The appearance of this vibration only above a transition in the cyclohexane derivatives would be explained if at low temperatures the *trans* form (Fig. 12) of the cyclohexane ring were the more stable. With rising temperature vibration of increasing amplitude might be expected to attain sufficient energy to carry some carbon over the potential hump between it and the pliable *cis* position. The difficulty of exciting atomic group vibrations should then decrease abruptly by a large factor, in line with the reduction of the Einstein parameter Θ. Absorption of energy and increased heat capacity would result. In a crystal the effect might spread from one molecule to another very rapidly, in the manner suggested by Pauling[3] for rotational transitions, because of the relaxation of crystal forces which might be expected to accompany a great reduction of the rigidity of the molecule, and an expansion of its "average volume." The same reduction of crystal forces should greatly facilitate the molecular rotation which is also believed to rise at the transition temperature.

An alternative possibility is that, even if the *cis* form were stable at the lowest temperatures, dipole forces, London forces and steric hindrance might prevent pliable rotation and vibration until a transition temperature was reached.

The accepted model of the *d*-camphor molecule is not pliable in the same sense as the cyclohexane ring, and hence gives very little indication of being subject to the type of extremely low frequency vibration required by our hypothesis. It is possible that the mechanical model does not reveal all types of vibration actually present in the molecule. However, the model does show that camphor contains a cyclohexane *cis* ring, whose tendency to pliable vibration would be elastically opposed only by the two bonds in the "bridge" of the molecule.

Summary

1. A rise of about 18 cal./°C./mole in the specific heat of *d*-camphor at the transition at $-30°$ is observed and is explained by assuming that unusually energetic intramolecular vibration arises at the transition temperature.

2. The rise of molecular rotation at the transition which is evidenced by the dielectric data of Morgan and Yager is qualitatively explained in terms of Pauling's theory of molecular rotation in crystals, by assuming that the forces opposing this rotation are greatly diminished by a transition which reduces the rigidity of the molecules.

3. The behavior of total polarization of *d*-camphor in dilute solution is explained by assuming that at room temperature dipole moment is smaller and atomic polarization much larger than formerly supposed, and that the latter polarization begins to decline with temperature below $-10°$ when the transition to the more rigid molecules begins.

4. Transitions very similar to that occurring in *d*-camphor are observed in crystals of derivatives of cyclohexane, whose carbon ring is supposed to be very pliable in the *cis* form and relatively rigid in the *trans* form. These transitions appear to be due to transformations from the rigid to the pliable forms of the molecules of which the crystals are composed.

NEW YORK, N. Y. RECEIVED JUNE 18, 1935

[CONTRIBUTION FROM GATES CHEMICAL LABORATORY, CALIFORNIA INSTITUTE OF TECHNOLOGY, NO. 488]

A Quantum Mechanical Discussion of Orientation of Substituents in Aromatic Molecules

BY G. W. WHELAND AND LINUS PAULING

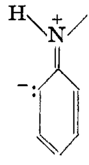
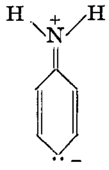
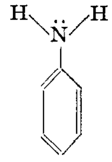
Introduction

When a substituent is introduced directly into an aromatic molecule, it may enter into certain of the available positions more readily than into others. This phenomenon of *orientation* has been exhaustively studied, and empirical rules have been found which describe the experimental results fairly satisfactorily. In a monosubstituted benzene C_6H_5R, for example, the introduction of a second substituent is governed by the nature of the atom or group, R, already present. Thus, for R = F, Cl, Br, I, OH, NH_2, and so on, the further substitution takes place largely in the ortho and para positions, while for R = COOH, CHO, NO_2, $(CH_3)_3N^+$, SO_3H, and so on, the substitution takes place largely in the meta position. Most ortho–para directing substituents, with the exception of fluorine, chlorine and bromine, activate the molecule so that the second substitution takes place more easily than in benzene itself. The meta directing substituents, on the other hand, produce deactivation. For other aromatic molecules similar rules have been found: in naphthalene substitution takes place largely at the α-position; in furan, pyrrole, and thiophene it takes place at the α-position; in pyridine it takes place at the β-position; and so on. In all the cases named except the last the molecule is more reactive than benzene.

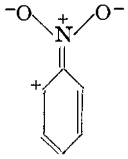
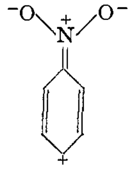
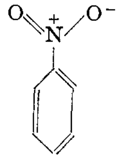
Several qualitative theories have been proposed to explain the phenomenon of orientation, but of these we shall mention only a single one, which has been developed during the last fifteen years. This theory, which has achieved a considerable measure of success, and which we believe to be essentially correct, presents the following picture of the process of directed substitution. For definiteness we shall consider a mono-substituted benzene, since this is the type of molecule which has been most carefully studied from both the experimental and the theoretical standpoints. In the molecule C_6H_5R, with R attached to carbon atom 1, the electron distribution may be such as to place an excess or a deficiency of electrons on the various carbon atoms 2–6. Moreover, the electron distribution may be changed somewhat on the approach of a group R′ to one of the carbon atoms ("polarization" of the molecule by the group). We assume that *the rate of substitution of R′ for hydrogen on the* i*th carbon atom increases with increase in the negative charge of the* i*th carbon atom when the group R′ approaches it.* We thus take into consideration, in addition to the permanent charge distribution, the changes in it caused by the approaching group.

There are two principal ways in which the charge distribution is affected by the substituent R. The first, called the *inductive effect,* results when the electron affinity of this group is appreciably larger, or smaller, than that of the hydrogen atom which it replaces. Thus, for example, if its electron affinity is larger, it will attract to itself a larger proportion of the electrons which form the bond between it and the carbon atom. This carbon atom will accordingly be left with a small positive charge, compared with the others, and its own electron affinity will be correspondingly increased. Consequently, it will tend to appropriate some of the electrons belonging to its neighbors, and so on around the ring. The net effect is, then, a transference of negative electricity from the ring to the substituent R. In an analogous manner, if R has a smaller electron affinity than the hydrogen atom which it replaces, it will give up negative electricity to the ring. In order to obtain agreement with experiment, it is necessary to assume that the *o*- and *p*-positions are affected by this redistribution of charge to approximately equal extents, and that the *m*-positions remain nearly unaffected. This is at variance with the simple classical picture, which predicts the order ortho > meta > para, but it finds its explanation in the quantum mechanical treatment given later.

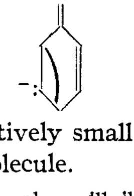
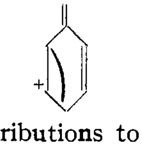
The second way in which the substituent R affects the charge distribution of the molecule is called the *resonance effect* (or sometimes the *tautomeric* or *electromeric effect*). This results when the molecule resonates among several electronic structures. For example, for aniline the structures


and

as well as


contribute to the normal state of the molecule, and for nitrobenzene


and

as well as


In the first molecule the resonance produces an increased concentration of electrons on the *o* and *p* carbon atoms, and in the second it produces a decreased concentration. As with the inductive effect, the meta positions remain nearly unaffected, since structures of the type


and


make relatively small contributions to the state of the molecule.

Some examples will illustrate the application of the theory to particular molecules. In toluene, the observed direction of the dipole moment shows that electrons have been transferred from the methyl group to the ring. Hence substitution takes place readily, particularly at the *o*- and *p*-positions, which have the greatest excess of electrons.

In nitrobenzene the nitro group has a large electron affinity, and accordingly draws electrons away from the ring. The resonance effect works in the same direction, and, as a result, all positions have a deficiency of electrons. The meta positions are least affected, and the substitution takes place there with difficulty. In aniline, the inductive effect and the resonance effect oppose each other, but the latter wins out, and very easy *o-p* substitution takes place.

These ideas have arisen through the combined efforts of a large number of different workers, among whom may be mentioned Fry, Stieglitz, Lapworth, Lewis, Lucas, Lowry, Robinson, and especially Ingold. For details the reader is referred to the excellent review articles of the last named author.[1]

With the development of a detailed quantum mechanical theory of the structure of aromatic molecules, it has become possible to attack the problem of orientation quantitatively. An attempt to do this was made by Hückel,[2] who showed that a polarizing substituent in benzene induces alternating charges in the ring. He limited his calculations, however, to the inductive effect, and neglected the resonance effect as well as the polarization of the molecule by the attacking group. As a result of these over-simplifications, he was led to make incorrect physical and chemical assumptions in order to obtain a rough correlation between his calculations and experiment. Thus, he assumed that "negative" groups, such as Cl, NH_2, and so on, increase the electron affinity of the attached carbon atom, while "positive" groups, such as NO_2, SO_3H, and so on, decrease its electron affinity. (Actually, all of the groups named tend to draw electrons away from the attached carbon atom, as far as the inductive effect alone is concerned, and hence to increase its electron affinity.) With this assumption, he then showed by a perturbation calculation that the negative groups induce positive charges on the *o*- and *p*-positions, while the positive groups induce negative charges there. This can be reconciled with the facts of orientation only if the ease of substitution is assumed to increase with increase of *positive* charge at the point of attack.

This is the exact opposite of the rule stated above, and appears to be in definite disagreement with other chemical evidence.[2] It is, furthermore, incapable of giving a consistent picture even of the process of orientation, as the following examples will show. (1) The observed direction of the dipole moment of toluene shows that electrons have been transferred from the methyl group to the ring. The methyl group is accordingly a positive one, in the sense of the foregoing discussion, and consequently should direct meta. Actually it directs ortho–para. (2) In pyridine the electron affinity of the nitrogen atom is larger than that of the carbon atoms. Consequently the charge distribution is the same as in benzene with a very strongly negative substituent. Substitution should accordingly take place in the α and γ positions (analogous to *o* and *p*) with great ease, but actually it takes place in the β position (analogous to *m*) with considerable difficulty.

The error in Hückel's treatment lies not in the quantum mechanical calculations themselves, which are correct as far as they go, but in the over-simplification of the problem and in the incorrect interpretation of the results. Consequently it has seemed desirable to us to make the necessary extensions and corrections in order to see if the theory can lead to a consistent picture. In the following discussion we have found it necessary to consider all of the different factors mentioned heretofore: the resonance effect, the inductive effect, and the effect of polarization by the attacking group. The inclusion of these several effects in the theory has led to the introduction of a number of more or less arbitrary parameters, and has thus tended to remove significance from the agreement with experiment which is achieved. We feel, however, that the effects included are all justified empirically and must be considered in any satisfactory theory, and that the values used for the arbitrary parameters are reasonable. The results communicated in this paper show that the quantum mechanical theory of the structure of aromatic molecules can account for the phenomenon of directed substitution in a reasonable way.

Outline of the Method

There are two principal methods available for the quantum mechanical treatment of molecular structure, the valence bond method and the molecular orbital method. In this paper we shall make use of the latter, since it is simpler in form and is more easily adapted to quantitative calculations.[3] We accordingly consider each electron

---

[1] C. K. Ingold, *Ann. Rep.*, **23**, 129 (1926); *Rec. trav. chim.*, **48**, 797 (1929); *J. Chem. Soc.*, 1120 (1933); *Chem. Rev.*, **15**, 225 (1934).

[2] E. Hückel, *Z. Physik*, **72**, 310 (1931). For criticism of this work from the chemical side see A. Lapworth and R. Robinson, *Nature*, **129**, 278 (1932); **130**, 273 (1932); see also E. Hückel and W. Hückel, *ibid.*, **129**, 937 (1932).

[3] For purely qualitative arguments, on the other hand, the former is quite convenient. The description of the theory up to this point has been, in fact, merely a statement of the qualitative conclusions of the valence bond theory.

to be moving essentially alone in a self-consistent field, which is produced by the nuclei and by all the electrons other than itself, and we approximate the individual one-electron wave functions, φ_j, by means of linear combinations of the orbitals, ψ_k, belonging to the different atoms.

$$\varphi_j = \sum_k a_{jk}\,\psi_k \qquad (1)$$

In systems of the type under consideration the aromatic rings lie in a single plane. Consequently the atomic s and p orbitals, ψ_k, to which the present discussion will be limited, can be divided into two classes, depending upon whether they are symmetric or antisymmetric with respect to reflection in this plane. Since orbitals from one of these classes do not combine with orbitals from the other, each of the summations of Eq. 1 needs to be extended only over the members of a single class. Electrons will of course be assigned to molecular orbitals, φ_j, of both types, but it will be sufficient for our purposes to consider only the ones which are formed from the antisymmetric atomic orbitals—the $[p]_h$ orbitals in Hückel's nomenclature. In energy calculations this is a legitimate simplification, since it affects merely the arbitrarily chosen zero point of energy, but in the present treatment it requires further justification, since we are now interested in the charge distribution. There is, however, considerable experimental evidence to show that π bonds, such as are formed by the $[p]_h$ orbitals, are much more polarizable than σ bonds, such as are formed by the symmetric orbitals. This is in accord with their smaller binding energy, and also with the fact that orientation phenomena are not common in saturated systems. Consequently we shall assume that the inclusion of the polarization of the symmetric orbitals would result in only a negligibly small correction.

We return now to those molecular orbitals, φ_j, of Eq. 1 which are expressed in terms of the $[p]_h$ functions. The variation method shows that the best values of the coefficients a_{jk} are those which satisfy the equations

$$\sum_{k=1}^{l} a_{jk}\,(H_{mk} - \Delta_{mk}\,W_j) = 0.\ j.m = 1,2,\ldots.l \qquad (2)$$

where l is the total number of $[p]_h$ functions (one for each atom present, except the hydrogens);

$$H_{mk} = \int \psi_m^*\, H\psi_k\, d\tau \qquad \Delta_{mk} = \int \psi_m^*\, \psi_k\, d\tau$$

and W_j is the energy of an electron occupying the molecular orbital φ_j.

For each value of j there is a set of equations, but the different sets are all identical except for the replacement of the subscript j by some other, say j'. The condition for the solubility of this set of equations is that their determinants vanish. The l roots of this secular equation, $W = W_j$ ($j = 1,2,\ldots,l$) give the energies of the l molecular orbitals. We then assign two electrons (one with positive and one with negative spin) to the lowest root, two more to the next lowest root, and so on until all have been used up. The values of W_j found in this manner are then substituted back into the Equations 2, from which the ratios of the values of the coefficients a_{jk} can be derived. The absolute magnitudes of the coefficients are determined by the normalization condition, $\int \varphi_j^*\, \varphi_j d\tau = 1$.

$$|H_{mk} - \Delta_{mk}\,W| = \begin{vmatrix} H_{11} - \Delta_{11}W & H_{12} - \Delta_{12}W & \cdots & H_{1l} - \Delta_{1l}W \\ H_{21} - \Delta_{21}W & H_{22} - \Delta_{22}W & \cdots & H_{2l} - \Delta_{2l}W \\ \cdots & \cdots & \cdots & \cdots \\ H_{l1} - \Delta_{l1}W & H_{l2} - \Delta_{l2}W & \cdots & H_{ll} - \Delta_{ll}W \end{vmatrix} = 0 \qquad (3)$$

An electron occupying the molecular orbital φ_j will spend a fraction of its time given by the expression $|\,a_{jk}\,|^2$ in the atomic orbital ψ_k. Consequently the total average charge on atom k will be

$$C_k = e\sum_j |\,a_{jk}\,|^2 \qquad (4)$$

where the summation is extended over all occupied orbitals φ_j (the orbital being counted twice if occupied twice),[4] and e represents the charge of the electron.

As usual in these calculations, we make the following assumptions in regard to the magnitudes of the integrals H_{mk} and Δ_{mk}.

$H_{kk} = q + \delta_k\beta$
$H_{mk} = \beta$, if ψ_m and ψ_k are on adjacent atoms,
$= 0$, otherwise, for $m \neq k$
$\Delta_{kk} = 1$
$\Delta_{mk} = 0$, for $m \neq k$

Here q represents the coulomb energy of an electron occupying a definite $[p]_h$ orbital in unsubstituted benzene; its value has been estimated to be about -2.7 v. e. $= -60$ kcal./mole.[5] β is a *resonance integral* between adjacent orbitals; its value has been estimated to be about -0.85 v. e. $= -20$ kcal./mole.[5] δ_k is a constant, the purpose of which is to allow for the different electron affinities of the different atoms. For $\delta_k > 0$, the

---

[4] Strictly speaking, the complete eigenfunction for the molecule should be made antisymmetric before the charge densities at the various positions are calculated. It is easily shown, however, that this further refinement in the treatment does not alter the results obtained.

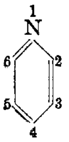
[5] G. W. Wheland, *J. Chem. Phys.*, **2**, 474 (1934).

atom k has a larger, and for $\delta_k < 0$, a smaller electron affinity than a carbon atom in unsubstituted benzene.

We have attempted to estimate values of the δ's for the different kinds of atoms with which we will deal later, but we have not been able to arrive at any very consistent figures. Our procedure was to determine the values which would bring the calculated resonance energies into agreement with the empirical ones. The figures obtained in this way averaged about +4 for oxygen and +2 for nitrogen, and accordingly were of the correct sign, and of a reasonable order of magnitude, but the spread in both cases was quite large. Moreover, it is questionable if δ values obtained from energy calculations can be applied directly to calculations of charge distribution, when the method of treatment is as rough as the present. Accordingly we shall not choose a definite set of δ's which we shall use throughout, but instead shall allow the values to vary within fairly wide limits, in order to see the effect of such a variation upon the calculated polarization. Very fortunately it turns out that the decisive factor in most cases is not the magnitude of δ but its sign, which can be predicted with almost complete assurance.

Our method of calculation is now quite straightforward. We account for the permanent polarization resulting from the inductive effect by introducing suitable δ's upon the proper atoms; we account for that resulting from the resonance effect by explicitly including in the treatment all the $[p]_h$ orbitals of the problem, and not merely those on the atoms of the ring; and finally we account for the polarizing effect of the reacting group by introducing a small δ at the point of attack. For the sake of simplicity, we shall first treat only the permanent polarization, and then later consider the polarizability at the different positions.

Application to the Calculation of the Permanent Polarization



Pyridine.—We shall begin the discussion with pyridine rather than with a substituted benzene, since it offers a particularly simple example, in which only the inductive effect is operative. We number the atoms in order, beginning with nitrogen. Then δ_1, which takes into account the electron affinity of the nitrogen atom, will be positive in sign, and of the order of magnitude of perhaps 2; $\delta_2 = \delta_6$ will also be positive in sign as a result of induction, but will be much smaller in magnitude than δ_1; the remaining δ's can be considered negligibly small. The secular equation is then

$$\begin{vmatrix} x+\delta_1 & 1 & 0 & 0 & 0 & 1 \\ 1 & x+\delta_2 & 1 & 0 & 0 & 0 \\ 0 & 1 & x & 1 & 0 & 0 \\ 0 & 0 & 1 & x & 1 & 0 \\ 0 & 0 & 0 & 1 & x & 1 \\ 1 & 0 & 0 & 0 & 1 & x+\delta_6 \end{vmatrix} = 0$$

where, for simplicity in writing, we have divided each row by β and set $(q-W)/\beta = x$. We cannot solve this equation rigorously unless we put into it definite numerical values for δ_1 and $\delta_2\ (=\delta_6)$. The perturbation theory, however, gives a method by which we can expand both the solutions, W_j, and the eigenfunctions, φ_j, in rising powers of the δ's. If we carry the procedure only as far as the first order, we obtain linear expressions which represent approximately the variation of the different quantities of interest for small values of the δ's. The unperturbed secular equation, which is obtained by setting $\delta_1 = \delta_2 = \delta_6 = 0$, is just the same as the one which arises in the treatment of benzene. Its roots are

$$W_j^0 = q + 2\beta\cos(2\pi j/6)$$
$$j = 1,2,\ldots,6$$

and the corresponding wave functions are

$$\varphi_j^0 = \frac{1}{\sqrt{6}}\sum_{k=1}^{6} e^{(2\pi i/6)jk}\,\psi_k$$
$$i = 1,2,\ldots,6 \tag{5}$$

The familiar equations of the first order perturbation theory now show that the corresponding perturbed quantities are

$$\left.\begin{aligned} W_j &= \int \varphi_j^{0*}H\varphi_j^0\,d\tau = W_j^0 + \int \varphi_j^{0*}H'\varphi_j^0\,d\tau \\ \varphi_j &= \varphi_j^0 + \sum_k{}' \frac{\int \varphi_k^{0*}H'\varphi_j\,d\tau}{W_j^0 - W_k^0}\varphi_k^0 \end{aligned}\right\} \tag{6}$$

where H' is the part of the complete Hamiltonian operator which refers to the perturbation, such that

$$\left.\begin{aligned} \int \psi_k^* H'\psi_l\,d\tau &= \delta_1\beta \text{ if } k = l = 1 \\ &= \delta_2\beta \text{ if } k = l = 2 \text{ or } 6 \\ &= 0 \quad \text{otherwise} \end{aligned}\right\} \tag{7}$$

and the prime on the summation sign indicates that the term with $k = j$ is omitted. When Eqs. 5 and 7 are inserted in Eq. 6 and use is made of Eq. 4, the following results are obtained.

$$C_N = C_1 = \left(1 + \frac{43}{108}\delta_1 - \frac{17}{54}\delta_2\right)e$$

$$C_\alpha = C_2 = C_6 = \left(1 - \frac{17}{108}\,\delta_1 + \frac{22}{54}\,\delta_2\right) e$$
$$C_\beta = C_3 = C_5 = \left(1 + \frac{1}{108}\,\delta_1 - \frac{14}{54}\,\delta_2\right) e$$
$$C_\gamma = C_4 = \left(1 - \frac{11}{108}\,\delta_1 + \frac{1}{54}\,\delta_2\right) e$$

(neglecting terms in δ_1^2, $\delta_1\delta_2$ and δ_2^2). These figures differ somewhat from the corresponding ones given by Hückel,[2] because he put $\delta_2 = 0$ and also, in applying the perturbation treatment, he made use of a rather different procedure which so complicated the calculations that he did not carry them through completely.

If, now, δ_2 lies within the limits $^1/_{28}\,\delta_1 < \delta_2 < ^1/_4\,\delta_1$, which seem quite reasonable, the over-all effect of the perturbation is a transference of negative electricity from the various carbon atoms to the nitrogen atom at position 1, with the consequence, in accordance with the rule formulated above, that the molecule is deactivated so that substitution is more difficult than in benzene; and furthermore the α and γ positions (2 and 6, and 4, respectively) are most affected by this transference of electricity, so that substitution will take place at the β positions (3 and 5), which have the smallest deficiency of electrons. Both of these conclusions are borne out by experiment.

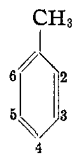
This calculation and its comparison with experiment hence provide us with a rough idea as to the magnitude of the inductive effect from atom to atom, as given by the ratio δ_2/δ_1. If this ratio were greater than $^1/_4$ the effect of δ_2 would overcome that of δ_1 and pyridine would substitute in the α and γ positions, whereas if it were less than $^1/_{28}$ substitution would occur in the β positions more readily than in benzene. A value of about $^1/_{10}$ for δ_2/δ_1 seems reasonable to us from a consideration of the electronic phenomena involved.

In order to see how accurate this perturbation treatment actually is, we have substituted numerical values for the δ's directly into the secular equation, and then solved it rigorously by numerical methods. The calculations are not given in detail, since they are quite straightforward and proceed along well-known lines. The results are shown in Table I.

TABLE I

δ_1	1	2	3	2
$\delta_2 = \delta_6$	0	0	0	0.2
$C_N = C_1$	1.37e	1.62e	1.79e	1.59e
$C_\alpha = C_2 = C_6$	0.85	0.76	0.71	0.83
$C_\beta = C_3 = C_5$	1.01	1.01	1.01	.96
$C_\gamma = C_4$	0.91	0.84	0.79	.82

It is seen that for $\delta_1 = 1$ the perturbation treatment is still fairly accurate, but that for larger values it is only qualitatively satisfactory. In any case, however, if we assume that the ease of substitution increases with increased negative charge at the point of attack, the calculations are in very good agreement with experiment; if, on the other hand, we follow Hückel in making the opposite assumption, the calculations can be reconciled with experiment only by making δ_1 negative, which seems highly improbable in view of the fact that the nuclear charge of nitrogen is greater than that of carbon.



Toluene.—The toluene molecule is complicated in that it contains atomic orbitals (in the methyl group) which are neither symmetric nor antisymmetric with respect to reflection in the plane of the ring. However, since it seems very unlikely that there will be any appreciable resonance between the methyl group and the ring, we are able to avoid the difficulty by formally replacing the substituent by a small δ value on the adjacent carbon atom. The dipole moment of the molecule requires that this δ be negative in sign. It is also presumably small enough so that the same perturbation treatment as was employed in the case of pyridine will represent a good approximation.[6] The results are then just the same as those given in Eq. 8, except that C_α, C_β, and C_γ are to be replaced by C_{ortho}, C_{meta}, and C_{para}, respectively, in order to be in accord with the customary nomenclature. Here, as before, all the the δ's except δ_1 and δ_2 can be set equal to zero. Since δ_1 is negative, we see, neglecting δ_2, that the ortho and para positions have excess negative charges, in agreement with the experimental result that substitution takes place there very readily. This result is not changed by the consideration of δ_2 so long as the ratio δ_2/δ_1 is less than $^1/_4$.

Phenyltrimethylammonium Ion, $N^+(CH_3)_3$-C_6H_5.—In this ion, as in toluene, we ignore the electrons involved in bonds from nitrogen to the attached groups, and consider only the inductive effect. The positive charge on the nitrogen atom increases its electron affinity to a value still greater than that for neutral nitrogen, so that we

---

(6) For larger δ values, the results for toluene can be read off at once from those for pyridine (Table I); reversing the sign of the δ's in a given secular equation of the present type merely reverses the direction of the induced polarization, and leaves its magnitude unchanged.

introduce in the above expressions (for toluene) a large positive value for δ_1 (and a small value for δ_2). This leads to greatly decreased values of C for ortho and para carbon atoms, with smaller decrease for meta, and in consequence to meta directed substitution, in agreement with experiment.

Nitrobenzene.—In the treatment of nitrobenzene we must consider nine $[p]_h$ orbitals and ten electrons, including an unshared pair on the negative oxygen atom in the indicated structure and one electron for each of the remaining atoms. We assume the resonance integral β to have the same value for O-N and N-C as for C-C. The results of the calculation are given in Table II, the next to the last column of which corresponds to the δ values which seem reasonable to us. It is seen that all positions are deactivated, the deactivation occurring least at the meta positions, where substitution should take place. These conclusions are in agreement with experiment.

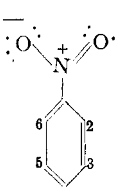


TABLE II

δ_O	0	0	4	4	4
δ_N (δ_C)	0	2	2	2	0.8
δ_1	0	0	0	0.2	0
C_O	1.50e	1.24e	1.95e	1.94e	1.92e
C_N (C_C)	1.00	1.53	1.37	1.32	0.68
C_1	1.00	0.96	0.62	0.70	.95
$C_{\text{ortho}} = C_2 = C_6$	1.00	1.01	.79	.79	.85
$C_{\text{meta}} = C_3 = C_5$	1.00	1.00	.95	.95	1.00
$C_{\text{para}} = C_4$	1.00	1.01	.63	.61	0.83

Benzoic Acid, C_6H_5COOH.—The treatment of benzoic acid is the same as for nitrobenzene except for the δ values. The results for $\delta_O = 4$ and δ_C (of the carboxyl group) $= 0.8$ are given in the last column of Table II; it is seen that the carboxyl group, like the nitro group, is meta directing, in agreement with experiment. The calculation, however, indicates a rather too small deactivation of the meta position. This effect appears only in the third decimal place of C_{meta}, and is consequently not shown in the table.

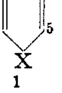
Benzaldehyde, C_6H_5CHO, **Acetophenone,** $C_6H_5COCH_3$, and **Benzonitrile,** C_6H_5CN.—The results obtained for these molecules by considering eight $[p]_h$ orbitals and eight electrons are given in Table III.

It is seen that the aldehyde, ketone and nitrile groups are all meta directing, in agreement with experiment.

TABLE III

δ_O (δ_N)	0	2	2	4	4
δ_C	0	0	0.2	0	0.4
C_O(C_N)	1.00e	1.76e	1.75e	1.92e	1.92e
C_C	1.00	0.49	0.55	0.41	0.56
C_1	1.00	1.03	1.02	1.02	.98
$C_{\text{ortho}} = C_2 = C_6$	1.00	0.90	0.89	0.88	.85
$C_{\text{meta}} = C_3 = C_5$	1.00	1.00	1.00	1.00	1.00
$C_{\text{para}} = C_4$	1.00	0.91	0.90	0.87	0.85

Furan, Thiophene, Pyrrole.—The calculations for this case, with five orbitals and six electrons, show that when $\delta_2 (= \delta_5)$ (on the α carbon atoms) is neglected the β positions (3 and 4) are more strongly activated than the α. When, however, these δ's are given a reasonable magnitude, greater than $^1/_{25}\ \delta_1$, the situation is reversed, and the α positions are found to be the more strongly activated, as is demanded by experiment. Some calculated charge values are given in Table IV.

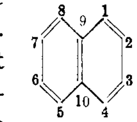


(X = O, S, NH)

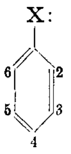
TABLE IV

δ_1	0.5	2.9	4	4	2	3
$\delta_2 = \delta_5$	0	0.29	0	0.5	0.25	0.375
$C_X = C_1$	1.38e	1.80e	1.88e	1.87e	1.69e	1.81e
$C_\alpha = C_2 = C_5$	1.14	1.08	0.98	1.13	1.10	1.13
$C_\beta = C_3 = C_4$	1.18	1.02	1.08	0.93	1.06	0.96

Naphthalene.—Naphthalene offers the first example of a molecule in which a consideration of the permanent polarization alone does not lead to the correct answer. Since this is a hydrocarbon, we might expect that no δ values would be necessary. In this case, however, the charge on each atom is found to be exactly 1.00e, so that substitution would take place at the α and β positions (1, 4, 5, 8 and 2, 3, 6, 7, respectively) with equal ease. If we attempt to avoid this difficulty by putting a δ upon the 9 and 10 carbon atoms, we find that its sign must be negative in order to produce an excess of electrons on the α positions. This seems unreasonable, however, since these carbon atoms carry no hydrogens, and would consequently be expected to have a larger rather than a smaller electron affinity than the others. We shall see later that a possible explanation of this discrepancy can be found by a consideration of the polarizing effect of the attacking group.



Aniline, Phenol, Fluorobenzene, etc.—The treatment of this group of molecules, involving seven $[p]_h$ orbitals and eight electrons, is comparatively unsatisfactory, in consequence of the fact that in them the inductive and the resonance effects work in opposite directions. In order to be able to predict which of the two will win out in any specific example, we need to have fairly accurate numerical values of the different δ's to be introduced in each case. In the absence of such information, the best we can do is to carry through a number of calculations, in order to see if any reasonable sets of values lead to agreement with experiment. The results are given in Table V.

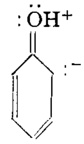
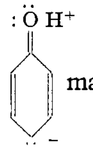


(X = NH_2, OH, F, etc.)

TABLE V

δ_X	0	2	4	6	2	4	6
δ_1	0	0	0	0	0.25	0.4	0.5
C_X	1.57*e*	1.91*e*	1.96*e*	1.98*e*	1.91*e*	1.97*e*	1.98*e*
C_1	1.00	0.93	0.95	0.95	1.02	1.10	1.15
$C_{\text{ortho}} = C_2 = C_6$	1.14	1.06	1.04	1.02	1.02	0.97	0.95
$C_{\text{meta}} = C_3 = C_5$	1.00	1.00	1.00	1.00	1.00	1.00	1.00
$C_{\text{para}} = C_4$	1.14	1.05	1.02	1.02	1.02	0.98	0.96

It will be seen that, as might be expected from purely qualitative considerations, the inductive effect, which deactivates the molecule and favors *m* orientation, is increased in importance with increase of the δ's, while the resonance effect, which activates the molecule and favors *o-p* orientation, is decreased in importance. This is in agreement with the fact that the *o-p* directing power of the halogens is in the order F < Cl < Br < I, which is the opposite of the order of electron affinities. On the other hand, it does not explain the marked difference between the halogens as a whole and the amino and hydroxyl groups—the former being relatively feeble in orienting power, and the latter being among the most powerful known. It also does not explain why the orders $NH_2 < PH_2 < AsH_2$ and OH < SH, which would be expected from analogy, are not correct.

It is possible that the explanation of these discrepancies is to be found in the fact that the resonance integral, β, may vary with the row and group of the periodic table. Such a variation must almost certainly exist, but it can be taken into account only with difficulty. Furthermore, the introduction of the large number of additional arbitrary parameters would deprive the whole procedure of much of its significance. A second possible explanation is that, with phenol for example, resonance with such structures as  and  may make relatively important contributions to the state of the molecule. The inclusion in the theory of such effects also would be difficult and would result in a further decrease in the significance of the treatment. Instead of resorting to such heroic measures to force the calculations into agreement with experiment, we prefer to let the matter stand as a possible weakness of the theory. It would not be surprising if our rough treatment should fall down in evaluating the relative importance of two opposed effects which differ so greatly in nature.

A further difficulty in the case of fluoro-, chloro- and bromobenzenes is that with them apparently no choice of the δ values seems to be reconcilable with the observed order of ease of substitution at the various positions: unsubstituted benzene > para > ortho > meta. Both the inductive and the resonance effects are seen to leave the charge on the *m*-position practically unchanged, and approximately equal to 1.00*e*, while the observed order demands a considerably smaller value. As in the case of naphthalene, however, we shall find later that this discrepancy can apparently be explained by taking into account the polarization of the molecule by the attacking group.

Polarization by the Attacking Group

The calculations so far have been concerned with the "permanent" charge distribution which the molecule would have when isolated, and which would be observed in its dipole moment. The factor that actually determines the orientation of the entering group, however, is the charge distribution at the instant of attack, which differs from the permanent charge distribution in consequence of polarization by the entering group. In most cases it is not necessary to take this latter effect into account explicitly, since it represents only a small correction. However, in certain cases it is apparently of decisive importance.[7]

In order to investigate the relative ease of substitution at the *i*th carbon atom, we shall introduce a small δ_i upon that atom, in addition

---

[7] This was suggested originally by Ingold, *cf.* Ref. 1.

to any other δ's that may be required by the permanent polarization. This δ_i will presumably be positive in sign in all cases, since the customary reacting groups are "electron-seeking" and will accordingly tend to increase by induction the electron affinity of the near-by carbon atom.

In benzene itself the charge at the point of attack will be not 1.00e, as has been assumed heretofore, but $(1.00 + (43/108)\delta_i)e$. (Since δ_i is small, the results of the perturbation treatment can be applied directly.) In other molecules the ease of reaction at the ith carbon atom will then be greater than that of benzene if the charge at that point is greater than $(1.00 + (43/108)\delta_i)e$, and conversely.

Let us apply these considerations to pyridine. If we set $\delta_N = \delta_1 = 2$, $\delta_\alpha = \delta_2 = \delta_6 = 0.2$, and $\delta_i = 0.2$, we obtain the results given in Table VI. As might be expected, the polarizability of the molecule at the various positions is practically constant, and is very nearly the same as that of benzene. Consequently the orientation is not affected by considering the polarization by the approaching group. This state of affairs can be expected to hold whenever the permanent polarization produces marked differences, as here, between the different positions.

TABLE VI

POLARIZABILITY OF PYRIDINE

i[a]	Charge at ith atom			Charge in benzene
	$\delta_i = 0$	$\delta_i = 0.2$	Difference	$(1.00 + \frac{43}{108} 0.2)e$
2	0.831e	0.920e	0.089e	1.080e
3	.960	1.038	.078	1.080
4	.825	0.905	.080	1.080

[a] For the numbering of the different atoms see the diagrams given previously.

In naphthalene, on the other hand, the permanent polarization leaves all atoms with the same or very nearly the same charge, and as a result any small difference in polarizability has an opportunity to make its effect felt. The results of perturbation treatments are given in Table VII. With $\delta_i > 0$ the α position is activated, and substitution would take place there readily, as is demanded by experiment.

TABLE VII

POLARIZABILITY OF NAPHTHALENE

i	Charge on ith atom	Charge in benzene
1	$(1.00 + 0.44\delta_i)e$	$(1.00 + 0.40\delta_i)e$
2	$(1.00 + .40\delta_i)e$	$(1.00 + .40\delta_i)e$

In the halogen benzenes the ortho and para positions are found to be more polarizable than the meta, as is shown in Table VIII.

TABLE VIII

POLARIZABILITY OF $C_6H_5\ddot{X}$

$\delta_X = 4.0$, $\delta_1 = 0.4$

i	Charge at ith atom			Charge in benzene
	$\delta_i = 0$	$\delta_i = 0.2$	Difference	
2	0.975e	1.055e	0.080e	1.080e
3	1.002	1.069	.067	1.080
4	0.984	1.064	.080	1.080

With the particular choice of the δ's used in the calculation the observed order[8] unsubstituted benzene > para > ortho > meta is not obtained, but agreement can be achieved by a suitable readjustment of the parameter values. If, with δ_X held fixed and equal to 4, the charges at the different positions are assumed to be linearly dependent upon δ_1 and δ_i (as is the case when these quantities are sufficiently small), the results of Table VIII can be expressed as

$$\begin{aligned}\text{Charge on 2} &= (1.036 - 0.152\delta_1 + 0.402\delta_i)e\\ \text{Charge on 3} &= (0.998 + 0.009\delta_1 + 0.339\delta_i)e\\ \text{Charge on 4} &= (1.024 - 0.099\delta_1 + 0.400\delta_i)e\end{aligned}$$

If we use the particular values $\delta_1 = 0.3$, $\delta_i = 0.2$ (any of several other sets would do as well), we find

$$\begin{aligned}\text{Charge on 2} &= 1.071e\\ \text{Charge on 3} &= 1.069e\\ \text{Charge on 4} &= 1.074e\end{aligned}$$

These figures are now in agreement with the observed order given above. While it is doubtful if much significance can be attached to parameter values obtained in such an arbitrary and *ad hoc* manner, it is nevertheless gratifying to observe that the ones found are not unreasonable.

Conclusion

By introducing reasonable values (about 2 for nitrogen, 4 for oxygen) for the electron affinity parameter relative to carbon, δ, and for the induced electron affinity for adjacent atoms ($\delta_2/\delta_1 \cong 1/10$), we have shown that the calculated permanent charge distributions for pyridine, toluene, phenyltrimethylammonium ion, nitrobenzene, benzoic acid, benzaldehyde, acetophenone, benzonitrile, furan, thiophene, pyrrole, aniline, and phenol can be satisfactorily correlated qualitatively with the observed positions and rates of substitution. For naphthalene and the halogen benzenes this calculation does not lead to results

---

(8) Iodobenzene substitutes more easily than benzene. In this case, however, δ_X is comparatively small, and the discrepancy does not exist.

in complete qualitative agreement with experiment; agreement is obtained, however, by the consideration of the effect of polarization of the molecule by the attacking group. Because of the presence of several adjustable parameters, this treatment cannot be considered as a pure quantum mechanical theory of directed substitution in aromatic molecules. However, the various effects which we have considered are certainly real, and there is no reason to doubt that they act in the directions calculated, and are of the general order of magnitude assumed; the parameters introduced are not completely arbitrary in value, because they are related to definite and fairly well understood atomic and molecular quantities. Furthermore, in the cases in which only one effect is important, as in pyridine, or in which all effects act in the same direction, as in nitrobenzene, the calculations lead to completely unambiguous conclusions, which are dependent to only a small degree upon the particular parameter values chosen. We accordingly feel that the degree of success realized in these quantitative considerations provides strong support for the underlying ideas as to the nature of the effects involved.

Summary

Using the method of molecular orbitals, a quantitative discussion of the charge distribution in aromatic molecules undergoing substitution reactions is carried out, taking into consideration the inductive effect, the resonance effect, and the polarizing effect of the attacking group. It is shown that, with reasonable values for the parameters involved, the calculated charge distributions for pyridine, toluene, phenyltrimethylammonium ion, nitrobenzene, benzoic acid, benzaldehyde, acetophenone, benzonitrile, furan, thiophene, pyrrole, aniline, phenol, naphthalene, and the halogen benzenes are in qualitative agreement with the experimental results regarding position and rate of substitution, the auxiliary hypothesis being made that the rate of substitution of the group R′ for hydrogen on the *i*th carbon atom increases with increase in the negative charge of the *i*th carbon atom when the group R′ is in position to react.

PASADENA, CALIFORNIA RECEIVED JULY 26, 1935

[CONTRIBUTION FROM THE CHEMICAL LABORATORY OF IOWA STATE COLLEGE]

Dibenzofuran. VII. Derivatives of Tetrahydrodibenzofuran

BY HENRY GILMAN, E. W. SMITH AND L. C. CHENEY

Introduction

The presence of a reduced dibenzofuran nucleus in morphine suggested an examination of nuclearly substituted reduced dibenzofurans. From the results of the present study on 1,2,3,4-tetrahydrodibenzofuran it appears (1) that the orientation of the parent dibenzofuran cycle is essentially unlike that of 1,2,3,4-tetrahydrodibenzofuran; (2) that the position assigned to the nitro group in nitro-1,2,3,4-tetrahydrodibenzofuran is incorrect; and (3) that the nuclear substituted *hexa*hydrodibenzofurans are in reality substituted *tetra*hydrodibenzofurans.

Nuclear substitution of dibenzofuran strikingly illustrates that the rule of conservation of the substitution type is not inviolate, and that the position assumed by substituents is significantly dependent on the kind of entering group. For example, with dibenzofuran, halogenation, sulfonation and the Friedel–Crafts reaction involve the 2-position; nitration involves predominantly the 3-position with only a small amount of 2-nitro compound being formed; and metalation (the replacement of hydrogen attached to carbon by a metal) takes place in the 4-position. With 1,2,3,4-tetrahydrodibenzofuran, sulfonation, the Friedel–Crafts reaction and nitration involve the 7-(or 3′-)position; whereas metalation proceeds as with dibenzofuran to give a 6-(or 4′-) organometallic compound. Bromination may be difficult because bromine is at this time a reagent of choice for dehydrogenation of tetrahydrodibenzofurans to dibenzofurans.

The position of the nitro group in nitro-1,2,3,4-tetrahydrodibenzofuran was established as follows