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One hundred years of the 'bond'

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The chemical bond has just crossed a century. The 'bond' has been central to science, second only to atoms perhaps, both of which together make the physical world.

In the last 100 years the bond has been manipulated to yield, among other things, sophisticated drugs.

In the past hundred years, the bond has been manipulated in numerous ways, which made drugs, polymers, plastics, dyes, detergents, agrochemicals, liquid crystals, and many others possible and in the process, the world around us has changed irreversibly. We synthesised over nine million compounds, making the chemical industry the largest in the world, second only to energy.

In a celebrated paper in the Journal of the American Chemical Society in 1916, G. N. Lewis introduced the electron pair bond. He proposed that two atoms may share from one to six electrons forming single, double or triple bonds. He introduced the cubical atom and six postulates to understand their chemical behaviour.

We can now design the bond at will and get desired properties from the material so produced. Atoms can be assembled the way we want, to get the shapes we need. Many of these capabilities have inputs from several branches of science, but the underlying reason is the understanding of the chemical bond itself.

In the 1916 paper, Lewis wrote, “An electron may form a part of the shell of two different atoms and cannot be said to belong to either one exclusively”. Remember that in 1916 quantum mechanics was nonexistent which helps us today to understand the atomic world. The Bohr atom model was introduced just around this time, in 1913. The wave-particle duality had come only in 1924. The Schrodinger wave equation came only in 1926, although the interpretation of the wave was unclear. At the time, sharing of an electron by two atoms, conceptually, was a giant step forward.

Lewis wrote about the electron shell model and proposed that atoms acquire a configuration containing eight electrons in the process of forming chemical bonds. He introduced a symbolism for the bond — A:B, where the colon implies the existence of an electron pair bond involving the sharing of two electrons between atoms A and B. He symbolised this as A: B and A :B implying polar bonds, where electrons are closer to A and B, respectively. He had, therefore, models for different types of bonds such as covalent and ionic — in the former, electrons are shared and in the latter, an electron from one atom is transferred completely to the other.

The foresight of Lewis is evident in the last paragraph of the paper in which he wrote about the properties he did not consider such as the great variety of colours of transition metal compounds. He wrote, “It seems probable that in these elements there is a possibility of the transfer of electrons either from one part of the kernel to another, or between the kernel and the outer shell, or possibly between two separate outer shells of the same atom, and that electrons which are suspended midway between two such stages are responsible for the absorption of light in these cases.” Lewis was talking about d-d transitions, perhaps!

Today, we teach the concepts of Lewis in classrooms across the world. Among the many concepts of chemistry that have become universal is the concept of electron pair bond and how the bond is represented.

We must not forget the long quest towards understanding chemical affinity. Newton wrote about the forces between particles at short distances explaining chemical operations. A theory of chemical combination was developed by Berzelius in 1819. Kekule is considered to be the principal founder of the theory of chemical structure and his structure of benzene is a landmark in chemistry. The first quantum mechanical description of the chemical bond came in 1927, with an understanding of the positively charged ion of the hydrogen molecule. The model, although quantitative, could not be extended to molecules with more electrons. Soon, many other models came into existence and today, we can calculate the structure and properties of complex molecules accurately.

The bond has gone beyond molecules to build supramolecular chemistry. Many different types of bonds such as hydrogen bonds came in with which we understand the unusual properties of systems such as water and ice.

That chemical bond decides the ‘horoscope’ of the molecule. It decides the properties of molecules such as colour, reactivity, solubility, etc. all of which decide their applications. Besides, it decides the physical properties of matter. Metals, ceramics, and semiconductors are what they are due to their chemical bonding. Today, structure and bonding are essential aspects of chemical pedagogy. It is a pity to note that the chemical bond did not get its due recognition in the form of a Nobel Prize. The all-pervading reach of the bond is elegantly captured in the celebrated book of Linus Pauling, The Nature of the Chemical Bond, which is dedicated to Lewis himself.

Looking through my window, I see that the world outside is built by bonds. Indeed, the window itself is made by bonds, so also the capability of the eye to capture the vibrant images outside. For the common man, the affinity between people is chemistry, although he does not recognize the quantum nature of bonds!

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