

Chemical Bonds- Interaction Between Atoms

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Abstract: *Living things are made up of atoms, but in most cases, those atoms aren't just floating around individually. Instead, they're usually interacting with other atoms (or groups of atoms).*

For instance, atoms might be connected by strong bonds and organized into molecules or crystals. Or they might form temporary, weak bonds with other atoms that they bump into or brush up against. Both the strong bonds that hold molecules together and the weaker bonds that create temporary connections are essential to the chemistry of our bodies, and to the existence of life itself.

Why form chemical bonds? The basic answer is that atoms are trying to reach the most stable (lowest-energy) state that they can. Many atoms become stable when their valence shell is filled with electrons or when they satisfy the octet rule (by having eight valence electrons). If atoms don't have this arrangement, they'll "want" to reach it by gaining, losing, or sharing electrons via bonds.

Introduction

A chemical bond is a lasting attraction between atoms that enables the formation of chemical compounds. The bond may result from the electrostatic force of attraction between atoms with opposite charges, or through the sharing of electrons as in the covalent bonds. The strength of chemical bonds varies considerably; there are "strong bonds" or "primary bond" such as metallic, covalent or ionic bonds and "weak bonds" or "secondary bond" such as Dipole-dipole interaction, the London dispersion force and hydrogen bonding.

Since opposite charges attract via a simple electromagnetic force, the negatively charged electrons that are orbiting the nucleus and the positively charged protons in the nucleus attract each other. An electron positioned between two nuclei will be attracted to both of them, and the nuclei will be attracted toward electrons in this position. This attraction constitutes the chemical bond. Due to the matter wave nature of electrons and their smaller mass, they must occupy a much larger amount of volume compared with the nuclei, and this volume occupied by the electrons keeps the atomic nuclei in a bond relatively far apart, as compared with the size of the nuclei themselves.

Discussion And Results

Early speculations about the nature of the chemical bond, from as early as the 12th century, supposed that certain types of chemical species were joined by a type of chemical affinity. In 1704, Sir Isaac Newton famously outlined his atomic bonding theory, in "Query 31" of his *Opticks*, whereby atoms attach to each other by some "force". Specifically, after acknowledging the various popular theories in vogue at the time, of how atoms were reasoned to attach to each other, i.e. "hooked atoms", "glued together by rest", or "stuck together by conspiring motions", Newton states that he would rather infer from their cohesion, that "particles attract one another by some force, which in immediate contact is exceedingly strong, at small distances performs the chemical operations, and reaches not far from the particles with any sensible effect."

In 1819, on the heels of the invention of the voltaic pile, Jöns Jakob Berzelius developed a theory of chemical combination stressing the electro negative and electropositive characters of the combining atoms. By the mid 19th century, Edward Frankland, F.A. Kekulé, A.S. Couper, Alexander Butlerov, and Hermann Kolbe, building on the theory of radicals, developed the theory of valency, originally called "combining power", in which compounds were joined owing to an attraction of positive and negative poles. In 1916, chemist Gilbert N. Lewis developed the concept of the electron-pair bond, in which two atoms may share one to six electrons, thus forming the single electron bond, a single bond, a double bond, or a triple bond; in Lewis's own words, "An electron may form a part of the shell of two different atoms and cannot be said to belong to either one exclusively." [2]

That same year, Walther Kossel put forward a theory similar to Lewis' only his model assumed complete transfers of electrons between atoms, and was thus a model of ionic bonding. Both Lewis and Kossel structured their bonding models on that of Abegg's rule (1904).

In 1927, the first mathematically complete quantum description of a simple chemical bond, i.e. that produced by one electron in the hydrogen molecular ion, H_2^+ , was derived by the Danish physicist Oyvind Burrau. [3] This work showed that the quantum approach to chemical bonds could be fundamentally and quantitatively correct, but the mathematical methods used could not be extended to molecules containing more than one electron. A more practical, albeit less quantitative, approach was put forward in the same year by Walter Heitler and Fritz London. The Heitler-London method forms the basis of what is now called valence bond theory. In 1929, the linear combination of atomic orbitals molecular orbital method (LCAO) approximation was introduced by Sir John Lennard-Jones, who also suggested methods to derive electronic structures of molecules of F_2 (fluorine) and O_2 (oxygen) molecules, from basic quantum principles. This molecular orbital theory represented a covalent bond as an orbital formed by combining the quantum mechanical Schrödinger atomic orbitals which had been hypothesized for electrons in single atoms. The equations for bonding electrons in multi-electron atoms could not be solved to mathematical perfection (i.e., analytically), but approximations for them still gave many good qualitative predictions and results. Most quantitative calculations in modern quantum chemistry use either valence bond or molecular orbital theory as a starting point, although a third approach, density functional theory, has become increasingly popular in recent years.

In 1933, H. H. James and A. S. Coolidge carried out a calculation on the dihydrogen molecule that, unlike all previous calculation which used functions only of the distance of the electron from the atomic nucleus, used functions which also explicitly added the distance between the two electrons. [4] With up to 13 adjustable parameters they obtained a result very close to the experimental result for the dissociation energy. Later extensions have used up to 54 parameters and gave excellent agreement with experiments. This calculation convinced the scientific community that quantum theory could give agreement with experiment. However this approach has none of the physical pictures of the valence bond and molecular orbital theories and is difficult to extend to larger molecules.

Ionic bonding involves a transfer of an electron, so one atom gains an electron while one atom loses an electron. One of the resulting ions carries a negative charge (anion), and the other ion carries a positive charge (cation). Because opposite charges attract, the atoms bond together to form a molecule.

The most common bond in organic molecules, a covalent bond involves the sharing of electrons between two atoms. The pair of shared electrons forms a new orbit that extends around the nuclei of both atoms, producing a molecule. There are two secondary types of covalent bonds that are relevant to biology —

polar bonds and hydrogen bonds.

Two atoms connected by a covalent bond may exert different attractions for the electrons in the bond, producing an unevenly distributed charge. The result is known as a polar bond, an intermediate case between ionic and covalent bonding, with one end of the molecule slightly negatively charged and the other end slightly positively charged.

These slight imbalances in charge distribution are indicated in the figure by lowercase delta symbols with a charge superscript (+ or -). Although the resulting molecule is neutral, at close distances the uneven charge distribution can be important. Water is an example of a polar molecule; the oxygen end has a slight positive charge whereas the hydrogen ends are slightly negative. Polarity explains why some substances dissolve readily in water and others do not.

Because they're polarized, two adjacent H₂O (water) molecules can form a linkage known as a hydrogen bond, where the (electronegative) hydrogen atom of one H₂O molecule is electrostatically attracted to the (electropositive) oxygen atom of an adjacent water molecule.

Consequently, molecules of water join together transiently in a hydrogen-bonded lattice. Hydrogen bonds have only about 1/20 the strength of a covalent bond, yet even this force is sufficient to affect the structure of water, producing many of its unique properties, such as high surface tension, specific heat, and heat of vaporization. Hydrogen bonds are important in many life processes, such as in replication and defining the shape of DNA molecules.

Like hydrogen bonds, London dispersion forces are weak attractions between molecules. However, unlike hydrogen bonds, they can occur between atoms or molecules of any kind, and they depend on temporary imbalances in electron distribution.

How does that work? Because electrons are in constant motion, there will be some moments when the electrons of an atom or molecule are clustered together, creating a partial negative charge in one part of the molecule (and a partial positive charge in another). If a molecule with this kind of charge imbalance is very close to another molecule, it can cause a similar charge redistribution in the second molecule, and the temporary positive and negative charges of the two molecules will attract each other.

Hydrogen bonds and London dispersion forces are both examples of van der Waals forces, a general term for intermolecular interactions that do not involve covalent bonds or ions. Some textbooks use the term "van der Waals forces" to refer only to London dispersion forces, so make sure you know what definition your textbook or teacher is using.

Conclusion

In the (unrealistic) limit of "pure" ionic bonding, electrons are perfectly localized on one of the two atoms in the bond. Such bonds can be understood by classical physics. The forces between the atoms are characterized by isotropic continuum electrostatic potentials. Their magnitude is in simple proportion to the charge difference.

Covalent bonds are better understood by valence bond theory or molecular orbital theory. The properties of the atoms involved can be understood using concepts such as oxidation number. The electron density within a bond is not assigned to individual atoms, but is instead delocalized between atoms. In valence bond theory, the two electrons on the two atoms are coupled together with the bond strength depending

on the overlap between them. In molecular orbital theory, the linear combination of atomic orbitals (LCAO) helps describe the delocalized molecular orbital structures and energies based on the atomic orbitals of the atoms they came from. Unlike pure ionic bonds, covalent bonds may have directed anisotropic properties. These may have their own names, such as sigma bond and pi bond.

In the general case, atoms form bonds that are intermediate between ionic and covalent, depending on the relative electronegativity of the atoms involved. This type of bond is sometimes called polar covalent.

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The chemical bond is the mutual adhesion of atoms in a molecule and the crystal lattice, as a result of the action of the force of attraction that exists between atoms. It is thanks to chemical bonds that various chemical compounds are formed; this is the nature of the chemical bond. How Chemical Bonds are Formed? The basic answer is that atoms are trying to reach the most stable (lowest-energy) state that they can. In such chemical bonds ions usually are simple, consisting of one atom of the substance. Characteristics properties of ionic bonds are the absence of saturation in it. As a result, a very different amount of oppositely charged ions can join an ion or even a whole group of ions. Chemical energy arises through reactions between atoms of different materials, whilst nuclear energy arises through reactions in the nucleus of the atoms. In chemical reactions the nuclei are not affected or changed. What the result of a chemical bond between two atoms from a simultaneous? The chemical bonding between atoms form a molecule. What is the main difference between between chemical and nuclear energy? its when an interaction holds atoms or ions togetherChemical bonds are formed when the valence electrons of one atom interact with the valence electrons of another atom. What causes chemical bonds to form between atoms? This force is the attraction between atoms. What results from a transfer of electrons between atoms? Definition: What is a Chemical Bond? Chemical bonds are forces that hold the atoms together in a molecule. They are a result of strong intramolecular interactions among the atoms of a molecule. The valence (outermost) electrons of the atoms participate in chemical bonds. When two atoms approach each other, these outer electrons start to interact. Although electrons repel each other, they are attracted to the protons within atoms. A bond between two atoms depends upon the electronegativity difference between the atoms. If the electronegativity difference is significantly high, the atoms transfer electrons to form ions and thereby form an ionic bond. If the electronegativity difference is zero or small, then the atoms combine to form covalent bonds.