3   CHEMICAL BONDS

Matter consists of atoms. Therefore, since chemistry is the study of matter, atomic theory is the logical foundation of chemistry. However, chemistry is not based on atoms alone. Chemistry first emerges when atoms combine to form molecules.

The process by which the character of atoms connecting with atoms is clarified, that is, clarifying the formation of the concept of chemical bonds, that process truly marks the path along which chemistry has developed. To understand the true nature of chemical bonds requires the support of quantum mechanics. Today quantum mechanics is actually an indispensable part of chemistry. So it is necessary that those who study chemistry consider that important new aspect of chemistry.

3.1   The theory of chemical bonds before the 20th century

(a)   Chemical affinity

The atomic theory is the premise for the concept of chemical bonds. However, the **affinity theory** favored by chemists in the 18th century may be regarded as the remote ancestor of the modern theory of chemical bonds, although it was a theory concerning chemical reactions.

The basis of affinity theory was the concept that "like attracts like". The French chemist Étienne François Geoffroy (1672-1731) made a table in which sixteen kinds of substances were listed in the order of their affinity to other substances (Fig. 3.1). The work had some historical significance as one could to some extent predict the outcome of a reaction with the aid of Fig. 3.1.

Figure 3.1 The table of chemical affinity by Geoffroy.
The symbols used are of alchemical origin. In the first row, the sixteen substances are listed. Under each of these substances, other substances are listed in their order of affinity with the substance in the first row.

Around the middle of the 19th century, chemists sought to measure chemical affinity quantitatively. The Danish chemist Hans Peter Jargen Julius Thomsen (1826-1909) and the French chemist Pierre Eugene Marcelin Berthelot (1827-1907) considered the heat generated during a chemical reaction as the measure of chemical affinity. However, some reactions are endothermic though most reactions are exothermic. It became clear that there was no simple proportionality between the heat generated during a reaction and chemical affinity.

(b)   Electrochemical dualism

Electrochemical dualism is the first rational theory of chemical bonds ever conceived, and this theory was advocated by Davy, Berzelius *et al*. in the first half of the 19th century. Berzelius reasoned as follows: atoms of various elements are either electrically charged positively or negatively in differing amounts, and charges are the driving force of formation of substances. For

instance, copper is (electro)positively charged while oxygen is (electro)negatively charged. Copper oxide formed by the combination of the two elements is slightly positive. That is why in general slightly positive metal oxide and slightly negative water react with each other to form hydroxides. The discovery that the electrolysis of alkali metal oxides yields a metal and oxygen was neatly explained by electrochemical dualism.

However, cases were found where the theory did not fit properly. According to the axiom of Berzelius, the hydrogen atom was positive and the chlorine atom was negative. Now according to Berzelius' theory, though acetic acid, CH_3COOH, was acidic, trichloroacetic acid, CCl_3COOH, should be basic. Berzelius believed that electric charge was the origin of acidity and basicity. Since exchanging hydrogen with chlorine, which have opposite charges, formed trichloroacetic acid, it should be basic. The fact was that trichloroaceteic acid was an acid even stronger than acetic acid. Electrochemical dualism thus gradually fell from favor.

(c) Valence theory

In the latter half of the 19th century, a more practical theory was proposed from the side of organic chemistry. Many organic compounds had been synthesized before that time, and their structures were determined by chemical analysis. Since it was found that many compounds had chemically similar properties (*e.g.*, in today's nomenclature, the series of carboxylic acid), chemists proposed several theories to classify and order these. According to one theory, a radical (*e.g.*, a benzoyl radical; C_7H_5O-) which was composed of several atoms was considered equivalent to an atom in inorganic compounds (Table 3.1). Another theory explained that the binding ability (chemical affinity) of a given atom attracted a definite number of other atoms.

Table 3.1 Some examples of compounds with a benzoyl radical

formula at that time	modern formula	name
$C_7H_5O \cdot H$	C_6H_5CHO	benzaldehyde
$C_7H_5O \cdot OH$	C_6H_5COOH	benzoic acid
$C_7H_5O \cdot Cl$	C_6H_5COCl	benzoyl chloride

The German chemist Stradouity Friedrich August Kekulé (1829-1896) and the British chemist Archibald Scott Couper (1831-1892) elaborated the second theory into **valence theory**. Kekulé thought that a carbon atom had four affinity units (in modern terminology, **valence**), and using these units combined with four hydrogen atoms to form CH_4. or combined with two oxygen atoms to form CO_2. He also suggested the possibility that a carbon atom could combine with another carbon atom, using one of the four valences, and each carbon atom could combine with other atoms including carbon, using the remaining three valences.

Kekulé devised a method of portraying molecules (which was called "Kekulé's sausage") as shown in Figure 3.2. At that stage, valence was merely a kind of index to indicate the ratio of atoms composing a molecule.

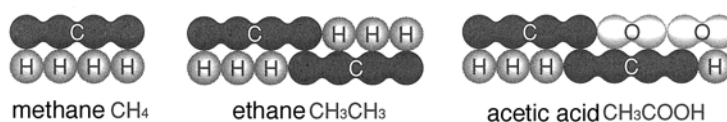methane CH_4 ethane CH_3CH_3 acetic acid CH_3COOH

Figure 3.2 Molecular structure as portrayed by Kekulé.
At this stage the concept of chemical bonds connecting atoms was not clear.

Couper formulated his theory in a similar manner, but he preceded Kekulé in that he employed the term "bond", which is used today, to indicate the binding of atoms. The fundamental concept of modern organic chemistry, that of a chain of carbon atoms, was gradually formulated. Thus the

concept of chemical bond fostered by Kekulé and Couper based on the valence theory and chemical bond was essentially identical with the modern concept of chemical bond. It must be pointed out that it was impossible, in the 19th century, to reply to the basic question: why a certain combination of two atoms will form a chemical bond while another combination of two atoms will not form a chemical bond.

3.2   The theory of chemical bond based on the Bohr theory

(a) Ionic bonds

In order to understand fully the nature of chemical bonds, the nature of the atoms forming the bond should be understood beforehand. In fact, from the beginning of the 20th century, the understanding of scientists as to the atomic structure was considerably deepened; this accelerated the development of the theory of chemical bonds.

The German chemist Albrecht Kossel (1853-1927) thought that the stability of rare gas elements (noble gases) was due to their closed shell electron configurations (*i.e.*, the electron configuration in which the outermost shell, the valence cell, was fully occupied). He attempted to extend this interpretation to other atoms. Atoms other than rare gases tend to receive electric charge (*i.e.*, electrons) from the outside or give electric charges to the outside, depending on whether the number of electrons in their outermost shell is less or more than that of the nearest rare gas in the Periodic Table. If an atom loses electrons, it will become a cation that has the same number of electrons as the nearest rare gas, while if it obtains electrons, it will become an anion that has the same number of electrons as the nearest rare gas. He concluded that the driving force of chemical bond formation was the electrostatic attraction between a cation and an anion. The chemical bond thus formed was called an **ionic bond**.

K and L shells of the sodium atom are fully occupied by electrons, but there exists only one electron in the outermost (M) shell. Hence sodium easily loses this electron to become a sodium ion Na^+ which has the same electron configuration as neon Ne $(1s^22s^22p^6)$. The electron configuration of a chlorine atom is $(1s^22s^22p^63s^23p^5)$. If a chlorine atom can acquire an electron in its M shell to make the shell fully occupied, the electron configuration will now become $(1s^22s^22p^63s^23p^6)$ which is identical with the electron configuration of argon Ar.

By that time, the crystal structure of sodium chloride had already been analyzed by X-ray crystallographic analysis, and the existence of the sodium ion and the chloride ion was established. There was no disagreement between Kossel's theory and experimental facts as far as compounds made of ions were concerned. However, this theory was also insufficient, as had been the case with electrochemical dualism, in that it failed to explain several experimental facts such as the formation of the hydrogen molecule or the absence of C^{4+} cations or C^{4-} anions.

(b) Covalent bonds

About 1916, two American chemists, Gilbert Newton Lewis (1875-1946) and Irving Langmuir (1881-1957), independently sought to explain what was inexplicable in Kossel's theory by expanding the theory to non-polar molecules. The crucial point in their theories was the sharing of electrons by two atoms as a means to make an outermost shell fully occupied when there were vacancies. Sharing of an electron pair by two atoms *i.e.*, by a **covalent bond** was a brand-new concept.

This theory was further extended to the **octet theory**. This theory explained that, in the case of rare gases (except He), the eight electrons in the valence shell are arranged as if these occupy eight corners of a cube (Fig. 3.3) while in the case of other atoms, some of the corners are not occupied by electrons. Formation of chemical bonds by sharing an electron pair is achieved by sharing an edge or a plane of the cubes.

It became possible to understand the nature of the chemical bond that forms a hydrogen molecule. However, the most fundamental question, why two hydrogen atoms combine, was not necessarily fully explained. The true nature of chemical bonds still remained unsolved.

Figure 3.3 Octet theory of Lewis/Langmuir
Models of atoms and simple molecules. A single bond is represented
by sharing an edge, and a double bond by sharing a plane.

Lewis developed a symbolism for electronic bonding to form a molecule (**Lewis structural formula** or **Lewis structure**) in the following way.

Rules for writing Lewis structures

1) All valence electrons are shown as dots around the atom.
2) A bond (in this case, a single bond) between two atoms is formed by sharing two electrons (an electron from each atom).
3) A line instead of a pair of dots is used to indicate each pair of bonding electrons.
4) Electrons which are not used for bonding remain as lone pair(s). Dots are used to represent lone pairs.
5) Except for the hydrogen atom (which will have two electrons when bonded), an atom has now in general eight electrons to satisfy the octet rule.

The following examples will show you how to write Lewis structures.

hydrogen chloride

water

carbon dioxide

(c) Coordinate bonds

By combining valence theory with the theories of ionic and covalent bonds, almost all chemical bonds known in the early 20th century could be understood to some degree. However, by the end of the 19th century, some compounds had already been reported that could not be explained by the theories of Kekulé and Couper. If the theories of Kekulé and Couper were used to interpret the structure of the luteo salt, a compound containing a metal cation and ammonia with the rational

formula of $Co(NH_3)_6Cl_3$, then a singular structure (Fig. 3.4(a)) had to be assigned.

Such a structure was not acceptable to the Swiss chemist Alfred Werner (1866-1919). He proposed that some elements including cobalt have an **auxiliary valence,** in addition to the valence defined by Kekulé and Couper, which Werner called the **main valence**. According to Werner, the cobalt atom in the luteo salt combined three chloride anions with its main valence (trivalent) and six ammonia molecules with its auxiliary valence (hexavalent) to form an octahedron with the cobalt atom in its center (Fig. 3.4(b)).

(a)   (b)
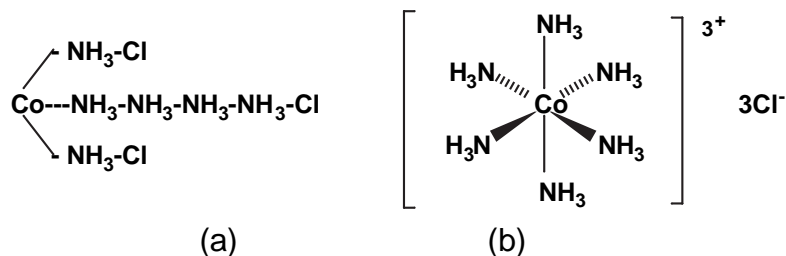
Fig. 3.4 Two proposals for the structure of luteo salt.
The wedge ◀━ starts from Co in the plane of the drawing at the narrow edge of the wedge and bonds to NH_3 above that plane. The short parallel line bonds to NH_3 are below the plane.

After many debates, the correctness of Werner's theory was generally accepted, and it was found that there were many other compounds with auxiliary valences. In these compounds, the atom (or an ion) that plays the role of cobalt was called the **central atom**, and the molecules that play the role of ammonia were called the **ligand**.

The true nature of auxiliary valence was revealed by the British chemist Nevil Vincent Sidgewick (1873-1952). He proposed a kind of covalent bond in which the shared electron pair was provided by one of the atoms, the **coordinate bond**. Thus the atom receiving an electron pair should have a vacant orbital to accommodate the electron pair. Kekulé had expressed ammonium chloride as $NH_3 \cdot HCl$. According to Sidgewick, a coordinate bond was formed between the nitrogen atom of ammonia and a proton to yield the ammonium ion NH_4^+, which in turn formed an ionic bond with the chloride ion to give ammonium chloride.
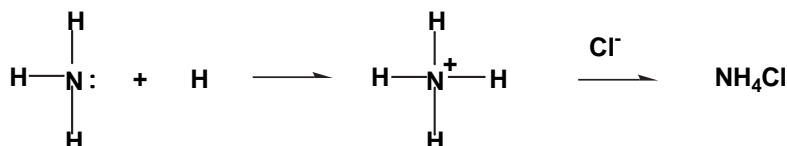
Ammonia is an **electron donor** since it donates an electron pair, while the proton is an **electron acceptor** since it accepts an electron pair in its vacant orbital.

In the case of the luteo salt, the cobalt ion has six vacant orbitals to each of which an ammonia molecule forms a coordinate bond. The reason why the cobalt ion has six vacant orbitals will be described in Ch. 5.

Sample exercise 3.1 Electron configuration of a boron compound
Write down the Lewis structural formula of trifluoroboron BF_3, and explain the reason why this compound forms an adduct (addition compound) $BF_3 \cdot NH_3$.
Solution
Boron $_5B$ is a group 13 element (*cf.* Ch. 5), and its electron configuration is $1s^2 2s^2 3p^1$, *i.e.*, it has three valence electrons. Each of the three valence electrons can make a covalent bond with one of the valence electrons of fluorine. Then, the boron atom has only six valence electrons, two electrons less than the eight associated with a stable atom. By forming a coordinate bond with the electron pair of ammonia, both boron and nitrogen can acquire the eight valence electrons associated

with a stable atom.

3.3 Quantum chemical theory of chemical bonds

(a) The method of Heitler and London

As described in Ch. 2.3, the Bohr theory, although it was revolutionarily novel, failed to explain why atoms form bonds. The Lewis-Langmuir theory of the covalent bond was essentially qualitative, and failed to give an answer to the fundamental question as to why chemical bonds are formed by two atoms, or rather, why a molecule is more stable than the two atoms which compose it.

The problem was solved by means of quantum mechanics (wave mechanics). Soon after quantum mechanics was introduced, the German physicist Walter Heitler (1904-1981) and the German/American physicist Fritz London (1900-1954) successfully explained the formation of a hydrogen molecule by approximate solution of the wave equation for a system consisting of two hydrogen atoms. In a system of two protons and two electrons, an attraction exists between each proton and each electron, and repulsion between the protons and between the electrons (Fig. 3.5(a)) They calculated the energy of the system as a function of inter-proton distance and found that the deep valley associated with the minimum energy value observed in experiments (*e.g.*, bond length) were not reproduced. They took another approach: they considered a system in which the electrons exchanged positions (Fig. 3.5(b)), and recalculated under the assumption that the two systems should contribute equally to the formation of the bond. They found that the possibility of bond formation increased, and a result very similar to experiment was obtained.
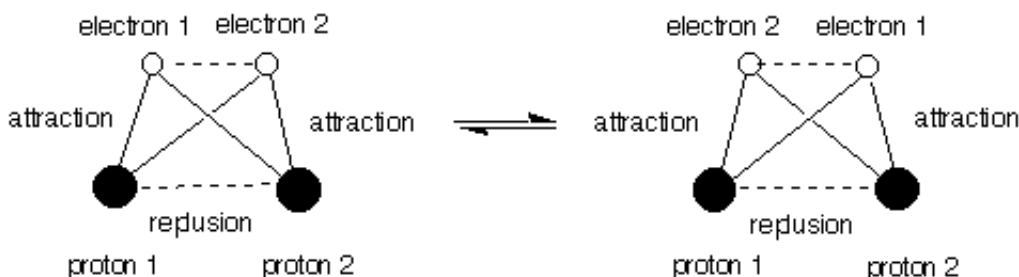

Figure 3.5 The states of a system consisting of two hydrogen atoms.
The locations of the electrons were exchanged between (a) and (b).
The system resonates between states (a) and (b).

The two states in Fig. 3.5 are called "in **resonance**". The energy difference between plots (a) and (b) is called **resonance energy**. The energy of Fig. 3.6(d) is that of the state where the spins of the two electrons are parallel. In this state, repulsion dominates, which destabilizes the bond, *i.e.*, the **antibonding** state. The method of Heitler and London was the first to explain quantitatively the nature of the covalent bond. The method had a potential to explain not only the bond that forms the hydrogen molecule, but also chemical bonds in general.
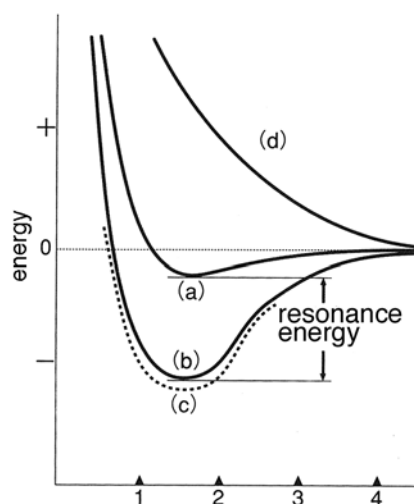
Figure 3.6 The energy of a system composed of two hydrogen atoms.
The values along the abscissa are the interatomic distances in units of the Bohr radius. (a) The energy when one state is taken into consideration. (b) The energy when resonance structures are taken into consideration (c) The experimental value. (d) The energy when the spin of the two electrons are parallel (antibonding orbital).

(b) Valence-bond approach

Let us examine the method of Heitler and London in detail. When two hydrogen atoms in the ground state are infinitely far apart, the wave function of the system is either $1s_1(1)1s_2(2)$ (which corresponds to the state where electron 1 combines with proton 1 and electron 2 combines with proton 2 as shown in Fig 3.5(a) (or $1s_1(2)1s_2(1)$ which corresponds to the state where electron 2 combines with proton 1 and electron 1 combines with proton 2 as shown in Fig 3.5(b)). When two protons approach, it becomes difficult to differentiate the two protons. In this case, the system can be approximated conveniently by a linear combination of two wave functions. Thus,

$$\Psi_+ = N_+ \left[1s_1(1)1s_2(2) + 1s_1(2)1s_2(1)\right] \tag{3.1}$$

$$\Psi_- = N_- \left[1s_1(1)1s_2(2) - 1s_1(2)1s_2(1)\right] \tag{3.2}$$

where N_+ and N_- are the constants that normalize[a] the wave function. By solving these equations, eigenvalues E^+ and E^- were obtained which corresponded to Fig. 3.6(a) and Fig. 3.6(b), respectively.

The method described above was called **valence-bond method** (the **VB method**). The premise of the VB method is that molecules can be expressed by the wavefunctions of atoms that constitute the molecule. When two electrons are shared by two atomic nuclei, and the spins of the two electrons are antiparallel, a stable bond is formed.

(c) Molecular orbital approach

The VB method was further developed by American scientists including John Clarke Slater (1900-1978) and Linus Carl Pauling (1901-1994). However, today the **molecular orbital method** (the **MO method**) is much more popular. The basic concept of the MO method can be explained conveniently by studying the simplest molecule, the hydrogen molecular ion H_2^+ (Fig. 3.7).

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[a] Recall that Ψ^2 gives the probability of finding the electron in each region of space. If you were to add up the values of this function over all points, you would have the total probability of finding the electron, which should equal unity. Orbitals are ordinarily **normalized** so as to satisfy this requirement, *i.e.*, $\int \Psi^2 \, dxdydz = 1$.

$$(-h^2/8\pi^2 m)\nabla^2\Psi + V\Psi = E\Psi \qquad (2.21)$$

Thus,

$$(-h^2/8\pi^2 m)\nabla^2\Psi + e^2/4\pi\varepsilon_0 \left[(-1/r_1) - (1/r_2) + (1/R)\right] \Psi = E\Psi \qquad (3.3)$$
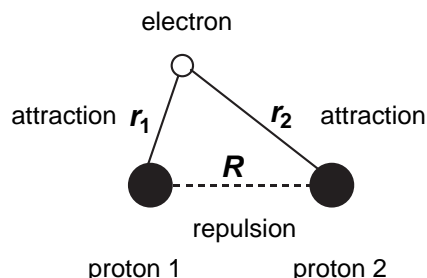

Figure 3.7 Hydrogen molecular ion.
This species is the smallest molecule, consisting of two protons and one electron.

The wave function of this system can be obtained by substituting the potential term into Eq. 2.21. When the electron is in the vicinity of nucleus 1, the influence of nucleus 2 may be negligible, and the orbital can be approximated by a 1s wave function of hydrogen around nucleus 1. Similarly, when the electron is in the vicinity of nucleus 2, the orbital can be approximated by a 1s wave function of hydrogen around nucleus 2.

Here a linear combination of the two 1s wave functions is introduced as an approximate molecular orbital of H_2. For each of the electrons 1 and 2, the following orbitals are obtained.

$$\phi_+(1) = a[1s_1(1) + 1s_2(1)]$$
$$\phi_+(2) = a[1s_1(2) + 1s_2(2)] \qquad (3.4)$$

The orbital for the whole hydrogen molecule should be the product of these two atomic orbitals. Thus,

$$\Psi_+(1, 2) = \phi_+(1)\cdot\phi_+(2) = a[1s_1(1) + 1s_2(1)] \times a[1s_1(2) + 1s_2(2)]$$
$$= a^2[1s_1(1)\,1s_1(2) + 1s_1(1)\,1s_2(2) + 1s_1(2)1s_2(1) + 1s_2(1)\,1s_2(2)] \qquad (3.5)$$

This orbital embraces the whole molecule, and is called the **molecular orbital function**, or simply, the **molecular orbital (MO)**. Accordingly, one electron orbital for an atom is called the **atomic orbital function** or simply the **atomic orbital (AO)**. The method of approximating a molecular orbital by a linear combination of atomic orbitals is called the **linear combination of atomic orbital (LCAO)** MO method.

Sample exercise 3.3 The VB method and the MO method
The difference between the VB method and the MO method lies in the extent to which they consider the electronic states of the molecule. Find this difference by comparing Eqs. 3.1 and 3.5.
Answer
Except for the constant terms, the second and third terms of Eq. 3.5 are identical with the two terms of Eq. 3.1. The electronic states described by these terms represent the neutral state of molecule because each electron belongs to a different orbital. On the other hand, the first and the fourth terms of Eq. 3.5 represent the ionic state H^+ - H^- of the molecule because two electrons occupy the same atomic orbital. Eq. 3.1 is devoid of corresponding terms. In conclusion, MO theory takes ionic states into consideration which are not considered in the VB method.

3.4 Other types of chemical bonds

(a) Metallic bonds

After the discovery of electrons, the high conductivity of metals was explained in terms of the **free electron model**, *i.e.*, the idea that metals are rich in free electrons that freely travel in the metal. That was, however, no more than a model. At the advent of quantum mechanics, around 1930, the MO theory similar to that for the hydrogen molecule was introduced into the theory of metal crystals.

Electrons in a metal crystal belong to the orbitals with discontinuous energy values, and the situation is similar with electrons circling around the atomic nucleus. However, as the number of interacting atomic orbitals increases, the gap of energy levels of the MO theory become narrower, and finally the difference becomes negligibly small. As a result many energy levels combine to form a band with a certain energy width. This theory is called the **band theory**.

The energy level of metallic magnesium offers a good example of band theory (Fig. 3.8). Electrons in the 1s, 2s and 2p orbitals are in the vicinity of the nucleus, and hence localized within the orbitals. This corresponds to the bottom of the ravine in Fig. 3.8. However, the 3s and 3p orbitals overlap and mix each other to form the MO. The MO is occupied by electrons only partly, so that these electrons are continuously accelerated by an electric field to generate electric current. Thus, magnesium is a **conductor**.

When orbitals of the valence s are fully occupied, these electrons cannot be moved by an electric field unless the electrons jump over the occupied orbitals filled with electrons to the upper vacant orbitals. This is what happens in an **insulator**.
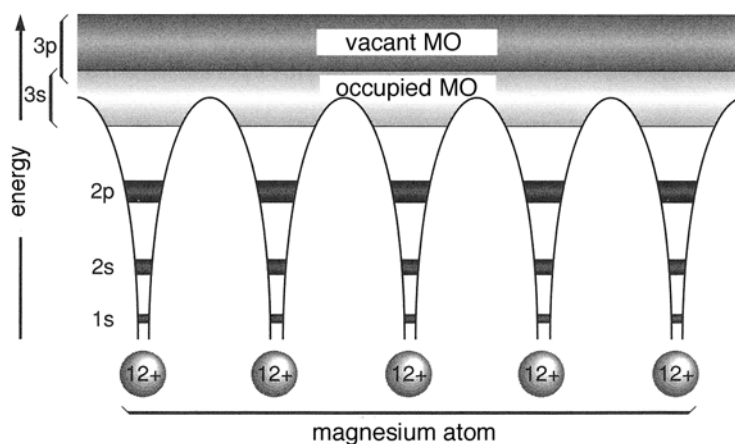

Figure 3.8 A schematic representation of MO of magnesium.
The 1s, 2s and 2p orbitals are localized, as is indicated in the figure.
However, the 3s and 3p orbitals overlap to form a band structure.

(b) Hydrogen bonds

Initially the reason why hydrogen fluoride HF has higher boiling and melting points as compared with other hydrogen halides (Fig. 3.9) was that HF exists in a polymeric form. The reason was not clear for many years. In the early 1920s, it was made clear that the polymer was formed by a bond formation between two fluorine atoms that sandwiched a hydrogen atom.

Very high boiling and melting points of water were also a matter of great concern. At the beginning of the 1930s, it was shown that two oxygen atoms of water form a bond that sandwiched hydrogen in between as was the case with HF (Fig. 3.9). It was recognized that such a type of bond is rather general and was called the **hydrogen bond**.
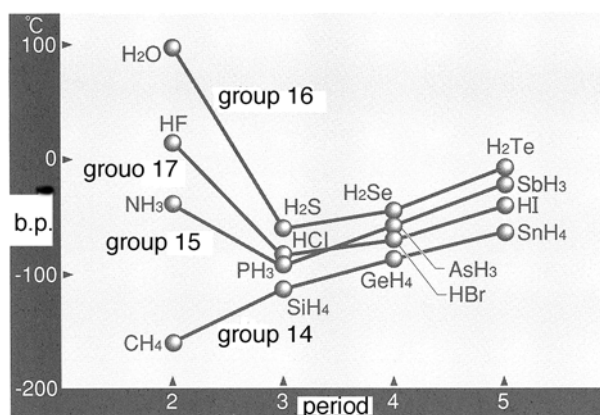
Figure 3.9 Boiling points of hydrogen compounds.
You can see that the boiling points of the 2nd row elements (N, O, F) are irregular.

Hydrogen bonds are easily formed when a hydrogen atom is bonded to such electronegative atoms as oxygen or nitrogen. The fact that some organic compounds with a hydroxy group –OH or an amino group $–NH_2$ are relatively soluble in water is due to their forming hydrogen bonds with water molecules. Dimerization of carboxylic acids such as acetic acid CH_3COOH is also well known as a good example of hydrogen bonding.
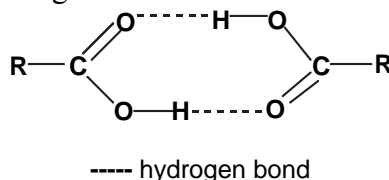
----- hydrogen bond

(c) Van der Waals force

The driving force of hydrogen bonding is the stationary uneven distribution of electric charge in the molecule, *i.e.*, the polarity of molecules (permanent dipole). Polarity of molecules is the cause of aggregations of molecules into liquid or solid states. However, nonpolar molecules such as methane CH_4, hydrogen H_2 or He (monoatomic molecule) can be liquefied, and at extremely low temperature, may also be solidified. This means that there exists some kind of aggregating force also in these molecules. Such a force is called the **intermolecular force**.

The hydrogen bond discussed above is a type of this intermolecular force. The representative intermolecular force for nonpolar molecules is the **van der Waals force**. The origin of this force is the momentary uneven distribution of electric charge (momentary dipole) caused by fluctuation in the electron cloud around the nucleus. If other conditions are equal, the more electrons a molecule has, the more easily it will polarize because the electrons are widely spread. When two clouds of electrons approach each other, dipoles are induced as the clouds polarize in such a fashion as to stabilize each other by opposing opposite charges. By van der Waals force a system will be stabilized by 1 kcal mol^{-1}. Compare this value with the stabilization achieved by formation of a chemical bond (in the order of 100 kcal mol^{-1}). Chemists are now very much interested in supramolecules which are formed by aggregation of molecules by intermolecular forces.

Exercises

3.1 Strength of an ionic bond

The energy of interaction between two electric charge Q_1 and Q_2 (both a positive or negative integer) which are separated by r (nm) is $E = 2.31 \times 10^{-19} Q_1 Q_2 / r$ (J nm). Calculate the energy of interaction in the two cases given below. (1) the interaction between Na^+ and Cl^- with $r = 0.276$ nm; (2) the interaction between Mg^{2+} and O^{2-} with $r = 0.25$ nm.

3.1 Answer

(1) $E = 2.31 \times 10^{-19}$ (+1)(-1)/(0.276) = -8.37 x 10^{-19} (J); or per mol,

E(mol) = -8.37 x 10^{-19} x 6.022 x 10^{23} J = 5.04 x 10^5 J = 504.0 kJ.

(2) $E = 2.31 \times 10^{-19}$ (+2)(-2)/(0.205) = -4.51 x 10^{-18} (J); or per mol,

E(mol) = -4.51 x 10^{-18} x 6.022 x 10^{23} J = 5.04 x 10^5 J =2710 kJ

The reason why the latter value is larger is firstly the larger charge and secondly the shorter interatomic distance for the latter.

3.2 Polarity of a bond

The magnitude of the polarity of a bond, i.e., the extent of uneven distribution in the bonding electron pair, is determined by the difference in electronegativity of the two atoms forming the bond. Arrange the following bonds in the order of increasing polarity. You can use Table 5.7 for electronegativity values.

H-H, O-H, Cl-H, S-H, F-H
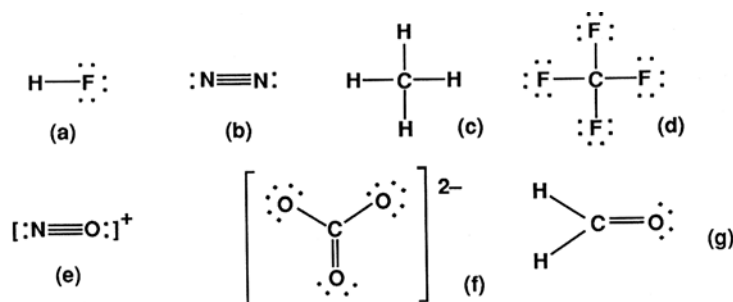
3.2 Answer

H-H < S-H < Cl-H < O-H < F-H

3.3 Lewis structural formula

Following the octet rule, write down the Lewis structural formulas for the following compounds:

(a) hydrogen fluoride HF   (b) nitrogen N_2 (c) methane CH_4 (d) carbon tetrafluoride CF_4

(e) nitrosyl cation NO^+ (f) carbonate ion CO_3^{2-} (g) acetaldehyde HCHO

3.3 Answer

3.4 A boron-nitrogen compound

Answer the following questions.

(1) Write down the electron configuration of boron in the ground state.

(2) Draw the Lewis structural formula for BF_3.

(3) Draw the Lewis structural formula for NH_3.

(4) The reaction between NH_3 and BF_3 gives an addition compound. Explain why this reaction takes place, and suggest the structure of the addition compound.

3.4 Answer

(1) $1s^2 2s^2 3p^1$

(2), (3)   (4)