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THE DEVELOPMENT OF VIEWS REGARDING THE NATURE OF CHEMICAL FORCES[1]

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AT the beginning of my lectures before the department of chemistry of Cornell University, I wish to express my great pleasure at having the opportunity to spend a few months in this wonderful laboratory, beautifully situated on a knoll of the picturesque campus, and to come into close personal contact with my distinguished colleagues as well as with the students. I am sure to learn many things from the splendid organization of this department which will be of great use in my own institute. I thank Professor Dennis heartily for the honor of his invitation to the George Fisher Baker Non-resident Lectureship and for the opportunities he thus affords me.

It is a pleasant custom that the holder of this lectureship begins his activities with a general address for a wider circle of listeners than will attend his later lectures in his special field. Several of my predecessors have chosen for this introductory lecture subjects that were more or less far removed from the particular lines of their own research. If you will permit me, I will join the other group and discuss a problem which lies in my own field, for I regard the problem of chemical forces as one that may be of interest not alone to the specialist but to the general public as well. These forces affect deeply not only the surrounding, inanimate world but also the life processes within our own organism, and I shall endeavor to present to you the problem of their true nature.

This question has greatly interested theorists for many centuries. Let me briefly recount at least a few of the steps in the older development before I pass on to the discussion of the views that are held at the present time. There is a characteristic trend that has been common to most of the various theories which have arisen. Chemistry, as you know, serves as a

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[1] Introductory public lecture.

basis of many other sciences, such as mineralogy, crystallography and biology, and chemistry in turn rests in many respects upon the results obtained in physics. And we shall see that the views concerning the nature of chemical forces in the passage of time often have been modified and shaped by the physical theories of the moment.

In the seventeenth century there was great development in the field of mechanics. It is necessary to mention only the names of Galileo, Kepler and Newton to make clear the powerful influence that mechanics exerted upon the concepts of that period. Even a century later, explanations of phenomena were expected to be based upon mechanical processes. As early as the seventeenth century the French chemist Lemery, in an attempt to explain the action of an acid on a metal, sought for a mechanism upon which its chemism might be based.

The sensation of taste that is produced by acids can be described as biting or prickling, and it was quite natural, from the point of view of a crude mechanical concept, that when a metal is dissolved by an acid the small particles of the acid were assumed to possess a pointed form which would bore into the cavities of the metal. Naturally such a concept can not be recognized as a sound scientific theory. In general the mechanical ideas have not proved to be fruitful in explaining chemical phenomena. To enable the particles of an acid to act upon a piece of metal, the particles of acid dissolved in water must by virtue of their motion in the liquid come into the immediate neighborhood of the metal. Thus far, the process can be clearly conceived as mechanical, but the real chemism begins only when the reacting particles of the acid and metal collide with one another, and at this point the bridge to the usual mechanical explanation is lacking.

It is interesting to note, however, that to-day mechanics is again beginning to come into closer relationship with what is essentially chemical action. This newer view is based not upon the classical mechanics of the seventeenth and eighteenth centuries, but on the refined, modern theory of wave mechanics or quantum mechanics associated with the names of de Broglie, Heisenberg and Schrödinger. We must, however, ascend through several stages to pass from the mechanical chemistry of a Lemery to this latest view.

When we leave the field of mechanics we enter at once another branch of physics, electricity, which we must consider if we are to understand the views advanced by the great Swedish chemist Berzelius in the early part of the nineteenth century. By that time two fields of knowledge had been developed that are of permanent value for our problem. There was brought forward a more exact conception of the building up of substances from small particles, an idea familiar to the Greeks. Dalton developed the atomic theory which required the existence of as many different atoms as there are non-decomposable elements, the smallest particle of a chemical compound, the molecule, being composed of one or more elemental atoms. The question at once arose: What forces hold these atoms together in the molecule?

The answer to this question appeared to be brought nearer because, in this period, it was found possible to overcome these attracting forces by means of a new and powerful agent, the electric current. At the end of the eighteenth and the beginning of the nineteenth centuries there appeared the epoch-making discoveries of Galvani and Volta as to the manner of producing an electric current. With the aid of the current, Davy was able to bring about sensational chemical reactions such as the separation of caustic soda into metallic sodium and oxygen. This was a decomposition of a chemical compound into its elements, a splitting of a molecule into atoms with the help of electricity.

Analysis of this phenomenon of electrolysis seemed to indicate a close relationship between electricity and the various kinds of atoms. Distinction had long been made between positive and negative electricity, and the new discoveries by Davy and others indicated that many elements, such as the metals and hydrogen, behaved during electrolysis as if their atoms carried positive charges, while the atoms of chlorine, oxygen and similar elements seemed to be negatively charged. But positive and negative electricity exert an attractive force upon each other, while charges of the same sign repel each other. This brought Berzelius to the idea that the chemical force which binds the oxygen atom either to the sodium atom in an alkali or to the hydrogen atom in the water molecule is of an electrical nature and is dependent simply upon the attraction of oppositely charged atoms as represented by the following slightly modernized formulas:

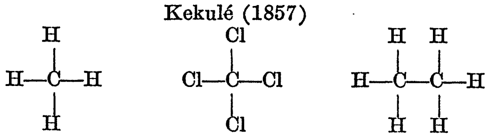
Berzelius (1812)

$$\overset{+}{Na_2}\overset{-}{O},\ \overset{+}{Na}\overset{-}{Cl},\ \overset{+}{H_2}\overset{-}{O}$$

This conception of the electrostatic nature of chemical forces Berzelius sought to extend to all chemical substances. He held that each atom contained an excess of either positive or negative electricity, and that because of this fact an atom is able to combine with oppositely charged atoms. Indeed, this electrostatic theory of Berzelius seemed to be satisfactory in explaining the behavior of many substances, and it was used by chemists as long as their researches were limited chiefly to inorganic compounds. But about the middle of the nineteenth century a new branch of chemistry blossomed into life—organic chemistry. This concerned itself with the substances found in animate nature, compounds whose most important constituent is carbon. Among the simplest of such compounds we have methane (CH_4) and carbon tetrachloride (CCl_4). Both these substances are stable compounds in which the atoms are held together by strong forces. If now one endeavors to explain these attractive forces on the basis of the theory of Berzelius, contradiction concerning the charge on the carbon atom immediately appears. According to the evidence furnished by the electrolysis of inorganic compounds, hydrogen is positively charged, from which it follows that the carbon atom in CH_4 should carry a negative charge. Yet in CCl_4, since chlorine is shown by electrolytic experiment to be negatively charged, the carbon atom should have a positive charge. Such contradictions, which arose in many analogous cases, led to the conviction that the electrostatic theory of Berzelius was untenable, and therefore chemists gradually abandoned it. We shall see later that this theory did nevertheless contain a germ of truth.

In order properly to follow the historical development of the subject, however, we must now leave the field of physics and turn our attention to a theory that forgoes definite assertion concerning the physical nature of chemical forces and contents itself with a formal representation of the union of atoms in molecules. I refer to the valence theory advanced in 1857 by the German chemist Kekulé. Kekulé represented the union of the atoms in the molecule by means of lines, the maximum number of other atoms with which a given atom may combine being represented by the number of lines proceeding from it. This number is termed the valence of the atom. The carbon atom has a valence of four, the atom of hydrogen and that of chlorine in carbon tetrachloride each has a valence of one.

Kekulé (1857)



$$\begin{array}{c}H\\|\\H-C-H\\|\\H\end{array}\qquad\begin{array}{c}Cl\\|\\Cl-C-Cl\\|\\Cl\end{array}\qquad\begin{array}{c}H\quad H\\|\quad\;|\\H-C-C-H\\|\quad\;|\\H\quad H\end{array}$$

Methane Carbon tetrachloride Ethane

It is a strikingly primitive picture that is used here to represent the chemical forces, and yet it confirms in most convincing manner the truth of Goethe's words, *In der Beschränkung zeigt sich erst der Meister,* for the scheme of Kekulé has proved to be extremely fruitful in chemical research up to the present day, and in particular in the field of the chemistry of the compounds of carbon. We find in almost every case that the carbon atom possesses four of such valence forces, four units of chemical attraction each of which is able to seize and hold an atom. The constancy of this number 4 in most of the carbon compounds constitutes the great, although purely systematic, value of Kekulé's suggestion.

From the physical view-point Kekulé's theory stands in some contradiction to that of Berzelius. On the basis of the Berzelius theory one can also draw lines between the atoms that he regarded as oppositely charged and as electrostatically bound. Such connecting lines would then have a positive and a negative end. One positive atom can be replaced only by a positive atom. The link would be polar. According to Kekulé, on the other hand, atoms of widely varying nature can be attached to the outer ends of the valence forces emanating from the atoms of carbon, and this stands in agreement with the facts of organic chemistry. Thus the carbon atom may attach itself to other carbon atoms, as is shown in the above formula of ethane. The linkage is here non-polar. Specially convincing illustrations of the existence of such non-polar linkages are found in the molecules of certain elements: the molecule of ordinary gaseous hydrogen (H—H) and also the molecule of chlorine (Cl—Cl) are each composed of two atoms. In such molecules there is no reason to assume that one atom carries a positive and the other a negative charge. The above successful application of Kekulé's scheme to inorganic molecules also seemed to justify its use for representing such compounds as sodium chloride, for which the theory of Berzelius seemed to be quite satisfactory. Until recently this has been written as Na–Cl without endeavoring to bring out the opposite polar character of sodium and chlorine as revealed by electrolysis. As may so often be observed in the history of the development of scientific theories, a leap from one extreme to the other here took place. Both theories attempted to reach the same goal and to explain the linkages of the atoms in all substances in a uniform manner. Further developments, however, showed that this could not be achieved in so simple a way.

In the last decade of the past century there appeared a new branch of chemistry, physical chemistry, associated with the older branches, inorganic and organic chemistry. The difference between physical chemistry and these two older divisions does not lie in the nature of the substances that are employed in the sphere of its investigations, for it concerns itself both with inorganic and organic compounds. However, the method of research used in physical chemistry, which was indeed only partially novel, lays particular emphasis upon the study of chemical compounds and reactions from the standpoint of physics. The electric current was again invoked for the investigation of the behavior of substances but not alone for the purpose of effecting chemical decomposition. The question as to whether various substances when dissolved in water are capable of conducting the electric current was very carefully studied and striking differences were found to exist. Aqueous solutions of compounds such as ordinary salt conduct the electric current very well. In order that conduction of the current may take place, however, charged particles must be present. It was Arrhenius who, in the year 1887, brought forward the theory, at that time regarded as extremely bold, that in such a solution of salt the molecule NaCl decomposes into a positively charged sodium atom and a negatively charged chlorine atom, the so-called ions. This concept of Arrhenius may to-day be regarded as fully verified. From this it would follow that when a solution of salt is evaporated and the water is removed the ions are brought closer together from the relatively greater distances that existed between them in solution, and the electrostatic forces between them must become active in some manner, since according to Coulomb's law the intensity of the electrostatic force is inversely proportional to the square of the distance between the charges. On the other hand, a substance like H_2 in aqueous solution does not conduct the current to any noticeable extent and consequently there is no reason for assuming the existence of ions of hydrogen in this solution.

These facts have led us since the beginning of this century (Abegg, Nernst) to distinguish between two different kinds of chemical linkages: on the one hand the polar linkage, also termed hetero-polar or dualistic linkage, which is characteristic of salts, and on the other hand the non-polar, homo-polar or unitary linkage, which is characteristic of the molecules of elements and of many organic compounds. And thus we can understand why the dualistic theory of Berzelius could maintain itself only so long as the investigation of organic compounds was not the center of interest. However, the unitary theory of Kekulé which replaced that of Berzelius also does not satisfactorily explain all the phenomena with which we have become acquainted. Both theories are entitled to exist side by side.

The theory of Berzelius in its original form was very weak in one respect; it assumed, for example, that the sodium atom and the chlorine atom always carry an excess either of positive or negative electricity. This, however, is not in agreement with the facts, for metallic sodium as well as its free atom in sodium vapor under normal conditions is electrically neutral. The same holds true for chlorine. The question arises, therefore, as to the manner in which the positively charged sodium ion and the negatively charged chlorine ion present in the solution are formed from these neutral atoms.

A satisfactory answer to this question was possible only after we had obtained a deeper insight into the nature of the atom itself and had acquired the knowledge that the atoms of the elements are not really the indivisible building-stones from which the material world is constructed. Physical research which began with the investigation of cathode rays and the rays from radioactive substances, and with which the names of Lenard, J. J. Thomson and Rutherford are intimately connected, revealed that the atoms of all elements are themselves composite structures and are in the last analysis made up of positive and negative electricity. These electrical charges have an atomistic structure; the magnitude of the smallest charge, the elementary electrical quantum, was measured by your countryman, Millikan, in his brilliant investigations. The charges of these elementary quantities of positive and negative electricity are equal and differ only in sign.

According to the Rutherford-Bohr theories of atomic structure, the atoms of all elements are composed of these atoms of positive and negative electricity. The simplest and lightest element, hydrogen, contains in its neutral condition one elementary quantum of positive electricity and one of negative electricity. The two do not, however, play precisely the same rôle, because their masses are not the same. The elementary negative charge, termed an electron, possesses a mass only 1/1830 that of the total mass of a hydrogen atom. The remainder of the mass lies in the positive charge. The lighter electron revolves around the heavy, positively charged nucleus of the hydrogen atom, the proton, and because of this fact the two attracting charges do not fall together but remain at a definite distance from each other, as do the sun and its planets.

When we consider the heavier atoms we find that they also contain a positively charged nucleus, whose charge, however, varies from element to element and always is a whole-numbered multiple of the elementary quantity, *e.g.*, 11 for the sodium atom, 17 for that of chlorine. These nuclei have dimensions which are very small compared to the dimensions of the whole atom even in the case of the heaviest atoms. The negative charge in the heavier atoms is divided among a corresponding number of negative electrons, and this number in the neutral state is therefore 11 for sodium and 17 for chlorine. All these electrons move in more or less complicated paths around the nucleus.

How now are the ions formed from such neutral atoms? Since, as has been mentioned, a positive elementary charge carries practically the whole mass of a hydrogen atom and such positive charges are localized in the nucleus of the heavier atoms, the removal of one or more positive charges from an atom would deeply affect its internal constitution and would mean a conversion of one element into another such as we have actually observed in the radioactive processes. The chemical processes, as well as the formation of ions from neutral atoms, take place only in the outer sphere of the atom, in the electron shells. A positive ion is formed when one or more electrons are removed from a neutral atom, for then the positive charge of the nucleus will exceed the total negative charge by the number of negative electrons removed. Negatively charged ions are obtained by attaching one or more electrons to a neutral atom. The process of forming a molecule of the sodium chloride type from neutral atoms is therefore most simply understood by assuming that an electron is removed from a sodium atom and attached to a chlorine atom; the electrostatic attraction of the resulting positive and negative ions then unites them into a molecule or, if they are present in larger number, into a crystal of the salt.[2] This is the essential content of the electrostatic theory of polar linkage outlined by J. J. Thomson in 1904 and especially developed by Kossel in 1916. It represents the modernization of the Berzelius theory.

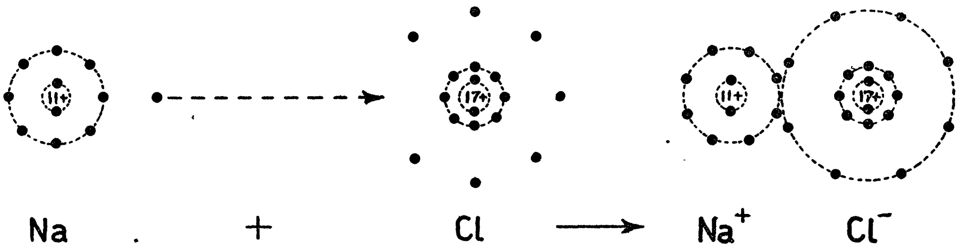


FIG. 1

As far as the other type of linkage is concerned, we are indebted to the idea of Niels Bohr, which shows how it can be brought about by the electrical interaction of positive and negative charges in the atom without the necessity of assuming oppositely charged ions in such molecules and crystals. Let us consider the simplest non-polar molecule, that of hydrogen. It contains two atoms of hydrogen, and accordingly, two positive nuclei and two negative electrons. The mechanism for its constitution (see Fig. 2a) suggested by Bohr, was as follows. The two electrons move along a path whose plane is perpendicular to the line joining the two nuclei. The positive nuclei are held together by the negative electrons, which draw them toward the middle even though the positive nuclei mutually repel each other.

This model has not proved satisfactory from the quantitative point of view, and it is now assumed that the true structure of the hydrogen molecule approximates more closely that shown in a model designed by Heitler and London (1927) on the basis of the wave mechanics referred to at the beginning of this lecture. According to this newer view, the electrons perform such complicated motions within the atoms and the molecules that, in the course of time, an electron passes through every point of the space occupied by the atom or molecule, although the probability that an electron will arrive at any given point depends upon the distance of the point from the nuclei. Therefore, it is only possible to indicate a time average of the spatial distribution of the negative electricity. From the results obtained in this direction, it is particularly important for the discussion that is to follow, that this distribution of the negative electricity with reference to the nuclei of the two linked hydrogen atoms be quite identical, and in this respect the wave mechanics model and that in Fig. 2a are in perfect agreement. In both models also each of the electrons is, in fact, as often at a given distance from the one nucleus as it is from the other.

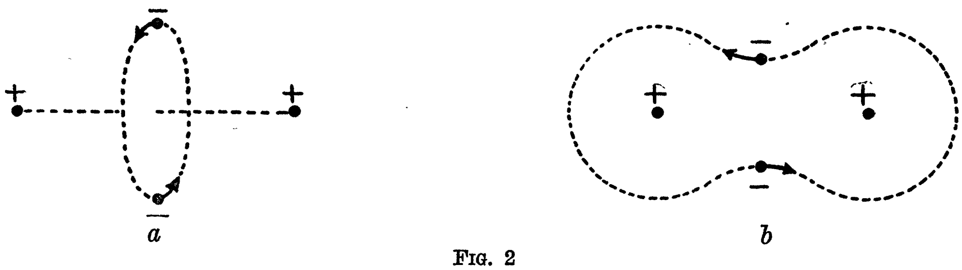


FIG. 2

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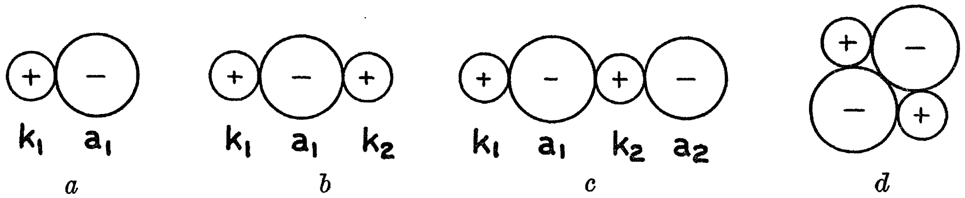
[2] Of great interest in this connection is the fact that both ions contain eight electrons in the outermost shell, like the rare gases neon and argon.

The relation between nuclei and electrons in the new model may be represented in a simplified way, sufficient for our purpose, by considering that the two electrons describe a path around the two nuclei, as indicated in Fig. 2b, and as has been assumed by a number of investigators in the past, particularly by Knorr (1923). The characteristic point common to all models proposed is, then: the linkage is effected by two electrons which are shared by both atoms in an identical manner. Here there is actually no polar contrast between the two bound atoms, and each contributes one electron to the linkage.



a *b* *c* *d*

FIG. 3

As has been shown very successfully by your countrymen, Lewis and Langmuir, this principle is also applicable to the molecules of other elements and to many other non-polar compounds. The general principle is, as a comparison of the following formulas with those of Kekulé shows, that two electrons held in common by the two atoms are associated with each one of Kekulé's valence lines—hence the term covalent linkage in contrast to ionic linkage.

G. N. Lewis (1916)

$$\begin{array}{ccc}
\mathrm{H} & \mathrm{Cl} & \mathrm{H\ \ H} \\
\cdot\cdot & \cdot\cdot & \cdot\cdot\ \ \cdot\cdot \\
\mathrm{H:C:H} & \mathrm{Cl:C:Cl} & \mathrm{H:C:C:H} \\
\cdot\cdot & \cdot\cdot & \cdot\cdot\ \ \cdot\cdot \\
\mathrm{H} & \mathrm{Cl} & \mathrm{H\ \ H} \\
\text{Methane} & \text{Carbon tetrachloride} & \text{Ethane}
\end{array}$$

The chemical and physical behavior of both types of compounds stands in good agreement with the above conceptions of their structure. We will mention here three properties: the type of lattice in which the substance crystallizes; the ease with which it passes from the solid or liquid to the gaseous state, *i.e.,* its volatility and the ability to form ions in aqueous solution and therefore to conduct the electric current. In this respect extremely large differences may be mentioned between the typical representatives of the two classes of compounds (see Table I). Rocksalt can be vaporized under ordinary pressure only at 1440°. Hydrogen, however, is a gas at ordinary temperatures and must be cooled to $-253°$ to liquefy it. In solution the former decomposes to a large extent into ions, whereas the latter remains as the molecule H_2 unchanged. The tendency to go over from the solid or liquid states to the vapor or gaseous state and *vice versa* is dependent upon the strength of the forces between the molecules. The attractive forces between the molecules of NaCl must, obviously, be much greater than those between the molecules of H_2. Indeed, even though the molecule NaCl is electrically neutral as a whole, it exerts, as the schematic representation in Fig. 3 indicates, a considerable electrical effect in its immediate neighborhood, because in this molecule the centers of gravity of the positive and negative charges, two poles (Fig. 3a), are separated. If one brings a positive ion k_2 to such a structure, which one calls a dipole, it will be attracted to the negative ion in the position which is indicated in Fig. 3b. The attractive force between the opposite charges a_1 and k_2 is greater than the repulsive force between the like charges of k_1 and k_2, because the distance between the former is but one half that between the latter. And if, in turn, one now adds the negative ion a_2 (Fig. 3c) it will, for the same reason, be attracted by the positive ion k_2 more strongly than it will be repelled from the negative ion a_1. Furthermore, there is a definite, although weak, additional attractive force exerted upon a_2 by the positive ion k_1. What we have before us is equivalent to the adhesion of two such neutral molecules, k_1a_1 and k_2a_2, as a result of electrostatic forces which are completely comparable with those which effect the cohesion of the ions within the molecule. Something entirely analogous holds for the action in the other two directions of space (Fig. 3d), and this explains, as shown by Kossel, the great tendency of the molecules of such salts to group themselves in large complexes and to form liquids and crystals. The validity of this view is attested by the recent determinations of the position of atoms and ions within such crystals by means of X-ray analysis. In Fig. 4 is portrayed the crystal structure of rocksalt. The individual points show the positions of alternate, positively charged sodium ions and negatively charged chlorine ions in a minute part of such a crystal. Each ion within the crystal is surrounded by six oppositely charged ions, all at exactly the same distance from the former. The electrical attractive forces existing between neighboring ions are equal in all three directions of space, so that it is not possible to regard any two ions as forming a molecule. Such a lattice therefore is called a coordination lattice. Similar relationships hold for many other salts. (See footnote 3.)

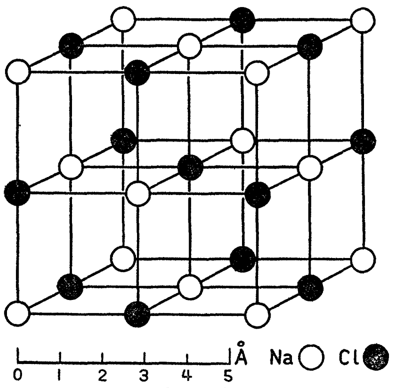


FIG. 4

The circumstances are entirely different in the case of H_2. Here the centers of gravity of the positive and negative charges coincide, so that as long as no shift of the charges takes place in the molecule, no electrical forces can be exerted on the outside. Hence, there is but a slight tendency for the H_2 molecules to combine with one another to form a liquid or solid. And there can be no doubt that the crystal structure of solid hydrogen is quite different from that of rock-salt. Corresponding to the very strong forces between two hydrogen atoms within the molecule and the very weak forces between distinct molecules, the distances between the atoms belonging to different molecules must be considerably greater than between those within one molecule, as indicated in the diagram (Fig. 5). Here we have individual molecules in the crystal lattice, and therefore we term it a molecule lattice.

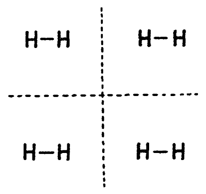


FIG. 5

Finally, the great difference between NaCl and H_2 with respect to the third property as given in Table I, *viz.*, the degree of their ionizations in aqueous solution, is readily understood without special discussion. Indeed, in the case of NaCl, the molecule and the crystal already contain ions and they need only to be separated by the action of the forces between the water and the ions, while in the case of such non-polar substances as H_2 energy first must be expended in order to form ions by shifting electrons to one of the atoms.

In such cases as that of sodium chloride and hydrogen, the relations are obviously quite clear. There is here not the least doubt concerning the group to which we must assign the substance, the ionic or the covalent non-polar type of compounds or linkages.

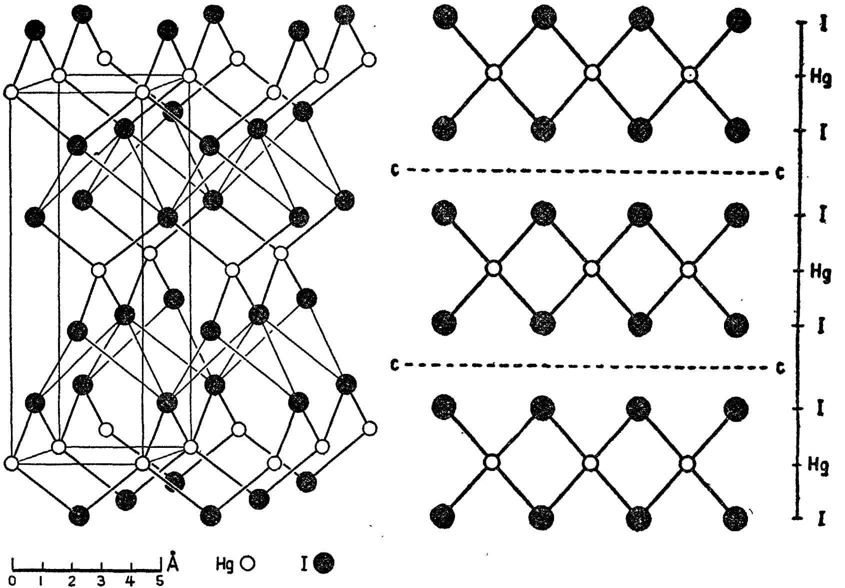
TABLE I

	Ionic linkage	Non-polar linkage	Transition type
	$\overset{+}{Na}\,\overset{-}{Cl}$	H : H	I Hg I
Type of lattice	Ionic coordination lattice	Molecule lattice	Layer lattice
Boiling-point at 760 mm	1440°	-253°	351°
Degree of ionization in aqueous solution	high	zero	very low

It is immaterial whether we choose as our criterion the volatility, the type of lattice or the ability to dissociate in solution. The same holds for many other substances, and there is a wide-spread tendency to classify all substances under one or the other of these two types of linkages. However, there is a very large number of compounds for which this classification is so indefinitely indicated that it is impossible to apply it with any certainty, and naturally one finds differences of opinion regarding the type of structure to which they really belong.

An example is afforded by mercuric iodide, whose behavior with reference to volatility, crystal structure and ionization in solution is intermediate between that of hydrogen and that of sodium chloride. Mercuric iodide is solid at ordinary temperatures but boils at 351°, and may be vaporized with relative ease similarly to the closely related mercuric chloride, popularly termed corrosive sublimate. Both compounds have a low solubility in water as compared with sodium chloride and, what is here of particular importance, their degree of ionization is measurable but very small.

Of particular interest is the structure of the crystal lattice of mercuric iodide (see Fig. 6a). The characteristics of this lattice that are essential to our problem are emphasized in Fig. 6b which is a simplified presentation in one plane. It is seen that the two kinds of atoms are arranged in alternating layers —hence the type name proposed for it by Hund, "layer lattice." Just as in the lattice of sodium chloride, and in contrast to a molecule lattice, no individual molecule HgI_2 can be differentiated here. Each Hg is linked in an identical way with four, not with two I's; and each I is linked, not with one but with two Hg's. In another respect, however, this layer lattice recalls a molecule lattice more than it does an ionic coordination lattice. In the latter each ion has attracted for its immediate neighbors, on every



a *b*

FIG. 6

side, only ions of opposite charge.[3] In the lattice of mercuric iodide, on the other hand, the layer sequence I, Hg, I is followed by I again, the intervals between two neighboring I-layers being greater than that between an Hg- and its neighboring I-layer. Obviously the attractive forces acting between two neighboring, similar layers must be different, and in fact they are weaker than those acting between two unlike layers, just as the forces between the molecules of H_2 are weaker than the forces within H_2. These weaker forces in the lattice cause the cleavability along the planes c-c (Fig. 6b) and, to some extent, the volatility of such crystals.

It is thus evident that none of the three properties mentioned in Table I permits one definitely to decide whether mercuric iodide should be classified as a polar or a non-polar compound. It behaves in an intermediate manner.

The difficulty of classification here encountered is, in the speaker's opinion, inherent in the method of classification itself. The Berzelius theory has failed by reason of its attempt to apply to all substances the same structural principle. It is, moreover, impossible, as we have seen, to understand satisfactorily the behavior of all substances from the view-point of the unitary concept. But classification into two sharply divided groups also is too narrow to explain satisfactorily all the facts. The distinction between the two types of structure mentioned probably will retain its fundamental significance for all time, but it is the speaker's conviction that one should regard them rather as limiting types between which it is possible to have a series of transition forms into which indeed many known substances can be fitted. The limiting case of non-polar linkage is found in the quite symmetrically built molecules: hydrogen, H_2, or chlorine, Cl_2. The other limiting case, the ideal ionic linkage, would be obtained if free positive and negative ions, formed by complete transfer of one or more electrons from one kind of atom to another—as from sodium to chlorine—could be associated with each other as rigid spheres without mutual influence. Real molecules and crystals approximate this limiting case only to a greater or lesser degree, for when two such ions come together a certain change always occurs, a deformation of the ions themselves.

The existence of such deformations can clearly be demonstrated in the case of the formation of mercuric iodide from its ions Hg^{++} and I^-. Aqueous solutions of mercuric nitrate and potassium iodide, containing these ions, are colorless. When these solutions are brought together the ions combine, yielding the intensely red mercuric iodide. This change of color indicates an alteration in the properties of the ions, namely, in the state of their electronic systems. Closer considerations show that in most cases where ions combine to form molecules and crystals it is specially the electronic system of the anion that is changed by the attractive force of the positive charge of the cation. The electronic system of the anion is drawn toward the cation, and thus the transfer of the negative electricity from the one kind of atoms

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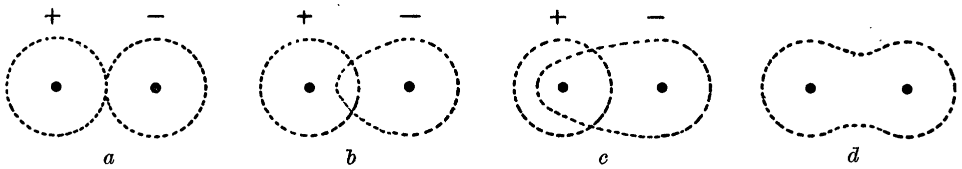
[3] This holds for any typical ionic lattice, not only for compounds of monovalent elements, such as NaCl (six neighbors each), CsCl (eight neighbors), but also, *e.g.*, for the lattice of fluorite (CaF_2), where each Ca^{++} has eight F^- and each F^- has four Ca^{++} as its neighbors.

to the other during the formation of the ions is reversed to some extent. This deformation (polarization) of the anion indicated in Fig. 7b and c thus leads to a diminution in the polarity of the compound. The degree of this deformation varies greatly in different cases and is, *e.g.,* in Fig. 7, greater in c than in b. It is therefore possible to conceive of an extended series of transition cases between the two limiting types of chemical linkage: one in which the ions are to be considered as rigid spheres which undergo practically no change when combining to form molecules (Fig. 7a), the other in which the electrons of the anions are so strongly drawn toward the cation as to become symmetrically disposed about the two nuclei (Fig. 7d). This latter case would be realized if a negatively charged hydrogen ion, H^-, composed of one proton and two electrons were to combine with a positive hydrogen ion, H^+, consisting of a single proton. This would yield the perfectly symmetrical hydrogen molecule, H_2, in which the two electrons are shared quite equally by both nuclei.



a b c d

FIG. 7

In general, as was pointed out by the speaker and confirmed by Grimm, extensive experimental material shows that the degree of deformation of the anions increases and the character of the linkage in a molecule or a lattice departs further from the ideal ionic linkage and approaches more closely the non-polar type, the smaller the cation, the higher its charge and the greater the deformability of the anion. Other things remaining the same, in those compounds whose cations lack the character of rare gases the degree of deformation of the anion is greater and the compounds approach more closely the non-polar limit than in the case of those having cations of the rare-gas type. Mercuric iodide stands much farther from the ideal ionic linkage than does sodium chloride, because I^- is much more easily deformable than Cl^-, as refractometric data show, and because Hg^{++} has twice the charge of Na^+. The most important factor in this case is that Na^+ has the structure of neon (eight electrons in the outermost shell, Fig. 1) but Hg^{++} has no rare-gas structure (eighteen outermost electrons). To the intense deformation of the iodine ion by the mercuric ion is due the transition type of mercuric iodide.

Though the classification loses much in simplicity through the presence of these transition types, by taking account of them we attain a single system satisfactory for widely different substances which the dualistic and unitary theories attempted in vain.

I realize that the subject which I have discussed is a difficult one to present to a general audience, and I can only hope that I may have succeeded in giving you some idea of how scientists have attacked this problem of the nature of chemical forces, and of the conclusions which they have reached.

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SCIENTIFIC EVENTS

HISTORIC APPARATUS[1]

THE Institute of Physics a short time ago appointed a committee to inquire into the possibility of drawing up a catalogue of apparatus which had been used in the course of investigations leading to important discoveries. At the request of the committee the institute is now making a general appeal to all who have charge of historic apparatus. The appeal, which is signed by Dr. W. H. Eccles, Sir F. W. Dyson, Sir W. H. Bragg, Sir C. A. Parsons, Sir J. J. Thomson, Sir R. T. Glazebrook and Sir H. G. Lyons, is made with the object of obtaining information as to the existence of apparatus and the researches it was used in, so that students of the history of science may be able to visit, identify and study it. It is also suggested that the committee could possibly assist the owners, if they so desire, to place the apparatus where it could be per-

[1] From *Nature.*