SUBSTITUTION AND ORIENTATION IN THE BENZENE RING

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The purpose of this paper is to summarize recent investigations of the mechanism of substitution reactions in benzenoid aromatic compounds and to present an interpretation of the nature of the influence exerted by substituent groups on the course of these reactions. In the latter connection, a method for the precise prediction of the orienting effect of a large variety of substituent groups is described, the basis for which is the accurate evaluation of the relative electrostatic influence of the various groups in polarizing the double bonds of the aromatic ring.

I. THE STRUCTURE OF BENZENE

In spite of the great number of formulas proposed for benzene, each with unique structural features to account for its "aromatic" properties,[1] the structural formula proposed by Kekulé for benzene has become more and more firmly established. In fact, it is now generally accepted that the many characteristic reactions, as well as the unusual stability of the aromatic nucleus, are simply manifestations of such a cyclic conjugated system of double bonds.

In this connection, several authors (13, 60) have recently pointed out that the compound prepared by Willstätter (66, 67), supposedly cycloöctatetraene, actually had chemical and physical properties closely resembling

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[1] For references and critical discussion see L. F. Fieser in *Organic Chemistry*, edited by H. Gilman, Volume I, pp. 52–149, John Wiley and Sons, Inc., New York (1938).

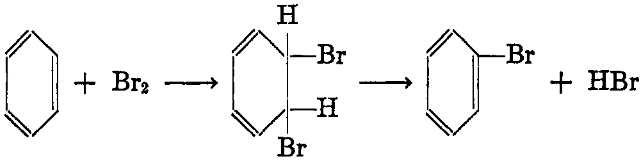
those of an isomer, styrene. Thus, the experimental evidence cited as one of the chief arguments against the "aromaticity" of a cyclic conjugated system of double bonds may be entirely erroneous.

The Kekulé formula in its resonant forms will therefore be used as the basis for the present discussion. Indeed, the facility with which the various experimental facts may be interpreted on this basis affords still further indication as to the validity of this formula.

II. SUBSTITUTION REACTIONS

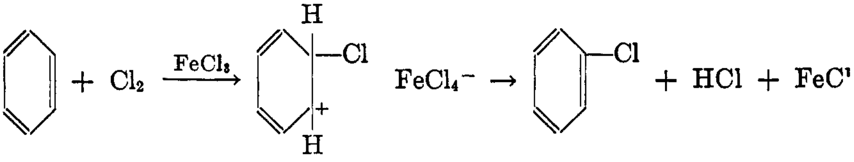
A. Halogenation

Since many of the reagents which will *substitute* in benzene will also *add* to olefins, and since, in the Kekulé formula, benzene is written with double bonds, it has frequently been postulated that the latter process is an intermediate in the former. Thus, for instance, the bromination of benzene has been represented by Holleman (18) and others as addition of bromine, followed by elimination of hydrogen bromide.

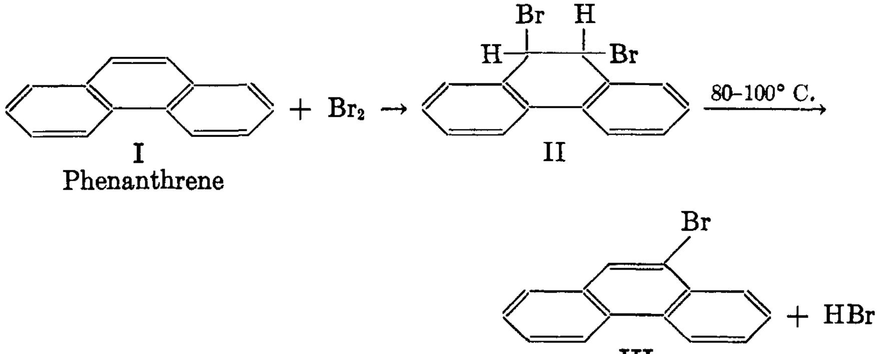


Prins (48) suggested that only one atom of the benzene ring could be involved rather than two, as in addition to a double bond, since the dihydrobenzenoid intermediate above should be too readily attacked further by the substituting reagent. van Alphen (57), however, found that 4,5-dihydrophthalic acid yielded phthalic acid derivatives on treatment with various substituting reagents, invalidating this particular objection to the addition–elimination theory.

In the same year, Pfeiffer and Wizinger (39), on the basis of the colored ionic intermediates formed in the catalytic addition and substitution reactions of the double bond in unsymmetrical diaryl ethylene derivatives, suggested a mechanism for aromatic substitution involving the intermediate formation of an ionic complex. In the case of halogenation, the purpose of a halogen carrier (A), such as ferric chloride, iodine, antimony pentachloride, etc., is to promote the polarization of the halogen molecule and thus the formation of the ionic intermediate.



Both the mechanism involving addition followed by elimination, as suggested by Hollemann, and that involving an ionic intermediate, as suggested by Pfeiffer and Wizinger, were, however, supported chiefly by analogy rather than by any direct experimental evidence. For example, it has never been possible to isolate or detect the presence of such dihydrobenzenoid derivatives as those postulated as intermediates in the addition–elimination mechanism for the case of substitution in the benzene ring. With phenanthrene, however, the reaction can actually proceed by such a mechanism. Bromine will add to form a dibromide which is stable at room temperature but which, on warming, loses hydrogen bromide to yield 9-bromophenanthrene. This reaction has, therefore, frequently been considered as evidence for the addition–elimination hypothesis.



I
Phenanthrene

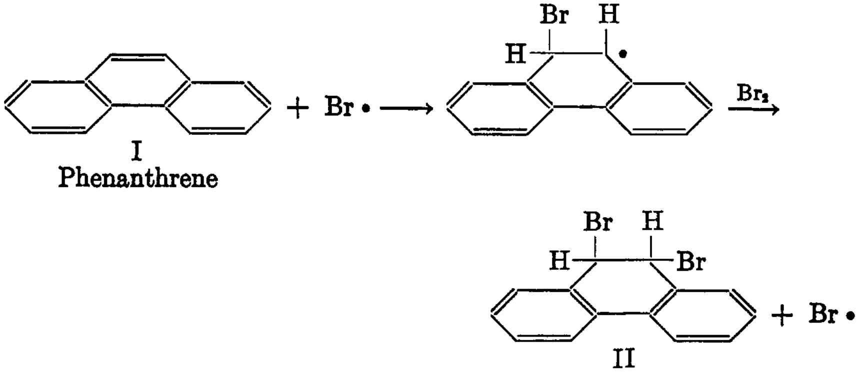
II

III
9-Bromophenanthrene

On this account a study of the bromination of phenanthrene offered an excellent opportunity for the determination of the mechanism of substitution. A detailed investigation of this reaction (10, 40, 41, 44), with particular regard to the effect of halogenation catalysts, has shown that the catalytic substitution does *not* proceed according to this addition–elimination scheme, but that the experimental facts are in complete accord with the ionic mechanism suggested by Pfeiffer and Wizinger.

In the first place, it was found that, at 25°C. in carbon tetrachloride solution with no halogenation catalysts present, only the addition reaction occurred and that it proceeded to a measurable equilibrium by a chain mechanism which was entirely light-catalyzed (24, 44). The chain initiating step in the photochemical addition of bromine to a double bond in a non-polar solvent has been found to be the photochemical dissociation of bromine into bromine atoms (2, 33, 52, 61). These active atoms then initiate a chain reaction leading to the addition of bromine to the double bond.

$$Br_2 + h\nu \rightarrow 2Br\bullet$$

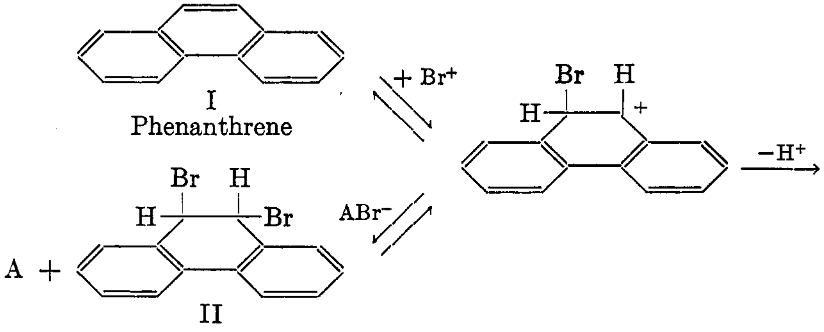


I
Phenanthrene

II

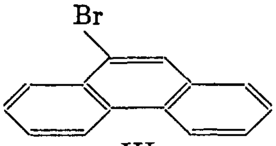
In the presence of a halogen carrier such as iodine or ferric chloride, the reaction of phenanthrene with bromine followed a different course (41). It proceeded smoothly in the dark and was accompanied by the evolution of hydrogen bromide and the formation of 9-bromophenanthrene. The effect of these catalysts could not have been simply the elimination of hydrogen bromide from the dibromide (II), since, under equivalent conditions, the rate of formation of 9-bromophenanthrene (III) from phenanthrene was very nearly the same as from the dibromide and both reactions proceeded at about the same rate as establishment of equilibrium for the addition reaction. These facts would indicate that the reactions were all proceeding through a common intermediate. This intermediate could not be the free radical of the photochemical addition, since iodine, an effective bromination catalyst, strongly inhibits the photochemical addition of bromine to phenanthrene. Since, however, iodine does slowly catalyze the addition reaction in the dark, the reaction in the latter case must be ionic rather than atomic. Allowing A to represent the halogen carrier (iodine, ferric chloride, antimony pentachloride, etc.), the catalytic reaction in the dark may be represented as follows:

$$A + Br_2 \rightleftharpoons Br^+(ABr)^-$$



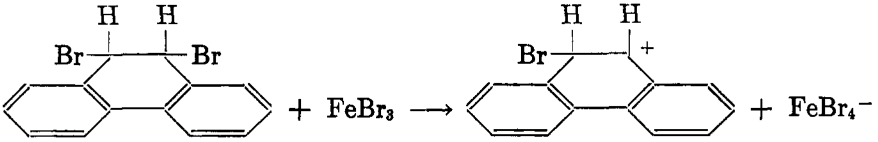
I
Phenanthrene

II



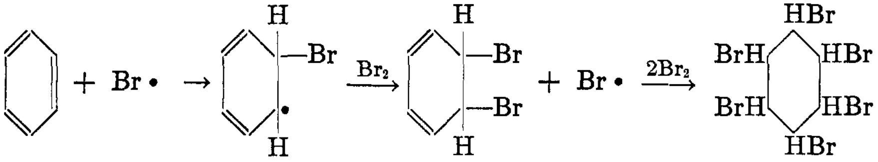
III
9-Bromophenanthrene

The initial step in the reaction from the dibromide (II) involves ionization of a carbon–bromine bond. Since the bond is to a secondary carbon atom alpha to a benzene ring, the reaction could certainly proceed by such a mechanism.[2] It has frequently been suggested that such reactions, which usually follow first-order kinetics in a polar solvent such as alcohol, are only pseudo-unimolecular and actually involve attack by the solvent molecule. The reaction cannot have been unimolecular in the case under consideration, in which a non-polar solvent, carbon tetrachloride, was used, since the pure dibromide was stable in this solvent, neither dissociating to phenanthrene and bromine nor decomposing to 9-bromophenanthrene and hydrogen bromide. It was only after the addition of a halogen carrier that both reactions occurred simultaneously. The initial step must be electrophilic attack by the catalyst on the bromine atom, a step similar to the polarization of a halogen molecule by the catalyst.



Of these three reactions of phenanthrene with bromine, i.e., the light-catalyzed atomic addition and the chemically-catalyzed ionic addition and substitution, two may be realized for benzene. It is a well-known fact that chlorine and bromine in an inert solvent will add to benzene photochemically to form the corresponding hexahalocyclohexanes (see, for example, 16 and 58).

$$Br_2 + h\nu \rightarrow 2Br\bullet$$



The catalytic substitution reaction with benzene must proceed just as with phenanthrene, except that the ionic intermediate evidently may be

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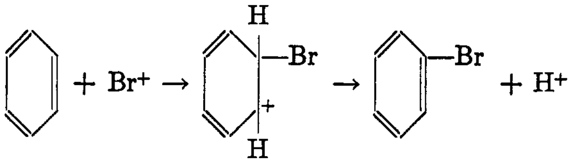
[2] For references and discussion see reference 14, pp. 131–183.

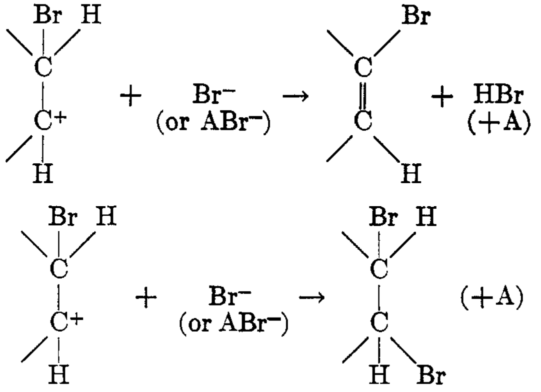
stabilized only by the loss of a proton, since no addition product is formed during substitution.

$$A + Br_2 \rightleftharpoons Br^+(ABr)^-$$



In fact, such elimination of a proton from this type of intermediate, leading to substitution, may be considered a characteristic of the aromatic nature of a substance. Olefinic character, on the other hand, is typified by the acceptance of an anion by such an intermediate, leading to addition. The double bonds of the benzene ring are thus wholly aromatic, while the 9,10-double bond of phenanthrene is only partially so.

Whether the reaction of ionic intermediates of this type will proceed by the acceptance of an anion, leading to addition, or by the loss of a proton, leading to substitution, must depend chiefly on the relative energy relationships for the two courses, i.e., the relative ease with which the bromide (or complex bromide) ion may act as a base and remove a proton from the active intermediate, acting as an acid. The basicity of the bromide ion will be largely dependent on the solvent medium, while the acidity of the intermediate will depend on its structure.



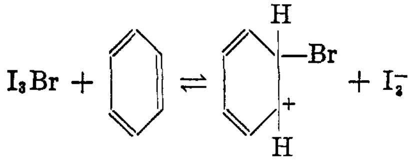
The identity of the mechanism of the iodine-catalyzed substitution in the two hydrocarbons, benzene and phenanthrene, is established by the kinetics of the two reactions since, in each case, the rate of reaction was found to be proportional to the three-halves power of the bromine concentration and to the five-halves power of the iodine concentration (44).

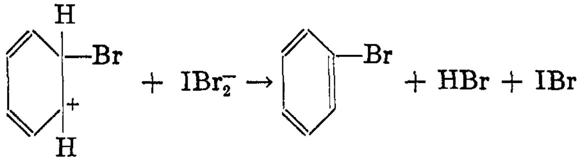
A reaction of such high order must be due to some sequence of reactions involving the two halogens. A series of reactions capable of accounting for the observed kinetics follows:

$$I_2 + Br_2 \rightleftharpoons 2IBr$$

$$IBr + I_2 \rightleftharpoons I_3Br$$

$$2IBr \rightleftharpoons IBr_2^- + I^+$$





If the rate-controlling step in this series of reactions is the last one, the rate of formation of bromobenzene may be expressed as

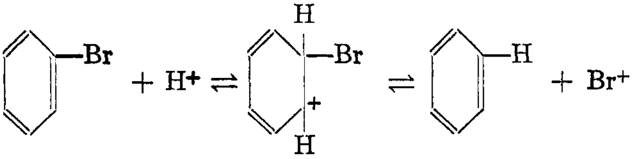
$$\frac{d[C_6H_5Br]}{dt} = k[C_6H_6Br^+][IBr_2^-]$$

Since $[C_6H_6Br^+]$ is proportional to $[I_3Br][C_6H_6]$, which in turn is proportional to $[I_2]^{3/2}[Br_2]^{1/2}[C_6H_6]$, and since $[IBr_2^-]$ is proportional to $[I_2][Br_2]$, the expression becomes

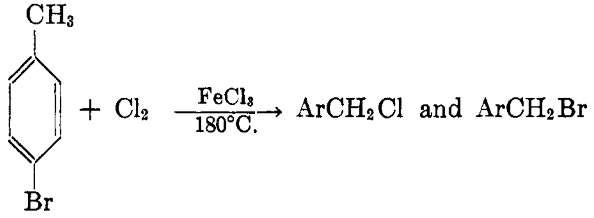
$$\frac{d[C_6H_5Br]}{dt} = k^2[C_6H_6][Br_2]^{3/2}[I_2]^{5/2}$$

The last step in the process is written as irreversible, since hydrogen bromide is not ionized in a non-polar solvent. In the presence of a strong acid, this step too becomes reversible. For instance, the removal and migration of bromine in certain instances of the Friedel–Crafts reaction (7) can be accounted for in the following manner:

$$AlCl_3 + HCl \rightarrow AlCl_4^-H^+$$



Since the bromine cation liberated in this step would be an active brominating agent, it might enter another molecule or a different position in the same molecule, either in the aromatic ring or in an aliphatic side chain. The wandering of bromine from the ring to the side chain, which has been observed during the chlorination of bromotoluenes (1), may be accounted for in a similar manner.

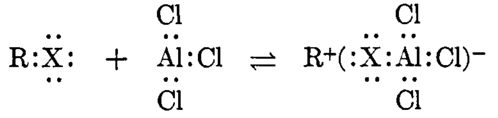


The second step in the course of such a replacement of bromine by hydrogen is the same as the second step for the reversal of the phenanthrene–bromine addition reaction in the presence of a polar catalyst such as iodine.

B. The Friedel–Crafts reaction

Another substitution reaction of aromatic compounds which has been studied extensively with regard to its mechanism is the Friedel–Crafts type of reaction, including acylation with acid chlorides, anhydrides, or esters, and alkylation with alkyl halides, alcohols, ethers, esters, or olefins. Such compounds as boron, aluminum, or iron halides, as well as sulfuric, phosphoric, or hydrofluoric acids, have served as the most useful catalysts for the condensations.

In these cases, also, the experimental evidence indicates an ionic type of mechanism. Wertyporoch and Firla (62) have clearly demonstrated by conductance studies the formation of an ionic complex between aluminum chloride and an alkyl halide.



$$\mathrm{R}:\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{X}}}: \;+\; \underset{\mathrm{Cl}}{\underset{\cdot\cdot}{\overset{\mathrm{Cl}}{\overset{\cdot\cdot}{\mathrm{Al}}}}}:\mathrm{Cl} \;\rightleftharpoons\; \mathrm{R}^+(:\underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{X}}}:\underset{\mathrm{Cl}}{\underset{\cdot\cdot}{\overset{\mathrm{Cl}}{\overset{\cdot\cdot}{\mathrm{Al}}}}}:\mathrm{Cl})^-$$

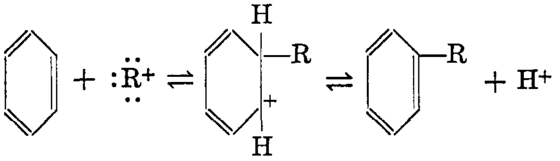
That catalysts for the Friedel–Crafts reaction actually do promote dissociation of the carbon–halogen bond in some manner is supported by an investigation of Bodendorf and Bohme (3). These authors found that optically active α-phenylethyl chloride ($C_6H_5CHClCH_3$) was racemized by the action of various metal salts which act as catalysts for the Friedel–Crafts type of reaction and that activity for the two reactions was roughly parallel:

$$\mathrm{HgCl} < \mathrm{ZnCl_2} < \mathrm{SnCl_4} < \mathrm{BCl_3} < \mathrm{SbCl_5}$$

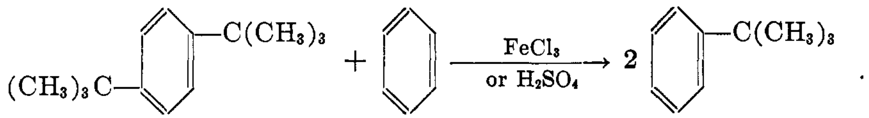
Lithium, silicon, tetramethylammonium, and arsenic chlorides, as well as hydrogen chloride, were ineffective as racemizing catalysts. Ulich and Heyne (56) have measured the equilibrium for the formation of certain of the catalyst–alkyl halide complexes postulated by Wertyporoch and, in addition, found that the rate of alkylation of benzene was directly proportional to the concentration of this catalyst–alkyl halide complex.

The reaction is then entirely analogous to halogenation, the electron-deficient carbonium ion (R^+), which may be either an alkyl or acyl group, reacting just as does the bromine cation (Br^+) to complete its octet by association with a pair of electrons from a double bond of the aromatic nucleus (9, 19, 45).

$$\mathrm{RCl} + \mathrm{AlCl_3} \rightleftharpoons :\!\underset{\cdot\cdot}{\ddot{\mathrm{R}}}{}^{+}(\mathrm{AlCl_4})^-$$



There have been many experimental verifications of the reversibility of the alkylation reaction. Ipatieff and Corson (21) have found that *p*-di-*tert*-butylbenzene in benzene solution is readily converted to *tert*-butylbenzene in the presence of such catalysts as ferric chloride, sulfuric acid, or phosphoric acid.



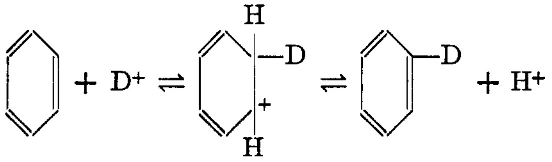
p-Di-*tert*-butylbenzene Benzene *tert*-Butylbenzene

Ipatieff and Corson found that these catalysts were effective only in removing tertiary alkyl groups. It has long been recognized, however, that more vigorous catalysis with aluminum chloride is effective for any type of alkyl group (4, 32; see also 68).

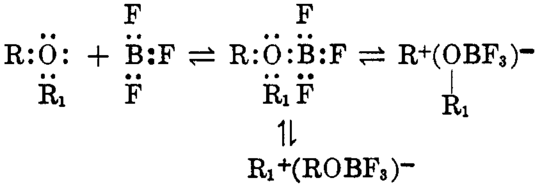
The relative ease of the two possible reactions for the active intermediate in the alkylation, i.e., loss of a proton to yield the alkylated derivative or loss of an alkyl cation to revert to the starting material, must depend on the tenacity with which the group R will retain the electrons it shares with the carbon atom of the ring. The fact that tertiary alkyl groups are most easily removed and primary least is in agreement with this prediction, since the order of electronegativity for alkyl groups has been found to be primary > secondary > tertiary (6, 64). Further evidence that this is clearly in accord with experimental facts will be discussed later under the consideration of orientation in the Friedel–Crafts reaction.

The mechanism for the reversibility of alkylation by the Friedel–Crafts reaction, as well as for the mobility of aromatic halogen under certain conditions, thus involves substitution by hydrogen through attack by a proton. A clear case of such substitution is the deuteration of benzene by deuterium chloride and aluminum chloride (25).

$$\mathrm{DCl} + \mathrm{AlCl_3} \rightarrow \mathrm{D^+AlCl_4^-}$$



The effect of the various catalysts on the oxygen-containing reagents, such as alcohols, ethers, esters, and anhydrides, is a polarization of the molecule similar to that exerted on the alkyl and acyl halides. For example, the general case of boron fluoride catalysis in such instances may be represented as follows:



$$\underset{\mathrm{R_1}}{\mathrm{R{:}\ddot{O}{:}}} + \overset{\mathrm{F}}{\underset{\mathrm{F}}{\mathrm{\ddot{B}{:}F}}} \rightleftharpoons \overset{\mathrm{F}}{\underset{\mathrm{R_1\ F}}{\mathrm{R{:}\ddot{O}{:}\ddot{B}{:}F}}} \rightleftharpoons \underset{\mathrm{R_1}}{\mathrm{R^+(OBF_3)^-}}$$

$$\rightleftharpoons \mathrm{R_1{}^+(ROBF_3)^-}$$

Complexes of this type in the case of ethers are well known; that of ether and boron fluoride, for example, is a stable liquid which boils at 126°C. with no decomposition. The complexes of ethers with boron chloride, however, decompose on warming to give alkyl chlorides and esters of boric acid (49).

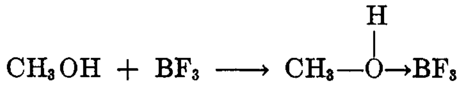
$$\mathrm{(CH_3)_2O + BCl_3 \longrightarrow (CH_3)_2O{\rightarrow}BCl_3}$$

$$\mathrm{(CH_3)_2O{\rightarrow}BCl_3 \longrightarrow CH_3Cl + CH_3OBCl_2}$$

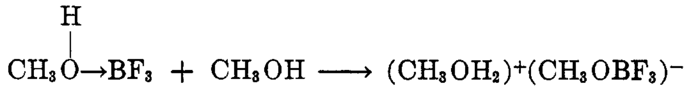
$$\mathrm{CH_3OBCl_2 + (CH_3)_2O \longrightarrow (CH_3)_2O{\rightarrow}BCl_2OCH_3}$$

$$\mathrm{(CH_3)_2O{\rightarrow}BCl_2OCH_3 \longrightarrow CH_3Cl + BCl(OCH_3)_2}, \text{ etc.}$$

Coördination with oxygen has also been suggested to account for the fact that the boron fluoride–alcohol complexes are strong acids. The association of the boron fluoride with the oxygen atom serves to promote the release of the proton (30, 38).



$$CH_3OH + BF_3 \longrightarrow CH_3-\overset{H}{\overset{|}{O}}\rightarrow BF_3$$



$$CH_3\overset{H}{\overset{|}{O}}\rightarrow BF_3 + CH_3OH \longrightarrow (CH_3OH_2)^+(CH_3OBF_3)^-$$

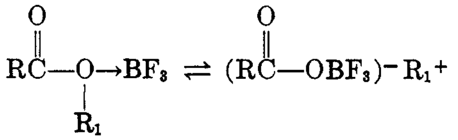
In the case of alcohols as well as ethers, the fluorides differ from the chlorides. While boron fluoride forms dialcoholates which, in the case of primary alcohols, are stable enough to be distilled (30), both boron and aluminum chlorides are vigorously decomposed by alcohols. In fact, the latter reaction has been recommended for the preparation of methyl, ethyl, and propyl chlorides (37).

$$AlCl_3 + CH_3OH \longrightarrow AlCl_2OCH_3 + HCl$$

$$AlCl_2OCH_3 \xrightarrow{100°C} AlOCl + CH_3Cl$$

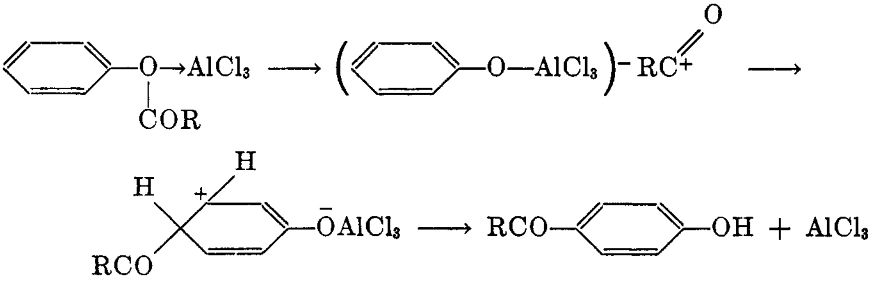
In the reactions of alcohols larger than propyl with aluminum chloride (37, 65) and of secondary alcohols with boron fluoride (46, 65), however, the principal product consists of polymeric material.

A particularly interesting case of the Friedel–Crafts reaction involving oxygen-containing compounds is that of esters. Since an alkyl group has considerably less affinity for electrons than an acyl group, it should be expected that ionization (or polarization) of the alkyl group as a cation would be predominant.



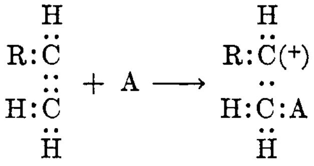
$$R\overset{O}{\overset{\|}{C}}-\underset{R_1}{\underset{|}{O}}\rightarrow BF_3 \rightleftharpoons (R\overset{O}{\overset{\|}{C}}-OBF_3)^-R_1^+$$

As would be expected on this basis, the reaction of esters in which R_1 is an alkyl group leads principally to alkylation (5, 29, 36). An aryl group, however, would be much less readily released as a cation than an alkyl group (6, 64). In the case of esters of phenols, therefore, it is to be expected that the acyl group might be liberated more readily. This may then recombine with a carbon in the phenol ring.



That the acyl group in the ester actually acts as a free acylating agent in the Fries rearrangement is indicated by the fact that Cox (8), in carrying out such a reaction in diphenyl ether, obtained the *p*-acyldiphenylether as well as the rearrangement product.

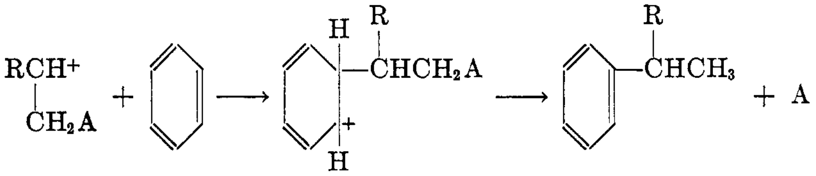
The condensation of olefins with aromatic compounds to yield alkyl derivatives has also been accounted for on the basis of a cationoid mechanism (19, 45). The electron-deficient catalyst in this case associates with a pair of electrons from the double bond of the olefin to give an active intermediate which is similar to that of the ordinary Friedel-Crafts reaction, in that one carbon atom has only a sextet of electrons.



$$\begin{matrix}\mathrm{H}\\ \mathrm{R}:\ddot{\mathrm{C}}\\ ::\\ \mathrm{H}:\underset{\cdot\cdot}{\mathrm{C}}\\ \mathrm{H}\end{matrix} + \mathrm{A} \longrightarrow \begin{matrix}\mathrm{H}\\ \mathrm{R}:\ddot{\mathrm{C}}(^+)\\ \cdot\cdot\\ \mathrm{H}:\mathrm{C}:\mathrm{A}\\ \mathrm{H}\end{matrix}$$

IV

This carbon atom may then complete its octet by association with an electron pair from the aromatic nucleus. The final step is elimination of the catalyst, followed by the α,γ-shift of a proton.

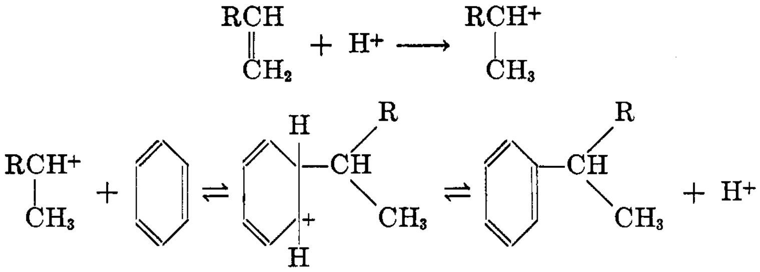


Completion of the octet of the electron-deficient carbon atom of IV by association with an electron pair of a second olefin molecule will, of course, result in polymerization (19).

A necessary corollary to such an intermediate complex formation for the olefins should be their *cis-trans* isomerization by these catalysts. This has been established for boron fluoride acting on *cis*-stilbene (47) and for the other Friedel-Crafts catalysts acting on diethyl maleate (11). In the latter case, the effectiveness of the catalysts in isomerization was closely parallel to that in condensation.

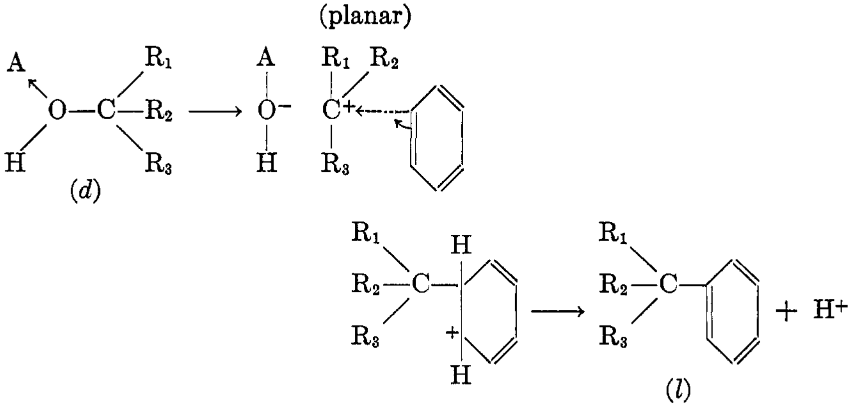
$$\begin{matrix}\mathrm{R{-}C{-}H}\\ \|\\ \mathrm{R{-}C{-}H}\end{matrix} + \mathrm{A} \rightleftharpoons \begin{matrix}\mathrm{RCH^+}\\ |\\ \mathrm{RCHA}\end{matrix} \rightleftharpoons \begin{matrix}\mathrm{R{-}C{-}H}\\ \|\\ \mathrm{H{-}C{-}R}\end{matrix} + \mathrm{A}$$

In the case of catalysis of any of these reactions by strong acids, such as sulfuric, phosphoric, or hydrofluoric acid or hydrogen chloride in the presence of aluminum chloride, the electron-deficient catalyst is evidently a proton.



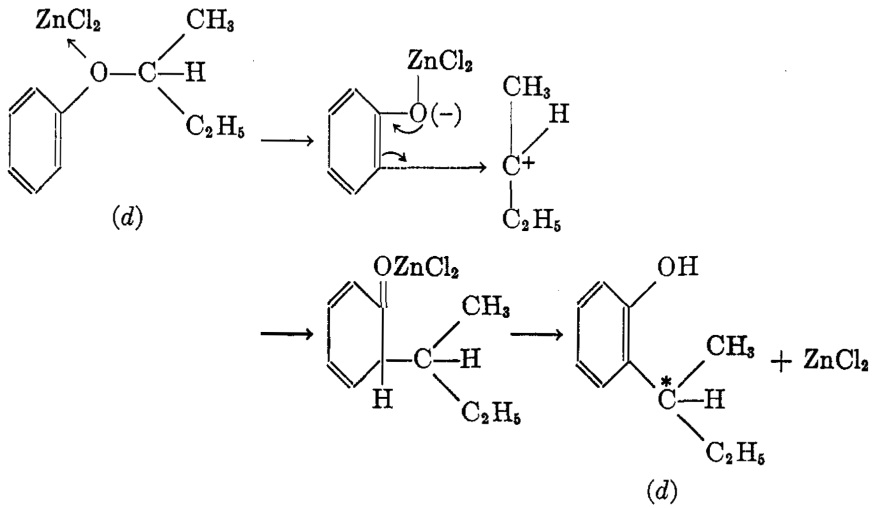
In accord with the views of Whitmore on molecular rearrangements (63), the electron-deficient carbon atom in the intermediate alkyl cation should tend to be secondary rather than primary, and tertiary rather than secondary. This tendency must account for the rearrangements of the alkyl group which are frequently encountered in Friedel–Crafts reactions (see, e.g., reference 22).

In those cases in which a possible rearrangement does not occur, such as alkylation using alcohols with aluminum chloride (20) or with boron fluoride (28) as catalysts, the existence of the alkyl cation may be too transient to allow for the migration of a proton (or an alkyl group), leading to a rearrangement product. This must certainly be true in the particular case of the boron-fluoride-catalyzed alkylation by means of alcohols, since *d*-*sec*-butyl alcohol was found to yield *l*-*sec*-butylbenzene (46). The rotation of the hydrocarbon ($[\gamma]_D^{20^\circ} = -0.16$) indicates not only an inversion of the configuration of the asymmetric carbon atom but also simultaneous racemization to the extent of about 95 per cent (27). This is the characteristic behavior of those replacement reactions at a saturated asymmetric carbon atom in which the rate-controlling step is the primary dissociation of the reagent into ions (reference 14, pages 131–183). If the asymmetric alkyl cation reacts almost simultaneously with the process of ionization, the only avenue of approach is at the face opposite that being vacated by the anion. The process will then result in almost complete inversion of configuration. The longer the life of the cation, however, the more

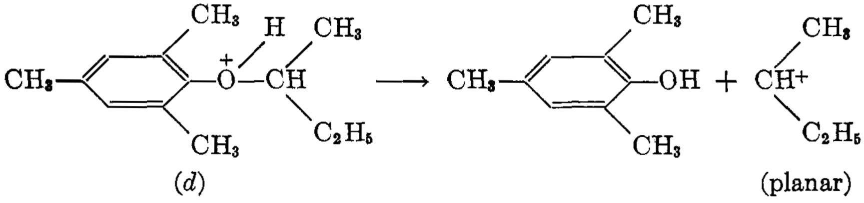


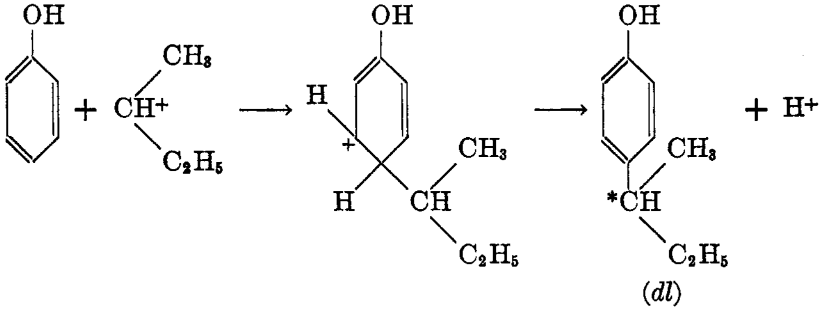
extensive will be its racemization before subsequent reaction. In the case of the alkylation of benzene, using an alcohol with boron fluoride as catalyst, the alkyl cation must have had a free enough existence to racemize considerably. It did, however, react so rapidly that inversion of configuration was a significant reaction.

It is of particular interest to compare these results of alkylation, using *d*-*sec*-butyl alcohol, with the results of Sprung and Wallis (55) on the rearrangement of *d*-*sec*-butyl aryl ethers. In each case the reaction involves scission of a carbon–oxygen bond, followed by union of the alkyl group thus liberated with the carbon atom of a benzene ring. In the case of the rearrangement, however, when the alkyl group moved to the ortho position in the same molecule, the reaction apparently proceeded with retention of configuration, accompanied by racemization. This may be readily accounted for on the basis that the planar intermediate alkyl cation never leaves the sphere of influence of the anion.



This course for the reaction is well substantiated by the fact that, when the alkyl group must react with a second molecule, as in the case of the alkylation of phenol with *d*-*sec*-butyl mesityl ether (12), the *sec*-butylphenol produced is racemic.





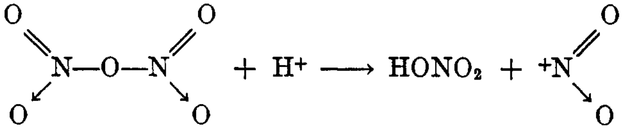
C. *Nitration and sulfonation*

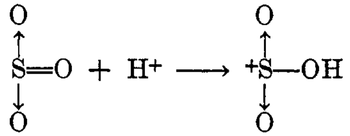
Although experimental data for the interpretation of the mechanism of nitration and sulfonation are not available, both reactions have been accounted for on the basis of a cationoid mechanism similar to that which has been established for the cases of halogenation and alkylation or acylation.[1] During the process, the nitric and sulfuric acids must dissociate in such a manner as to lose a hydroxyl group. It has been postulated that this occurs so as to leave the nitro and sulfonic acid groups as cations.

$$HONO_2 \rightarrow HO^- \; NO_2^+$$

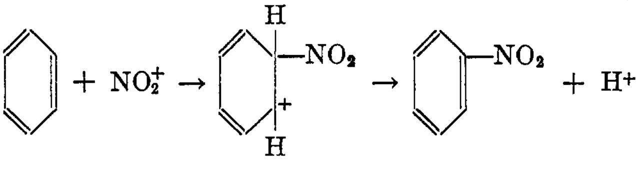
$$HOSO_2OH \rightarrow HO^- \; SO_2OH^+$$

These active intermediate cations may be formed by interaction of the respective acid anhydrides with a proton





The electron-deficient cations then react by picking up a pair of electrons from a double bond of the aromatic ring, and a proton is lost from the intermediate thus formed.



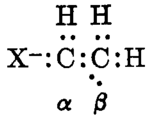
III. ORIENTATION

If the cationoid mechanism for substitution as presented above is to be considered entirely satisfactory, it should be capable of accounting for the influence of substituent groups in directing the entering group to certain preferred positions in the ring, as well as in affecting the ease with which the substitution can occur.

For those reactions proceeding by the cationoid mechanism,[3] the entering group in each case associates with a pair of electrons from a double bond of the aromatic nucleus. The substituent already present may therefore orient the incoming group by influencing the position of these electrons. The substituent may further influence the reaction (independently of its orienting effect) by conditioning the availability of these electrons, i.e. it may facilitate or hinder further substitution.

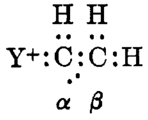
There are many manifestations of the effect of various substituent groups at a double bond on the position of the electrons of the bond. In particular, the mode of addition of most unsymmetrical reagents to a double bond is governed by the nature of the polarization of the bond imposed by the various neighboring groups.

If, for instance, the over-all effect of a group were negative, the electrons of the double bond, *particularly the extra labile pair,* would be repelled. This has been represented below in formula A, the α-carbon atom acquiring some measure of positive charge while the β-carbon atom becomes correspondingly negative. The second case (B), in which a group has a positive effect, will result in polarization of the double bond in the opposite sense.



$$X^-:\underset{\alpha}{\overset{\overset{H}{\cdot\cdot}}{C}}:\underset{\beta}{\underset{\cdot\cdot}{\overset{\overset{H}{\cdot\cdot}}{C}}}:H$$

A



$$Y^+:\underset{\alpha}{\underset{\cdot\cdot}{\overset{\overset{H}{\cdot\cdot}}{C}}}:\underset{\beta}{\overset{\overset{H}{\cdot\cdot}}{C}}:H$$

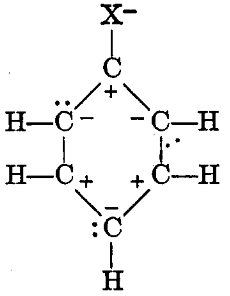
B

Addition of a reagent H^+A^- to these double bonds will then, of course, occur in such a manner that the negative portion of the adding reagent

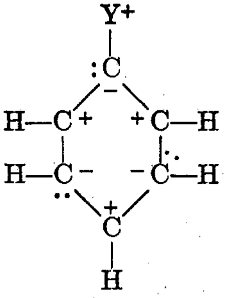
[3] There are a few reactions, proceeding by different mechanisms, which do not obey the ordinary rules for orientation. For those involving the reaction of an anion, such as the replacement of chlorine by hydroxyl ion (see Waters: *Physical Aspects of Organic Chemistry,* pp. 453–459, Routledge, London (1935)), the influence of substituent groups, both as to relative ease of replacement and relative position of activation, is opposite to that for the usual cationoid substitution reactions. This is as would be expected according to the mechanism for orientation outlined herein. Those reactions proceeding through free radical intermediates (Hey and Waters: Chem. Rev. **21**, 169 (1937)) lead chiefly to ortho and para derivatives, regardless of the nature of the substituent group present. Thus, on the basis of orientation, mercuration and metalation reactions may belong in the latter class.

will become attached to the more positively charged carbon atom. This is a statement of Michael's rule (31).

These two cases for a group attached to the conjugated system of double bonds in the aromatic nucleus may then be represented as follows (42):



(ortho–para directing)



(meta directing)

For each double bond, it is usually the polarity of the atom (or atoms) directly attached to the double bond which is the principal factor. This follows simply from Coulomb's law of electrostatic attraction and repulsion, the interaction of two charges decreasing in proportion to the square of the distance between them.[4]

Thus the group X, having a negative field with respect to the electrons of the ring, will shift the electrons in such a way as to increase the electron density at the ortho and para positions. Since the group which substitutes in the benzene ring is an electron-deficient group, tending to pick up a pair of electrons from the aromatic nucleus, it may obviously do so most readily at the ortho and para positions in this case. The group X will, therefore, be an ortho–para-directing group.

The group Y, with the opposite effect, will shift the electrons in the opposite direction, the electron density being greater in the meta positions. The entering group will thus tend to associate with the carbon atoms in the meta positions in this case. The group Y will therefore be a meta-directing group.

A. Ease of substitution

The ease of cationoid substitution in an aromatic compound will be governed by the ease of the initial attack involving acquisition of a pair of electrons from the ring. A substituent which will increase the electron density will increase the ease of substitution. One means of affecting the relative electron densities of the atoms in the ring must be the polarization of the bonds in the ring, induced by the substituent in the manner outlined above.

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[4] For the interatomic distances dealt with in this paper, the dielectric constant has been assigned the value of unity.

In addition, the substituent may evidently affect the over-all electron density of the ring by donating or withdrawing electrons. For example, the chlorine atom acquires a major share of the pair of electrons it shares with the ring, thus decreasing the over-all electron density of the ring. Chlorobenzene is thus more resistant to cationoid substitution than benzene.

Furthermore, it seems evident that the atom attached to the ring (referred to as the α-atom in table 2) may also share with the ring *a portion* of the charge which it acquires from some other atom or atoms. For example, the oxygen of the hydroxyl group will acquire a somewhat greater share in the electrons binding it to the ring than will the carbon atom to which it is bound. On this basis the hydroxyl group might resemble the chlorine atom and deactivate the ring to further substitution. The oxygen atom of the hydroxyl group, however, differs from the chlorine atom in that a major portion of its negative charge is acquired from the hydrogen atom. It apparently may share a portion of this negative charge acquired from the hydrogen with the ring, thus accounting for the fact that the hydroxyl group increases the ease of aromatic substitution.

Similarly, in the case of such groups as the carboxyl and trichloromethyl groups, further substitution in the ring must be hindered by the fact that a portion of the positive charge on the carbon atom in these two groups, acquired from the oxygen and chlorine atoms, respectively, is shared with the ring.

B. Rules of orientation

Many of the rules which have been suggested for predicting the orienting effect of a group substituted on the benzene ring, and particularly the more modern rules, such as those proposed by Hammick and Illingworth (17) and by Latimer and Porter (26), may actually be interpreted as empirical approximations for determining whether, in a substituent group, the atom which is directly attached to the ring is relatively positive or negative.

Since the over-all effect of many simple groups will be governed principally by the charge on this atom, which is closest to the electrons affected, the rules will correctly predict the orienting influence for these simple groups.

In order to determine whether the orienting influence of substituent groups could be more accurately correlated with their electrostatic effect, the charge on all the atoms of the substituent group should be taken into account, as well as their position relative to the electrons affected. This calculation has been carried out, using accurate molecular models, with charges assigned to the various atoms on the basis of bond moments determined from experimental dipole moment data (43). Many of these individual bond moments have already been estimated (53, 54). Others have been calculated in an analogous manner from more recent dipole moment data on saturated aliphatic derivatives containing the bond in question and from the respective atomic radii. These various individual bond moments are summarized in table 1.

TABLE 1

Individual bond moments and atomic charges in fractions of an electron

BOND	μ (BOND MOMENT)	d (BOND LENGTH)	Q (ATOMIC CHARGE) $= \mu \times 10^{-18}/d \times 10^{-8}$ E.S.U.	$Q/4.77 \times 10^{-10}$ = ATOMIC CHARGE IN FRACTIONS OF AN ELECTRON
	Debyes	Å.		
(+) (−)				
H—C	0.30	1.08	0.28×10^{-10}	0.06
C—O	0.85	1.43	0.60	0.13
C═O	2.40	1.24	1.93	0.40
C—S	0.95	1.81	0.53	0.11
C—N	0.40	1.47	0.27	0.06
C≡N	3.60	1.15	3.13	0.66
C—F*	1.60	1.44	1.11	0.23
C—F (3F)†	1.20	1.44	0.83	0.17
C—Cl	1.70	1.76	0.97	0.20
C—Cl (2 Cl)†	1.2	1.76	0.68	0.14
C—Cl (3 Cl)	0.85	1.76	0.48	0.10
C—Br	1.70	1.91	0.89	0.19
C—I	1.60	2.09	0.77	0.16
H—O	1.53	0.97	1.58	0.33
H—N	1.31	0.91	1.44	0.30
H—S	0.68	1.35	0.50	0.11
N→O	3.2	1.36	2.35	0.49
N═O	2.0	1.18	1.70	0.36
S—O	(0.25?)	1.70	0.15	0.03
S→O	2.5	1.70	1.47	0.31

* The values for the carbon-halogen bonds were calculated from primary halides, rather than from methyl halides.

† See Smyth and McAlpine: J. Chem. Phys. **1**, 190 (1933).

The charge on each atom associated with these bond moments was calculated simply by dividing the bond moment by the bond length. The basic charge distribution for the molecules was then assumed to be that derived from the various bonds for each atom. These charges, although undoubtedly modified by dipolar interaction, nevertheless formed a satisfactory basis for the calculations.

The effect of any particular atom on an electron of the adjacent double bond is then proportional to its charge divided by the square of its distance (r) from the electrons. That portion of this total force on the electrons effective in polarizing the double bond equals the total force multiplied by the cosine of the angle (θ) formed between the line connecting the electrons with the atom in question and the bond itself. This force, referred to as the polarizing force, is calculated for each particular atom (x) from the following expression,

$$\text{Polarizing force} = \frac{4.77 \times 10^{-10} \times e_x}{r^2} \cos \theta$$

in which 4.77×10^{-10} is the electronic charge in electrostatic units and e_x is the charge on the atom (x) in question.

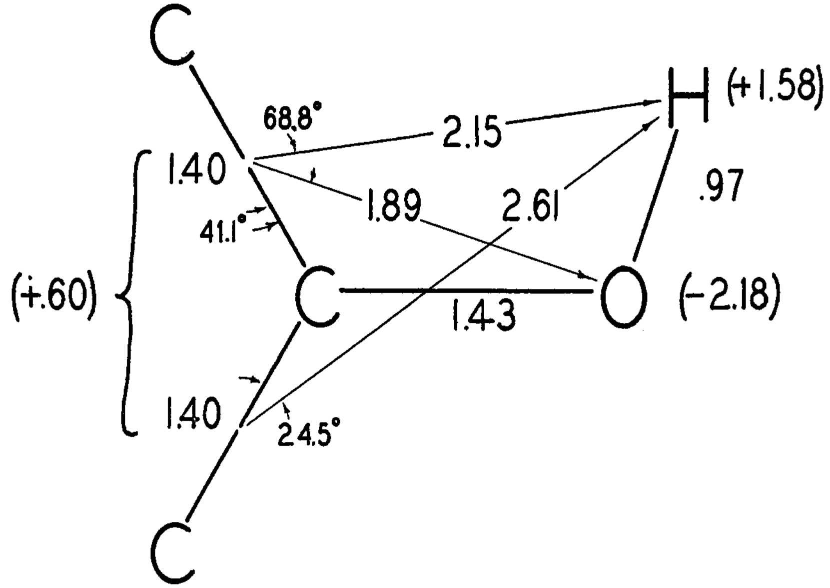

FIG. 1. Model for the hydroxyl group. Distances in Ångström units; charges in E.S.U. $\times 10^{10}$. Polarizing force = $[-2.18 \cos 41.1°/1.89^2 + (1.58 \cos 68.8°/2.15^2 + 1.58 \cos 24.5°/2.61^2)/2] \cdot 4.77 \times 10^{-4} = -1.46$ dynes.

In the calculations, free rotation of substituents[5] and resonance of the electrons in the aromatic nucleus have been assumed. When an atom could occupy several positions relative to the electrons of the double bond in the ring, the arithmetic mean of the polarizing force in the extreme

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[5] In several cases, such as that of the carbomethoxyl group, complete free rotation was impossible, owing to interference of the ortho hydrogen atoms. In such cases the extreme position of rotation was taken as that in which the groups approached each other as closely as the models would permit.

positions was used. Models for two particular cases, the hydroxyl group and the acetyl group, are illustrated in figures 1 and 2, respectively.

From such models the electrostatic effect of a large number of substituents of known orienting influence has been calculated. The correlation between these two factors, listed in table 2,[6] is so accurate that the degree of meta orientation may be estimated to within less than ±10 per cent from the value of the polarizing force.

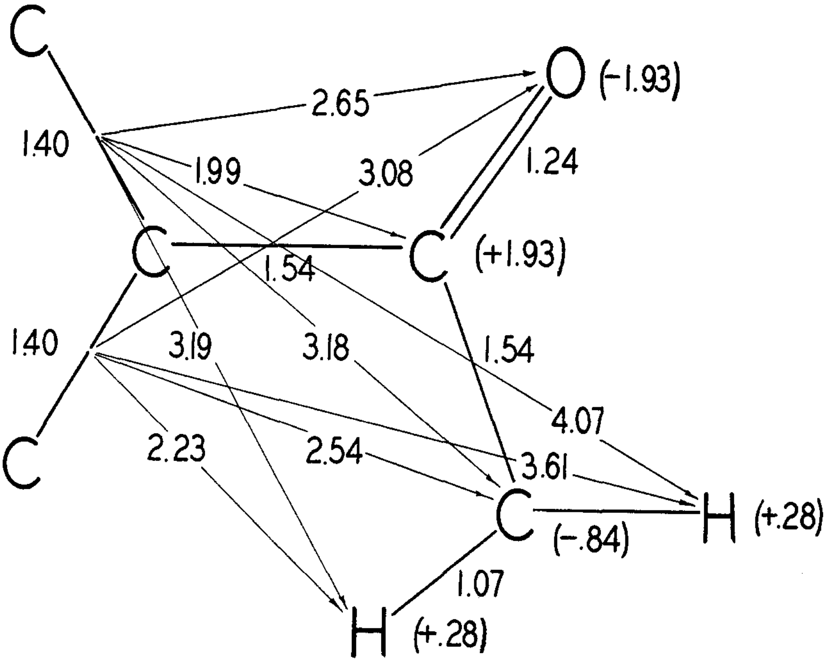

FIG. 2. Model for the acetyl group. Polarizing force $= [1.93 \cos 42.4°/1.99^2 + (-1.93 \cos 69.3°/2.65 - 1.93 \cos 27.7°/3.08^2)/2 + (-0.84 \cos 78.7°/2.54^2 - 0.84 \cos 19.9°/3.18^2)/2 + 3(0.28 \cos 103.3/2.23^2 + 0.28 \cos 1.0°/3.19^2 + 0.28 \cos 72.8°/3.61^2 + 0.28 \cos 29.8°/4.07^2)/4]\cdot 4.77 \times 10^{-4} = 0.89$ dyne.

It should be noted that the division between the ortho–para-directing and the meta-directing groups falls, as it should, at the value for the hydrogen atom; i.e., a group more negative than hydrogen is predominantly ortho–para directing, while one more positive is meta directing.

No substituents containing an olefinic double bond have been included in table 2, since no method is available for calculating the absolute degree of polarization, and thus the charge distribution, for such bonds. Fur-

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[6] The data listed in table 2 should be useful also for comparing various groups in regard to their polarizing effect on a simple olefinic double bond.

TABLE 2

The electrostatic polarizing influence of various substituent groups on an adjacent double bond, compared with their orienting influence

GROUP	CHARGE ON α-ATOM	POLARIZING FORCE	OBSERVED META ORIENTATION(a)
	E.S.U.	*dynes*	*per cent*
NH_2	-3.15×10^{-10}	-1.66×10^{-4}	
OH	−2.18	−1.46	3
$NHCH_3$	−1.98	−1.41	
OCH_3	−1.20	−1.21	
$N(CH_3)_2$	−0.81	−1.16	
F	−1.11	−1.12	
Cl	−0.97	−0.68	
$C(CH_3)_3$	0	−0.62	
Br	−0.89	−0.55	
I	−0.77	−0.39	
CH_3	−0.84	−0.39	4
COO^-	+2.53	−0.07	33(b)
CH_2CN	−0.56	+0.32	14(c)
CH_2Cl	+0.41	0.36	4 (11(d))
CH_2Br	0.33	0.36	7
CH_2F	0.55	0.39	17
H	0.28	0.45	(40)
$C(CH_3)_2NO_2$	0.27	0.52	30(e)
CH_2NO_2	−0.29	0.68	50
$CHCl_2$	+1.08	0.75	34
$N{=}O$	1.43	0.77	(f)
$CH_2NH_3^+$	−0.29	0.80	49
$CONH_2$	+2.20	0.84	69
$COCH_3$	1.93	0.89	55
CCl_3	1.44	0.93	64 (100(g))
CHO	1.65	0.98	79
$CH(CN)_2$	−0.28	1.05	68(c)
$CH_2N(CH_3)_3^+$	−0.29	1.07	88
$COOCH_3$	+2.53	1.23	73
COOH	2.53	1.25	82 (90(h))
SO_2CH_3	1.88	1.36	98(i)
CF_3	2.49	1.41	90(j)
COCl	2.90	1.73	90
$C{\equiv}N$	3.13	1.80	80 (88(c))
SO_3H	2.56	1.87	72 (95(k))
NO_2	3.78	2.21	93
NH_3^+	0.18	2.21	(95(l))
$N(CH_3)_3^+$	3.69	3.21	100

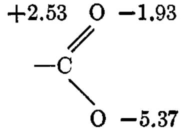
(a)Unless otherwise noted, these values have been taken from those given by L. F. Fieser in *Organic Chemistry* (edited by H. Gilman), p. 140, John Wiley and Sons, thermore, since the olefinic bond is a reactive center in the molecule, it may well be modified considerably while in contact with a reagent reactive enough to substitute in the aromatic ring.

The values for the percentage of meta orientation are principally those for nitration (footnote 1 (page 37) and reference 50). The more recent data are referred to in the table. These values may vary somewhat under differing experimental conditions or for different reagents. Significantly different values have been indicated in the table.

Although the data presented do not lead to any indication of the actual degree of polarization of the bonds in the aromatic ring, it is significant to compare the electrostatic polarizing forces in table 2 with two recent estimations of the effect of substituents on the charge distribution in the benzene ring (15, 51). If this force is the principal factor involved in polarizing the bonds of the benzene ring, then the difference in charge between adjacent atoms of the ring, such as the meta and para carbon atoms in a benzene derivative, should parallel the electrostatic polarizing force of the substituent.

Inc., New York (1938), and are principally for nitration.

[b] Smith: J. Chem. Soc. **1934**, 213. The calculation of the polarizing force was based on the ion as indicated with the following charge distribution:



Any hydration of this ion by hydrogen bonding with the solvent (water) would increase the value of the polarizing effect and thus bring this group into agreement with the others in the table.

[c] Flürscheim and Holmes: J. Chem. Soc. **1928**, 2230.

[d] Flürscheim and Holmes: J. Chem. Soc. **1928**, 1607.

[e] Baker and Ingold: J. Chem. Soc. **1926**, 2462.

[f] Hammick, New, and Sutton (J. Chem. Soc. **1932**, 742) have suggested that the reason the nitroso group directs ortho-para, contrary to predictions, is that nitrosobenzene actually exists in solution as a dimer.

[g] Lauer: J. prakt. Chem. **142**, 252 (1935).

[h] Reese: Chem. Rev. **14**, 55 (1934).

[i] Baldwin and Robinson: J. Chem. Soc. **1932**, 1445.

[j] Finger, Nachtrieb, and Reed: Trans. Illinois State Acad. Sci. **31**, 132 (1939).

[k] Lauer: J. prakt. Chem. **143**, 127 (1935).

[l] Since aniline is a weak base, the ortho and para derivatives frequently obtained from aniline salts may arise from the reaction of free aniline. This view is supported by the fact that increased acid concentration always results in increased meta orientation in aniline and its derivatives (Noelting and Collin: Ber. **17**, 261 (1884); Clemo and Smith: J. Chem. Soc. **1928**, 2414).

Hammett (15) has found that, for a large variety of reactions of benzene derivatives, the effect of a substituent in the meta or para position on the rate (or equilibrium) of the reaction may be expressed as

$$\ln k - \ln k_0 = \rho\sigma \qquad (1)$$

where k is the rate constant for the meta- or para-substituted derivative, k_0 is that for the unsubstituted compound, and ρ and σ are two constants. The reaction constant ρ depends solely on the nature of the reaction under consideration, e.g., the hydrolysis of an ester, esterification, the ionization of an acid, etc., whereas the substituent constant σ is dependent solely on the nature of the substituent in the meta or para position. The substituent constant is interpreted by Hammett as a measure of the electron density produced by a substituent at the reactive group, while ρ measures the susceptibility of a particular reaction of this group to changes in electron density.

TABLE 3

Comparison of the difference in the substituent constant σ for a substituent in the meta and para positions with the polarizing force of the substituent (table 2)

GROUP	σ_m*	σ_p	$\Delta\sigma$	POLARIZING FORCE
NH_2	−0.161	−0.660	−0.499	-1.66×10^{-4}
OCH_3	+0.115	−0.268	−0.383	−1.21
F	0.337	+0.062	−0.275	−1.12
Cl	0.373	0.227	−0.146	−0.68
Br	0.381	0.232	−0.149	−0.55
CH_3	−0.069	−0.170	−0.101	−0.39
I	+0.352	+0.276	−0.076	−0.39
$COCH_3$	0.306	0.874	+0.568†	+0.89
CHO	0.381	1.126	0.745†	0.98
COOH	0.355	0.728‡	0.373	1.25
$C{\equiv}N$	0.678	1.000‡	0.322	1.80
NO_2	0.710	1.27‡	0.56	2.21

* The values for σ are taken from reference 14, page 188.

† The only determinations from which these data were calculated were the rates of hydrolysis of the corresponding sulfonic acids.

‡ These data are for phenol derivatives. The evidence indicates that these values may be different for other types of compounds (15).

A comparison of the difference in values of σ for different substituent groups in the meta and para positions is presented in table 3, accompanied by the polarizing force of the substituents. The close parallel between the polarization as measured by σ and the polarizing force would indicate strongly that these two values must be directly connected.

Ri and Eyring (51) have estimated the charges on each nuclear carbon atom of benzene derivatives by analysis of the data in the literature for the competitive rates of nitration of various derivatives of benzene. They based their calculation on the assumptions that nitration proceeds by a cationoid mechanism and that the interaction between the nitro cation (NO_2^+) and the charge on the carbon atom with which it reacts changes the free energy of activation by the amount $e_y e_n/rD$, where e_y is the charge on the carbon atom undergoing nitration, e_n the charge on the nitro group, r the distance between the two charges in the intermediate activated complex, and D the dielectric constant of the medium separating the charges. The distance r was taken as 1.6Å., a value 10 per cent longer than the normal carbon–nitrogen bond. At this distance there can be no molecules separating the charges, so the dielectric constant is considered to be unity. The rates of substitution at a carbon atom in benzene (k_{H}) and at a carbon atom y in a benzene derivative (k_y) will then be

$$k_{\mathrm{H}} = \alpha e^{-\Delta F/RT} \tag{2}$$

$$k_y = \alpha e^{-(\Delta F + e_y e_n/rD)/RT} \tag{3}$$

from which

$$k_y/k_{\mathrm{H}} = e^{-e_y e_n/rDRT} \tag{4}$$

From the rates of nitration of a benzene derivative in each position relative to the rate for benzene itself, Eyring and Ri were thus able to estimate the charge on each atom in the ring by use of equation 4. The differences between this charge (e_y) for the positions meta and para to the various substituents are summarized in table 4, again compared with the electrostatic polarizing force of the substituent group. The agreement is even more satisfactory than for the values of σ (table 3).

It is of some interest to point out the theoretical relationship and implications of the work of Hammett and of Ri and Eyring as well as that of Jones (23). The latter investigator has found that, in the rate expression

$$k = \alpha e^{-E/RT}$$

for the halogenation of a large variety of aromatic derivatives, particularly of the type p-XC_6H_4OR, α remained constant while the effect of the various substituents on the activation energy E was additive. Expressing the rate of chlorination in the same manner as that of nitration (equation 3), the activation energy, E, may be represented as

$$E = \Delta F + e_y e_x/rD \tag{5}$$

where ΔF represents a constant, e_y the charge on a carbon atom y of the nucleus, e_x the charge on the entering substituent, in this case a chlorine cation, Cl^+, and r and D represent the distance and dielectric constant as before. If the effect of a substituent is principally on the charge, e_y, and if this influence for two groups in the same ring is substantially independent, then the effect of the two substituents on the activation energy for the reaction will be additive.

TABLE 4

Comparison of the differences between the charges on the meta and para carbon atoms, as calculated by Ri and Eyring (51), with the polarizing force (table 2)

GROUP	e_{ym}	e_{yp}	Δe_y	POLARIZING FORCE
	E.S.U. × *10*^{10}	*E.S.U.* × *10*^{10}	*E.S.U.* × *10*^{10}	
F	0.0933	0.0035	−0.0898	-1.12×10^{-4}
Cl	0.115	0.027	−0.088	−0.68
Br	0.116	0.0305	−0.0855	−0.55
I	0.0909	0.0065	−0.0844	−0.39
CH_3	−0.0128	−0.0532	−0.0404	−0.39
CH_2Cl	0.114	0.070	−0.044	+0.36
$CHCl_2$	0.117	0.104	−0.013	0.75
CCl_3	0.0876	0.0892	+0.0016	0.93
$COOCH_3$	0.0929	0.1230	+0.030	1.23
COOH	0.0411	0.0871	+0.046	1.25
NO_2	0.191	0.248	+0.157	2.21

$$E = \Delta F + \frac{(e_{y_1} + e_{y_2})e_x}{rD} \tag{6}$$

In this expression, e_{y_1} represents the charge induced on the carbon atom y by one substituent, and e_{y_2} that induced by a second substituent. The additivity of the effects of different groups may thus be adequately accounted for by equation 6.

The constancy of the probability factor[7] α, observed experimentally by Jones and implicit in the derivation of equation 4 from equations 2 and 3 by Ri and Eyring, leads to a direct interpretation of the ρ and σ constants of Hammett (equation 1). If we extend equation 2 to include the possibility that, in the benzene derivative under consideration (C_6H_5R), the carbon atom to which the reactive group R is attached may have a charge e_R, the equation can be represented as

$$k_H = \alpha e^{-(\Delta F^0 + e_R e_x / rD)/RT} \tag{7}$$

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[7] Hammett (14) has presented a discussion of the conditions necessary for the constancy of α, considered from the viewpoint of the transition-state theory.

A second substituent, y, in this compound will change the value of e_{R} to e_y.

$$k_y = \alpha e^{-(\Delta F^0 + e_y e_x / rD)/RT} \tag{8}$$

or

$$\ln k_{\mathrm{H}} - \ln \alpha = -\ (\Delta F^0/RT + e_{\mathrm{R}} e_x / rD \cdot RT) \tag{9}$$

and

$$\ln k_y - \ln \alpha = -\ (\Delta F^0/RT + e_y e_x / rD \cdot RT) \tag{10}$$

Subtracting equation 9 from equation 10, we find

$$\ln k_y - \ln k_{\mathrm{H}} = -(e_y - e_{\mathrm{R}}) e_x / rD \cdot RT \tag{11}$$

The term[8] $(e_y - e_{\mathrm{R}})$ is the change in electron density induced on the carbon atom of the benzene adjacent to R by the substituent and is independent of the reaction; thus it corresponds to σ. The term e_x/rD, the charge on the attacking reagent divided by its separation from the point of attack in the activated complex, is dependent solely on the nature of the reaction and thus corresponds to the ρ factor.

The factor $1/T$ is absorbed in the ρ factor. The only case for which ρ has been determined at two different temperatures is for the reaction of aniline derivatives with 2,4-dinitrochlorobenzene at 25° and 100°C. At the lower temperature, $\rho = -3.1910 \pm 0.07$; at the higher, $\rho = -2.581 \pm 0.16$. Multiplying by the absolute temperature, the first value becomes -950 ± 20 and the second -962 ± 60. A much more thorough test of the inverse proportionality of ρ and the absolute temperature is necessary.

The fact that equation 1 fails for cases in which the substituent is ortho to the reactive group may be accounted for very simply by the failure of the probability or steric factor α to remain constant, although a direct polar influence of the substituent on the reactive group, rather than one passed through the ring, is also undoubtedly of considerable importance.

The wide variety of reactions for which equation 1 may be applied implies the constancy of both α and ΔF^0 for all these reactions. If ΔF^0 actually does remain constant for all these diverse reactions, it can of course be included in α, and the rate expression becomes

$$k_y = \alpha_1 e^{-e_y e_x / rDRT} \tag{12}$$

In other words, the activation energy for this group of reactions may be determined solely by an expression involving the interaction between the electron density of the reactive group (e_y) and the electron density of the reagent (e_x) at the distance r which separates them in the activated complex.

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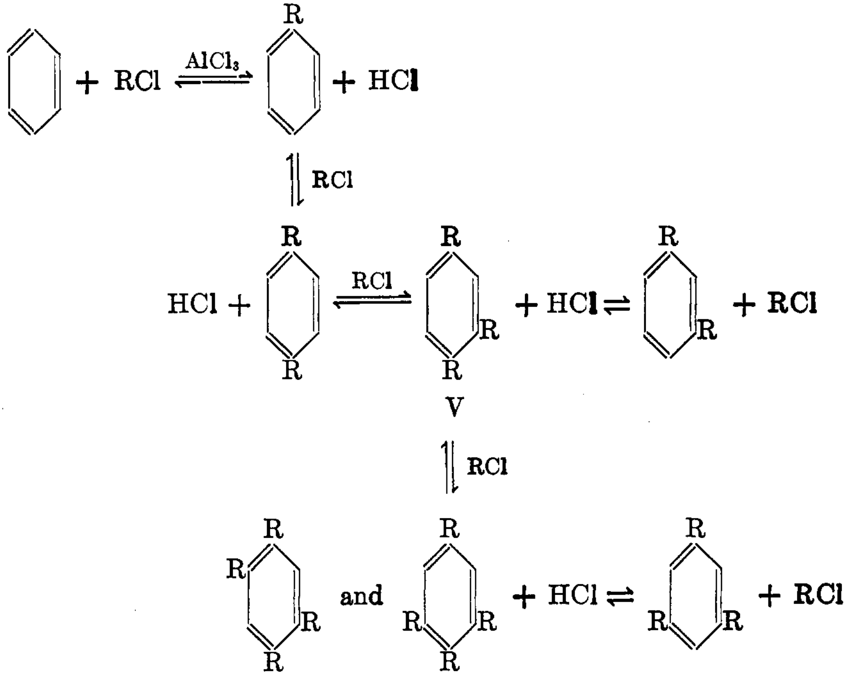
[8] It must be this *difference* in charge between a carbon atom in benzene itself and in a derivative which was calculated by Eyring and Ri, rather than the absolute value of the charge.

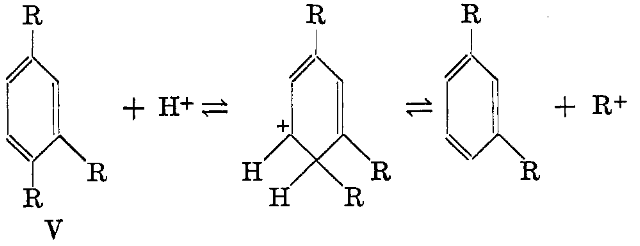
C. Anomalous orientation

One case of orientation which frequently gives results which have been considered anomalous is that of alkylation by the Friedel–Crafts reaction. In many cases of the alkylation of benzene, for instance, using an alkyl halide with aluminum chloride as catalyst, the dialkylated material consists chiefly of the meta isomer, whereas the first alkyl group should orient an incoming group to the ortho and para positions. It has been suggested (45) that this fact is due to the reversibility of the reaction.



Nightingale and Smith (35), in studying the action of aluminum chloride in rearranging compounds of the type V to the symmetrical isomer, have, in fact, isolated substantial amounts of the *m*-dialkyl compound. This replacement of the alkyl group in the 4-position by hydrogen would take place according to the following steps:

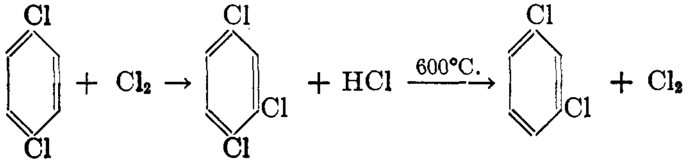
$$HCl + AlCl_3 \rightleftharpoons H^+ (AlCl_4)^-$$



The ease with which the R group will be replaced by hydrogen in this manner will be conditioned by its affinity for the pair of electrons binding it to the carbon atom of the benzene ring. This is indeed true, since a methyl group in this position is much more stable than other alkyl radicals, and *n*-propyl is more so than isopropyl (34).

It is thus apparent that an alkylation performed in such a manner that the reaction is reversible may lead to *m*-dialkyl derivatives, whereas the irreversible alkylations will yield the expected *o*- and *p*-dialkylbenzenes.

This same explanation may be applied to the results of Wibaut and van Loon (59) on the halogenation of benzene at high temperatures. For example, benzene reacts with chlorine above 600°C. to yield *m*-dichlorobenzene rather than the normal *o*- and *p*-isomers. The migration of aromatically bound halogen atoms under certain conditions has already been mentioned; evidently elevated temperatures in the presence of hydrogen chloride may lead to the same result.



IV. SUMMARY

Investigations of the usual substitution reactions in benzenoid compounds involving the replacement of a hydrogen atom, particularly halogenation and the Friedel–Crafts type of reaction, support a mechanism involving attack by an electron-deficient cationoid reagent.

On the basis of such a mechanism, the orientation of an entering group may be effected by the polarization of the double bonds of the benzene ring. This polarization may be accomplished by the electrostatic attraction or repulsion of a substituent group for the electrons of the adjacent double bond in the ring. Since the electrostatic polarizing forces for a large variety of substituents, calculated from dipole moment data and interatomic distances, closely parallel the observed orienting influences of the groups, this force is apparently the principal factor governing orientation.

The polarizing force for a hydrogen atom was estimated to be 0.45 dyne. Groups having a smaller but positive polarizing force are predominantly ortho–para directing, although significant amounts of the *m*-isomeride are usually obtained; groups having a negative (electron-repelling) polarizing force are almost exclusively ortho-para directing. Groups having a polarizing force larger (more electron-attracting) than that of hydrogen (0.45 dyne) but less than 1.0 dyne are principally meta directing, although significant amounts of the *o*- and *p*-isomerides are formed; groups having a polarizing force greater than 1.0 dyne are increasingly powerfully meta directing, usually yielding 90 per cent or more of the *m*-isomeride.

REFERENCES

(1) Asinger: Monatsh. **64,** 153 (1934).
(2) Berthoud and Beraneck: J. chim. phys. **24,** 213 (1927).
(3) Bodendorf and Bohme: Ann. **516,** 1 (1935).
(4) Boedtker and Halse: Bull. soc. chim. [4] **19,** 444 (1916).
(5) Bowden: J. Am. Chem. Soc. **60,** 645 (1938).
(6) Brown: J. Am. Chem. Soc. **61,** 1483 (1939).
(7) Copisarow: J. Chem. Soc. **119,** 442 (1921).
(8) Cox: J. Am. Chem. Soc. **52,** 352 (1930).
(9) Dilthey: Ber. **71,** 1350 (1938).
(10) Fieser and Price: J. Am. Chem. Soc. **58,** 1838 (1936).
(11) Gilbert, Turkevich, and Wallis: J. Org. Chem. **3,** 611 (1939).
(12) Gilbert and Wallis: J. Org. Chem. **5,** 184 (1940).
(13) Goldwasser and Taylor: J. Am. Chem. Soc. **61,** 1260 (1939).
(14) Hammett: *Physical Organic Chemistry,* pp. 112–131. McGraw-Hill Book Company, New York (1940).
(15) Reference 14, p. 188.
(16) Hammick, Hutson, and Jenkins: J. Chem. Soc. **1938,** 1959.
(17) Hammick and Illingworth: J. Chem. Soc. **1930,** 2358.
(18) Holleman: Rec. trav. chim. **42,** 356 (1923).
(19) Hunter and Yohe: J. Am. Chem. Soc. **55,** 1248 (1933).
(20) Huston and Hsieh: J. Am. Chem. Soc. **58,** 439 (1936).
(21) Ipatieff and Corson: J. Am. Chem. Soc. **59,** 1417 (1937).
(22) Ipatieff, Pines, and Schmerling: J. Org. Chem. **5,** 253 (1940).
(23) Jones: J. Chem. Soc. **1935,** 1831; **1936,** 1854.
(24) Kharasch, White, and Mayo: J. Org. Chem. **2,** 574 (1938).
(25) Klit and Langseth: Z. physik. Chem. **A176,** 65 (1936).
(26) Latimer and Porter: J. Am. Chem. Soc. **52,** 206 (1930).
(27) Marker: J. Am. Chem. Soc. **58,** 976 (1936).
(28) McKenna and Sowa: J. Am. Chem. Soc. **59,** 470 (1937).
(29) McKenna and Sowa: J. Am. Chem. Soc. **59,** 1204 (1937).
(30) Meerwein and Pannwitz: J. prakt. Chem. **141,** 123 (1934).
(31) Michael: J. Am. Chem. Soc. **32,** 1005 (1910).
(32) Milligan and Reid: J. Am. Chem. Soc. **44,** 208 (1922).
(33) Muller and Schumacher: Z. physik. Chem. **B42,** 327 (1939).
(34) Nightingale: Chem. Rev. **25,** 329 (1939).
(35) Nightingale and Smith: J. Am. Chem. Soc. **61,** 101 (1939).
(36) Norris and Arthur: J. Am. Chem. Soc. **62,** 874 (1940).
(37) Norris and Sturgis: J. Am. Chem. Soc. **61,** 1413 (1939).
(38) O'Leary and Wenzke: J. Am. Chem. Soc. **55,** 2117 (1933).
(39) Pfeiffer and Wizinger: Ann. **461,** 132 (1928).
(40) Price: J. Am. Chem. Soc. **58,** 1834 (1936).
(41) Price: J. Am. Chem. Soc. **58,** 2101 (1936).
(42) Price: Abstracts of the 98th Meeting of the American Chemical Society, Boston, September, 1939.

(43) PRICE: Abstracts of the 101st Meeting of the American Chemical Society, St. Louis, April, 1941.
(44) PRICE AND ARNTZEN: J. Am. Chem. Soc. **60,** 2835 (1938).
(45) PRICE AND CISKOWSKI: J. Am. Chem. Soc. **60,** 2499 (1938).
(46) PRICE AND LUND: J. Am. Chem. Soc. **62,** 3105 (1940).
(47) PRICE AND MEISTER: J. Am. Chem. Soc. **61,** 1595 (1939).
(48) PRINS: Rec. trav. chim. **43,** 685 (1924).
(49) RAMSER AND WIBERG: Ber. **63B,** 1136 (1930).
(50) REESE: Chem. Rev. **14,** 55 (1934).
(51) RI AND EYRING: J. Chem. Phys. **8,** 433 (1940).
(52) SCHUMACHER: J. Am. Chem. Soc. **52,** 3132 (1930).
(53) SIDGWICK: *The Covalent Link in Chemistry,* pp. 153–156. Cornell University Press, Ithaca, New York (1933).
(54) SMYTH AND MCALPINE: J. Chem. Phys. **1,** 190 (1933).
(55) SPRUNG AND WALLIS: J. Am. Chem. Soc. **56,** 1715 (1934).
(56) ULICH AND HEYNE: Z. Elektrochem. **41,** 509 (1935).
(57) VAN ALPHEN: Rec. trav. chim. **47,** 169 (1928).
(58) VAN DER LINDEN: Rec. trav. chim. **57,** 1075 (1938).
(59) VAN LOON AND WIBAUT: Nature **139,** 151 (1937); Rev. trav. chim. **56,** 815 (1937).
(60) VINCENT, THOMPSON, AND SMITH: J. Org. Chem. **3,** 603 (1939).
(61) WACHHOLTZ: Z. physik. Chem. **125,** 1 (1927).
(62) WERTYPOROCH AND FIRLA: Ann. **500,** 287 (1933).
(63) WHITMORE: J. Am. Chem. Soc. **54,** 3274 (1932).
(64) WHITMORE AND BERNSTEIN: J. Am. Chem. Soc. **60,** 2626 (1938).
(65) WHITMORE AND LAUCIUS: J. Am. Chem. Soc. **61,** 973 (1939).
(66) WILLSTÄTTER AND HEIDELBERGER: Ber. **46,** 517 (1913).
(67) WILLSTÄTTER AND WASER: Ber. **44,** 3423 (1910).
(68) WOODWARD, BORCHERDT, AND FUSON: J. Am. Chem. Soc. **56,** 2103 (1934).