electron. Thus, if we consider each bond as a pair of electrons being shared between the two atoms bonded together, one electron from the hydrogen atom and the other electron from the other atom, we see that each of the central atoms has the same number of electrons as the neon atom.



If was this reasoning that led G. N. Lewis to the notion of electron pair bonds and to the octet rule concept for covalent bonding.

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Author's Biography

Dr. M. Dale Alexander, Professor Emeritus of Chemistry in the Department of Chemistry and Biochemistry at New Mexico State University, received his B. S. (with honors) with a major in chemistry from NMSU in 1960 and his Ph.D. in the area of inorganic chemistry from The Ohio State University in 1964 at which time he joined the faculty at NMSU. During his career, he has received several awards including the Westhafer Award for Excellence in Teaching, NMSU (1985), the Burlington Northern Foundation Faculty Achievement Award, NMSU(1986), the PNM Foundation Distinguished Educator Award, state of New Mexico (1988), the Donald C. Roush Award for Teaching Excellence. NMSU (1987 and 1993), the Dennis W. Darnall Faculty Achievement Award for Excellence in Teaching, Research and Service, NMSU (1995), the James F. Cole Memorial Award for Service, NMSU Alumni Association (1988), and the President's Award for Service (1992), presented by President James E. Halligan, NMSU.

Dr. Alexander's current research interests lie in the area of chemical education, and he has had numerous publications in this area. Dr. Alexander retired in 1999, but continues to teach at NMSU on a part time basis, and he still pursues his research interests and his chemistry outreach program.