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JOURNAL REVIEW

Molecular Thermodynamics for Chemical Reaction Design

Various techniques of molecular thermodynamics and solution theory have been applied, through the transition state theory, to problems in chemical kinetics. These include experimental techniques for ascertaining the structure and properties of the transition states for various classes of reactions, plus analytical methods for using this information to predict the effect of solvents or reactant structure on chemical reaction rates. Such methods provide the chemical engineer with an additional parameter to use in the optimal design of reaction systems—that is, the chemistry of the reaction itself.

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One of the foremost aspects of chemical engineering practice is the design of processes involving chemical reaction steps. The reaction design engineer is well accustomed to considering the important mathematical and physical parameters in such a process, but the purpose of this review is to stress that the chemical parameters—although frequently neglected—may often be at least as important. A number of examples are now available where a knowledge of the molecular chemistry of the reaction has been applied to a rational design of the reaction system.

Such process improvements and optimization have most often been implemented by using transition state theory to reduce the rate problem to an equilibrium problem, tractable by the many approaches of thermodynamics. Often this involves the use of a thermodynamic method to characterize the mechanism of the reaction or the properties of the transition state, as from a determination of the activation parameters. An outstanding example of the application of such information is in the area of solvent selection and design. Commonly, solvent effects on chemical reactions are a factor of several orders of magnitude in the rate, and cases have been reported with factors as high as 10^9 (Brown and Okamoto, 1958; Cram et al., 1961). Thus a proper choice of a reaction solvent can have a very large effect on reaction rates, and molecular solution theories have been shown to be useful in designing solvents or solvent mixtures to optimize yields or selectivities of reactions or sets of reactions (Eckert, 1967; Wong and Eckert, 1969). For instance, often a solvent mixture is better than either pure component for separation processes (Grieger and Eckert, 1967), and correspondingly, examples exist where solvent mixtures markedly accelerate reaction rates (see Figure 1). Yet another extremely useful application of a thermodynamic approach

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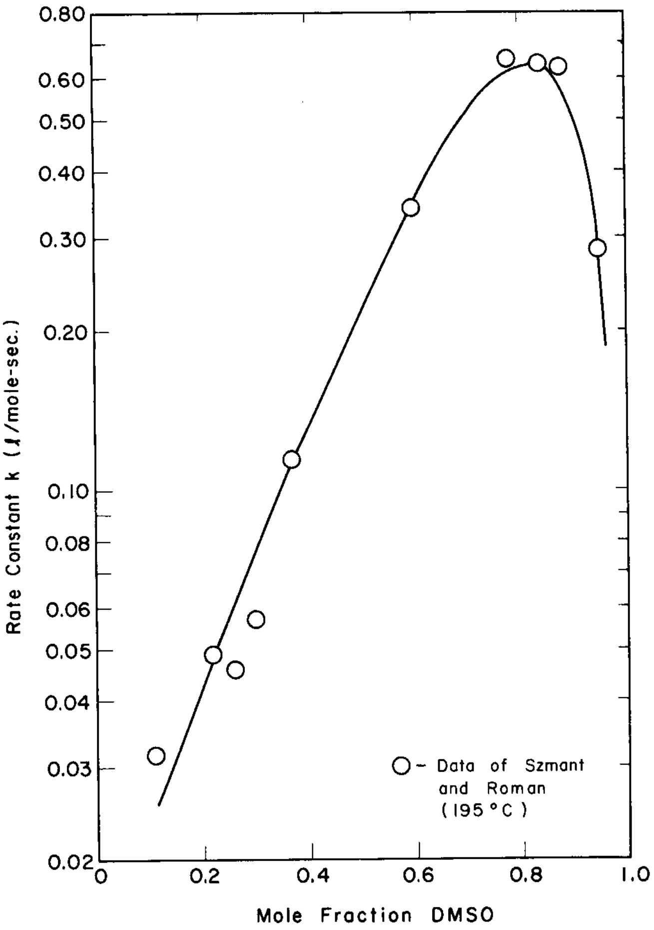

Fig. 1. The effect of solvent composition on the rate of a Wolff-Kishner reaction in mixtures of DMSO and butyl carbitol (Szmant and Roman, 1966).

to chemical reactions, largely overlooked by chemical engineers, is the use of linear free energy relationships to predict rates of reactions and chemical equilibria from data on similar reactions (Hammett, 1940). These provide the design engineer with a tool for predicting rates when only limited data, if any, are available. A number of other examples exist, and some will be discussed below, but they all have in common applications of thermodynamics, especially molecular thermodynamics, through transition state theory, to solve problems in chemical kinetics.

All chemical reactions occur in a gaseous, liquid, or solid solution; three closely interrelated steps are involved in applying a molecular solution theory to solve kinetic problems. In the first step, the rates of chemical reactions are related to the thermodynamic properties in solution of the reacting species and the reaction transition state. This has been done by the use of the classical transition state theory, which yields the enormous advantage of reducing a rate problem to an equilibrium problem. Both the theory and experimental techniques of solution thermodynamics are better understood and more accurate than those for kinetics. Moreover, many recent advances in molecular thermodynamics and applications of statistical mechanics provide new approaches to phase equilibria, which can be applied in conjunction with existing knowledge about intermolecular forces to characterize chemical reactions.

However, in order to predict the solution properties of a transition state, a molecule not isolable for study by its very definition, one must have some knowledge of its structure and properties. Thus the next step involves the characterization of the transition state, generally by the thermodynamic determination of its activation parameters and inferences drawn from these. From conventional kinetic experiments the energy and entropy of activation can be determined, and such techniques are well known. However, high pressure kinetics has proven to be another powerful experimental tool, as the activation volume, derived from the pressure dependence of reaction rates, yields precise molecular information about transition states not available otherwise—such as their structure (the mechanism of a reaction), their dipole moment, their specific solvent interactions, their degree of solvation, or even the details of their interaction with catalyst molecules.

The final part involves the use of some theory of solutions, some correlation, frequently borrowed from molecular thermodynamics, or preferably when possible, using as many actual experimental thermodynamic data as are available, to relate reaction rates (or equilibria) to known quantities. Here it must be emphasized that neither the methods discussed here, nor any other practical approach now available, can ever predict a priori rate constants. They can and do relate the rates to the rates of the same reaction or a similar reaction under other thermodynamic conditions.

TRANSITION STATE THEORY

The transition state theory, first proposed by Evans and Polanyi (1935) and applied by Glasstone, Laidler, and Eyring (1941), provides the bridge between thermodynamics and kinetics. This approach constitutes the thermodynamic framework within which the interpretation, correlation, and prediction of the rates of chemical reactions is feasible. It postulates for any elementary reaction the existence of an intermediate species M called the transition state

$$A + B \rightarrow M \rightarrow \text{Product}$$

where under normal conditions this species is in strict thermodynamic equilibrium with the original reactants.

$$A + B \rightleftharpoons M$$

Reaction occurs by the decomposition, at a fixed frequency (k_BT/h), of the transition state to product so that the rate is strictly proportional to the concentration, not the activity, of the transition state. This may be found from the chemical equilibrium relationship

$$(M) = K^{\ddagger} (A) \ (B) \cdot \frac{\gamma_A \gamma_B}{\gamma_M} \tag{1}$$

where $K^{\ddagger}$ is the equilibrium constant and the γ's are activity coefficients in the mixture. From this the rate constant k may be shown to be

$$k = \kappa \frac{k_BT}{h} \cdot K^{\ddagger} \tag{2}$$

where κ is the transmission coefficient, usually unity, for the probability that a decomposition of the transition state will result in product. This equilibrium constant in Equation (2) is related to a Gibbs energy in the usual manner,

$$\Delta G^{\ddagger} = - RT \ln K^{\ddagger} \tag{3}$$

and from $\Delta G^{\ddagger}$ the temperature and pressure derivatives lead to the energy, enthalpy, entropy, and volume of activation.

It is not necessary to believe in the existence of transition states to apply this very useful approach. For example, Marcus (1967) has shown that transition state theory is not dependent on an equilibrium assumption; rather, the same end result may be reached by assuming only the adiabatic postulate, that is, that as the molecules move along the reaction coordinate and as the internuclear distances change, the electronic energy levels may shift, but the quantum states remain constant.

Transition state theory was first applied to solution kinetics by Brønsted (1922) and Bjerrum (1924), who showed that the rate of reaction in a thermodynamically nonideal system is given by

$$\text{Rate} = k_0 \frac{\gamma_A \gamma_B}{\gamma_M} (A)\ (B) \tag{4}$$

where k_0 is the rate constant in the ideal reference system (where $\gamma_A = \gamma_B = \gamma_M = 1$). Then, if the activity coefficients are evaluated in terms of some theory of solutions, and the basic rate constant at atmospheric pressure k_0 is a function of temperature only, then the apparent rate constant k in any real solution is given by

$$k = k_0 \left(\frac{\gamma_A \gamma_B}{\gamma_M} \right) \tag{5}$$

They demonstrated this approach for reactions between ions in dilute aqueous solution using the Debye-Hückel treatment, and the results, which will be discussed below, are in quantitative agreement with experiment (Olson and Simonson, 1949).

As a result of this relationship, the problem of predicting reaction rates becomes a problem of predicting thermodynamics activities or Gibbs energies in solution of the reactants and of the transition state. In order to do this successfully, it is necessary next to determine the structure and properties of the transition state for various types of reactions, and finally, to apply workable techniques for predicting solution Gibbs energies from a knowledge of molecular properties.

STRUCTURE AND PROPERTIES OF THE TRANSITION STATE

The key to any understanding of chemical kinetics in solution is a knowledge of the structure and properties of the reaction transition state. Since the transition configuration exists only for a minute length of time in extremely dilute solution, inferential means are required to study its properties. For this purpose a thermodynamic approach is used, starting from Equations (2) and (3), which relate the rate constant to $\Delta G^\ddagger$, the Gibbs energy of activation.

Conventional experimental methods which measure the temperature derivative of $\Delta G^\ddagger$ measure an activation energy or entropy, which in turn allow estimation of the internal energy, entropy, or even the heat capacity of the transition state relative to the reactants. These quantities relate to some extent to a geometric effect, primarily of translational degrees of freedom, and more so to the internal degrees of freedom, rotation, vibration, and especially, the large electronic effects produced by shifts in energy levels as the transition state is formed. From these, an interpretation of the structure—that is, the reaction mechanism—may often be quite difficult. While such data as the activation energies are useful for many purposes, they are not always sufficient for some of the applications under consideration here.

On the other hand, the pressure derivative of the Gibbs energy of activation, equivalent to differentiation of the rate constant, gives an activation volume $\Delta v^\ddagger$

$$-\left(\frac{\partial \Delta G^\ddagger}{\partial P} \right)_T = RT \left(\frac{\partial \ln k}{\partial P} \right)_T = -\Delta v^\ddagger \tag{6}$$

where the rate constant must be expressed in pressure-independent units (Hamann, 1963) such as mole fraction or molality. The activation volume is rigorously the difference in partial molal volumes of transition state and reactants

$$\Delta v^\ddagger = \bar{v}_M - \bar{v}_A - \bar{v}_B \tag{7}$$

Thus from high pressure kinetic data coupled with partial molal volume determinations for the reaction substrates, Equations (6) and (7) yield the partial molal volume in solution of the transition state. This quantity is of relatively little use in gas-phase reactions, but in condensed phases it and its variations can give a great deal of very detailed information about the structure and properties of reaction transition states. For example, a study of the solvent variation of $\Delta v^\ddagger$ or $\bar{v}_M$ permits the separation of the volume into two portions. The structural contribution represents an intrinsic size or difference in molecular size between reactants and transition state, due in general to changes in bonding, and is useful in mechanistic interpretation and in elucidation of internal interactions. The other part is the solvent-dependent contribution showing changes in electrostriction of the solvation shell on activation. This second contribution gives indications of charge distributions within the transition state as well as specific solvent interactions.

Although activation volumes are quite important tools for determining reaction mechanisms, this approach had until recently fallen into some disfavor because of both the difficulty of the classical technique and the basic inaccuracies inherent in it. To achieve significant rate changes in a condensed phase, rather high pressures are required, generally in the range of 1000 to 10,000 atm. The equipment is specialized and cumbersome, and the mere generation of such high pressure gives quite significant volume changes and heat effects, even in condensed phases. Moreover, the actual data generally measured, concentration as a function of time, must be differentiated twice to evaluate an activation volume so that much precision may be lost in reducing data.

However, relatively recently new experimental techniques have been developed which provide substantially better data for significantly less experimental effort. For example, the in situ initiation technique (Grieger and Eckert, 1970a) eliminates many of the common difficulties, as do in situ analysis methods (Stewart and Weale, 1965; Hyne et al., 1966; Gay and Whalley, 1968, 1970; Baliga and Whalley, 1969; Mackinnon et al., 1970) or combinations of the two techniques (Eckert et al., 1974a). Such applications are discussed in detail in a number of recent reviews (Weale, 1967; leNoble, 1967; Eckert, 1972). Here we shall present such results for one type of reaction to demonstrate the application of high pressure kinetic data.

Quite extensive data are now available for Diels-Alder reactions, the condensation of a diene with a dienophile (generally a conjugated olefin) to form a cyclohexane ring. This reaction is important as a ring-forming reaction both industrially and in synthetic organic chemistry. However, much more important, it serves as an excellent example of the large class of molecular, nonpolar reactions. The term *nonpolar* is used in the organic chemist's sense of the word, meaning only that the polar effects, though present, do not dominate.

Until recently, considerable controversy existed over which of two possible mechanisms of the Diels-Alder reaction was correct (see Figure 2)— either a single-step multicenter addition (I) or a two-step process proceeding through a true diradical intermediate (II). It should be quite easy to distinguish between these two alternatives by measurements of the volume of activation, as it is well established that in the absence of strong electrostrictive effects, activation volumes are 10 to 15 cc negative per bond formed (Weale, 1967; leNoble, 1967). However, several questions were raised (Gonikberg, 1960; Benson and Berson, 1962) about earlier high pressure studies of such reactions (Raistrik et al., 1939; Walling and Peisach, 1958; Walling and Schugar, 1963) which appeared to indicate mechanism II, whereas the bulk of other evidence seemed to favor the concerted mechanism I (Lambert and Roberts, 1965; Charton, 1966).

Now newer and more precise high pressure kinetic data are available for at least ten different Diels-Alder reactions, five of which are reported in several solvents, primarily taken by the in situ initiation technique (Grieger and Eckert, 1970a, 1970b, 1970c, 1970d, 1971; Brun and Jenner, 1972; McCabe and Eckert, 1974). All give volumes of activation in the range of − 30 to − 45 cc/mole, explicable only in terms of the formation of two bonds in the transition state. Such results demonstrate conclusively that only mechanism I, the single-step four-center cycloaddition, can be valid for the reactions studied. In fact, the data showed that for addition of five dienes to maleic anhydride the partial molal volume of the transition state is smaller than that of the product in spite of the fact that two bonds in the transition state are only partially formed. Such results, coupled with other evidence, indicate strongly that as has been postulated by Hoffmann and Woodward (1965), there are strong secondary interactions (π-bonding) occurring between nonbonding centers in the transition state. These serve to make the transition state more compact, and corresponding to a maximum accumulation of double bonds, they help to determine isomeric products, corresponding to the well-known endo rule of cycloaddition (Alder and Stein, 1937). This result has a substantial impact not only on the stereochemistry of the product, but also in the rate of its formation, as these secondary interactions appear to lower the activation energy by several kcal.

In addition to internal or secondary interactions, high pressure kinetic studies have been useful in determining and understanding both solvent and substituent effects, in ascertaining the dipole moment of transition states, and additionally in elucidating mechanisms of catalysis. Such examples will be discussed below.

EFFECTS OF SOLVENTS

The choice of solvent is one of the most important factors in determining the rate and therefore selectivity of a liquid phase chemical reaction. Kinetic solvent effects on reaction rates can, as mentioned previously, be eight to ten orders of magnitude and variations in rate of the order of a factor of 100 to 1000 are rather common. Obviously, a method for predicting such effects would be a very useful tool in the area of exploratory reaction design since much trial and error could be eliminated. Fortunately, the application of Equations (4) or (5) in conjunction with an adequate theory of solutions does permit the estimation of rates with reasonable accuracy. However, it does have the disadvantage that some data are needed; only relative, not absolute, rates may be estimated.

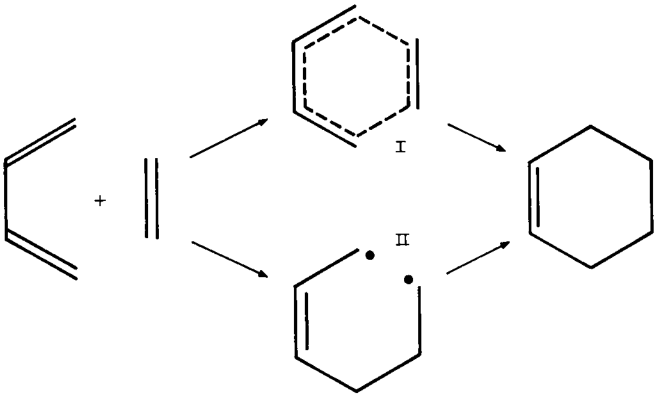

Fig. 2. Alternate mechanisms proposed for the Diels-Alder reaction: I-Single-step four-center ring closure; II-Two-step reaction with diradical intermediate.

In applying the Brønsted-Bjerrum equation in combination with some theory of solutions, the problem arises that most solution theories are strictly applicable only to one particular type of solution, such as dispersion force solutions, dipolar interactions, ionic solutions, etc., but the effects often overlap, such as a nonpolar reaction in a highly dipolar solvent. Nonetheless, in considering solvent effects on reaction rates, reactions will be classified in three broad categories; nonpolar reactions, polar reactions, and ionic reactions. We shall now consider each of these classes in conjunction with the molecular thermodynamic approach just discussed.

Nonpolar Solutions

While the influence of solvent on reaction rates has been extensively studied, the majority of such work has been directed toward ionic and highly polar reactions since they usually exhibit much more pronounced solvent effects. Therefore, solvent effects on nonpolar reactions (also called *nonionic* or *electroneutral*) have been studied relatively less.

The Diels-Alder cycloaddition reaction, in which solvent effects are generally less than a factor of 100, is a good example of a nonpolar reaction. Outside of the work in this laboratory only three comprehensive studies of solvent effects have been reported Sauer et al. (1964) studied the addition of fumaronitrile to 9,10-dimethylanthracene, and Dewar and Pyron (1970) and Silber (1971) investigated the isoprene-maleic anhydride reaction. Dewar and Pyron used ten solvents and tried to correlate their rate data with the dielectric constant of the solvent without much success. Silber measured the rate in 12 solvents and tried correlating the results with such basic solvent parameters as the dielectric constant, polarization, solubility parameter, relative internal pressure, and the empirical Y, Z, and E_t values. Silber found that some solvents formed complexes with maleic anhydride, and only after the rate constants were corrected for complex formation did he obtain a reasonable correlation with the E_t value. Correlation with the other solvent parameters were less successful.

Since the Diels-Alder reaction is a nonpolar reaction, a nonpolar theory of solutions should be useful for correlating solvent effects. Such an approach, which stresses van der Waals interactions, is the regular solution theory of Hildebrand and Scott (1962, 1964). Since regular solutions are characterized by the absence of any specific interactions between molecules, the theory should apply only to solutions of nonionic, nonpolar, or slightly polar molecules. The idea of applying regular solution theory to kinetics was first proposed by Glasstone (1941) and later implemented by Wong and Eckert (1969). According to regular solution theory, the activity coefficient of component i in a multicomponent mixture is given as

$$RT \ln \gamma_i = v_i(\delta_i - \bar{\delta})^2 \qquad (8)$$

where v_i is the liquid molar volume, δ_i is the solubility parameter (the square root of the cohesive energy density), and $\bar{\delta}$ is the volume-fraction average solubility parameter for the solution. For dilute solutions, the average solubility parameter can be approximated by that of the solvent. Substituting Equation (8) into (5) yields a relative rate expression, with the constants in mole fraction units,

$$\ln\left(\frac{k}{k_0}\right) = \frac{v_A(\delta_A - \bar{\delta})^2 + v_B(\delta_B - \bar{\delta})^2 - v_M(\delta_M - \bar{\delta})^2}{RT} \qquad (9)$$

Equation (9) also looks attractive as a method for predicting relative kinetic effects because the values of the molar volumes and solubility parameters for the reactants are readily available pure-component properties. The partial molal volume of the transition state may be evaluated from high pressure kinetic data, that is, activation volumes. δ_i is defined as the energy of vaporization of component i divided by its molar volume. For the transition state, this is equivalent to the liquid phase dissociation of M to A and B, vaporization of A and B and recombination to give M. To a good approximation the activation energies of dissociation and recombination are equal if there is negligible solvation so δ_M can be expressed in terms of known properties, and there are no adjustable parameters used in this approach.

$$\delta_M = \left[\frac{v_A\delta_A^2 + v_B\delta_B^2}{v_M}\right]^{1/2} \qquad (10)$$

The approach of Equations (9) and (10) should apply only where regular solution theory is applicable, strictly to solutions of nonionic nonpolar (or slightly polar) molecules. Polar forces or specific interactions (hydrogen bonds or charge-transfer complexes) should result in deviations. In fact, this analysis works rather better than one would expect. Wong and Eckert (1969, 1970a) used this approach to correlate, with no adjustable parameters, the rate data for the maleic anhydride- 1,3-butadiene reaction in seven solvents. In addition, the data of Grieger and Eckert (1970b, 1971); Dewar and Pyron (1970), Soula et al. (1966), Wasserman (1952), Coster and Pfiel (1968), and McCabe and Eckert (1974) were correlated. In all cases the data were found to be in good agreement with the theory even for some relatively polar solvents such as acetone. A comparison of the prediction with data on maleic anhydride with isoprene is shown in Figure 3. As expected, the most polar solvents, acetonitrile and nitromethane, show the largest deviations.

In addition, this approach has been used to calculate the relative yields of two competing reactions as well as for the design of optimum solvents or solvent mixtures for a given reaction (Eckert, 1967; Wong and Eckert, 1969).

Subsequent investigations (Wong and Eckert, 1970a) on the activity coefficients of the maleic anhydride + 1,3-butadiene reaction are helpful in explaining the good correlation obtained by regular solution theory even in relatively polar mixtures. It should be remembered that Equation (5) is based on the Brønsted-Bjerrum relation, which involves a ratio of activity coefficients. From direct measurements of reactant activities in solution, it was shown that the ratio of activity coefficients for maleic anhydride and the transition state $\gamma_{MA}/\gamma_{\ddagger}$ is nearly constant even though the individual γ's vary by orders of magnitude. Such a result is quite in agreement with Ingold's (1962) interpretation of the Hammond Principle (1955) that the electronic structure of the transition state should resemble the reactants more than the products for an exothermic reaction. Since a solute's electronic structure is the primary factor in determining its activity coefficient in a given solvent, it may be expected that the effects of specific solvent-solute interaction with the reactants and transition state, such as hydrogen bonding, will tend to be cancelled out in Equation (5). Whenever this type of compensation occurs, the regular solution theory can be expected to give a good approximation of solvent effects. However, deviations will still occur in solvents where strong specific interactions are present. Wong suspected that such an interaction was occurring with nitromethane. This has subsequently been confirmed by Silber (1971), who found that maleic anhydride forms a 1:1 complex with nitromethane.

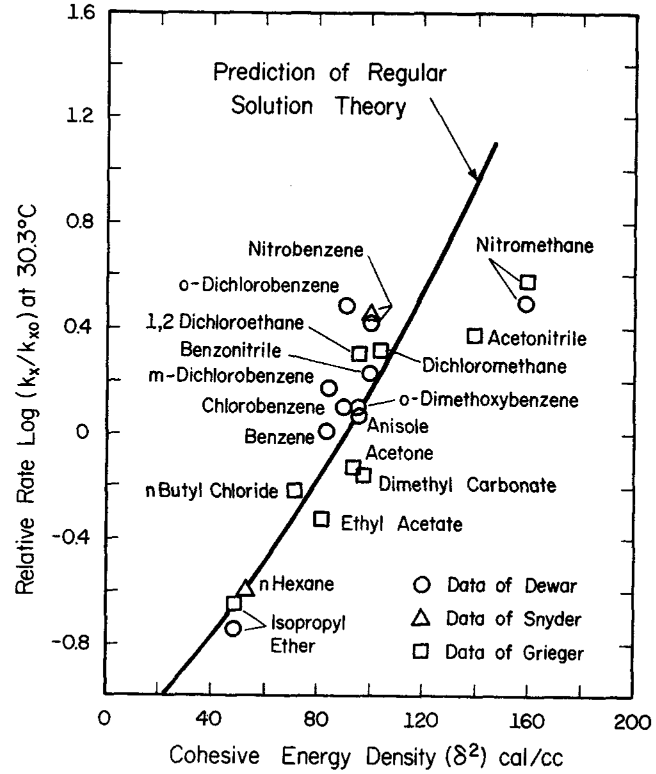

Fig. 3. Application of regular solution theory to kinetic solvent effects on the Diels-Alder addition of isoprene to maleic anhydride (Dewar and Pyron, 1970; Grieger and Eckert, 1970b; Snyder and Eckert, 1973).

Thus, the regular solution theory approach to solvent effect correlation for Diels-Alder cycloadditions does quite an adequate job, failing badly only with highly polar solvents or ones which are involved in specific interactions with the reactants. This emphasizes the fact the current solution theories are generally only strictly applicable to only one class of solution (regular solutions, dipolar solutions, etc.).

Polar Solutions

Although the regular solution theory does not apply to polar reactions or polar solutions, another approach has often been used, and this is derived from the Kirkwood (1934) expression for the work (or Gibbs energy at constant T and P) required to place an ideal dipole μ at the center of a spherical cavity of radius r within an isotropic medium of dielectric constant ϵ.

$$\Delta G = -\frac{\mu^2}{r^3}\left(\frac{\epsilon - 1}{2\epsilon + 1}\right) \qquad (11)$$

This expression neglects, of course, any nonelectrostatic interactions, such as van der Waals forces, hydrogen bonds, or charge-transfer complexes. When Equation (11) is used to calculate activity coefficients and combined with Equation (5), then the kinetic solvent effect for a bimolecular reaction is given as

$$\ln\left(\frac{k}{k_0}\right) = \frac{1}{k_B T}\left(\frac{\epsilon - 1}{2\epsilon + 1}\right) \cdot \left[\frac{\mu_M^2}{r_A^3} - \frac{\mu_A^2}{r_A^3} - \frac{\mu_B^2}{r_B^3}\right] \qquad (12)$$

where the logarithm of the rate constant is linear in the dielectric constant factor, with the slope proportional to the term in brackets, which might be termed a change in dipole moment density.

In general, the success of Equation (12) in predicting kinetic solvent effects for polar reaction has been rather sporadic. Many authors have pointed out (Benson, 1960; Baekelmans et al., 1964) that the nonelectrostatic contributions are rarely negligible. The reaction which has most often been compared with Equation (12) is the highly polar Menschutkin reaction of an alkyl halide with a tertiary amine, through a polar but nonionic transition state, to form the product, a quaternary ammonium salt.

$$R_3N + R'X \rightarrow R_3N \ldots . R' \ldots . X \rightarrow R_3R'N^+ + X^-$$

Numerous examples of solvent effects on such reactions are available (Kondo and Tokura, 1964, 1967; Hartmann et al., 1965, 1968; Heydtmann et al., 1966; Jungers et al., 1969; Eckert et al., 1974b), but all the results show that although Equation (12) may be qualitatively useful, quantitatively its success is rather variable. For example, Figures 4 and 5 show examples of the application of the Kirkwood method to rate data on Menschutkin reactions, all in aprotic solvents only, and agreement is good.

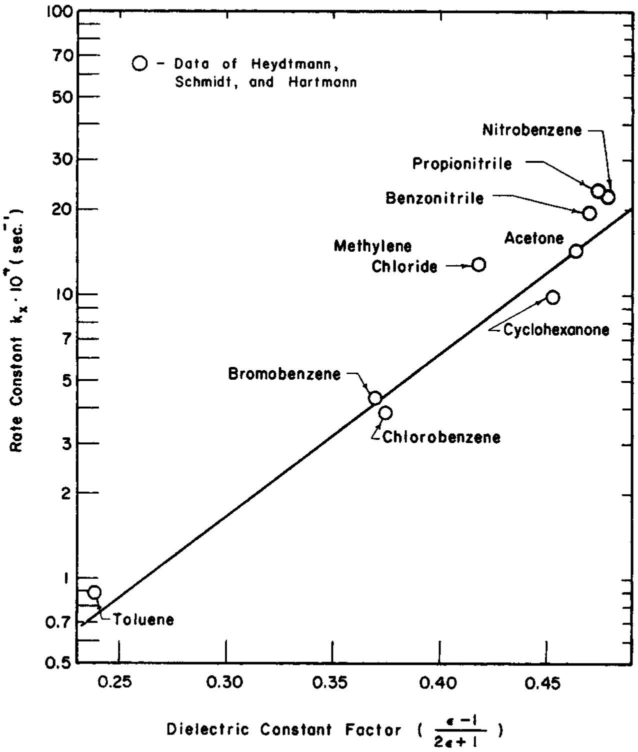

Fig. 4. Electrostatic approach to solvent effects on a Menschutkin reaction, α-picoline with ω-bromoacetophenone at 30°C (Heydtmann, et al., 1966).

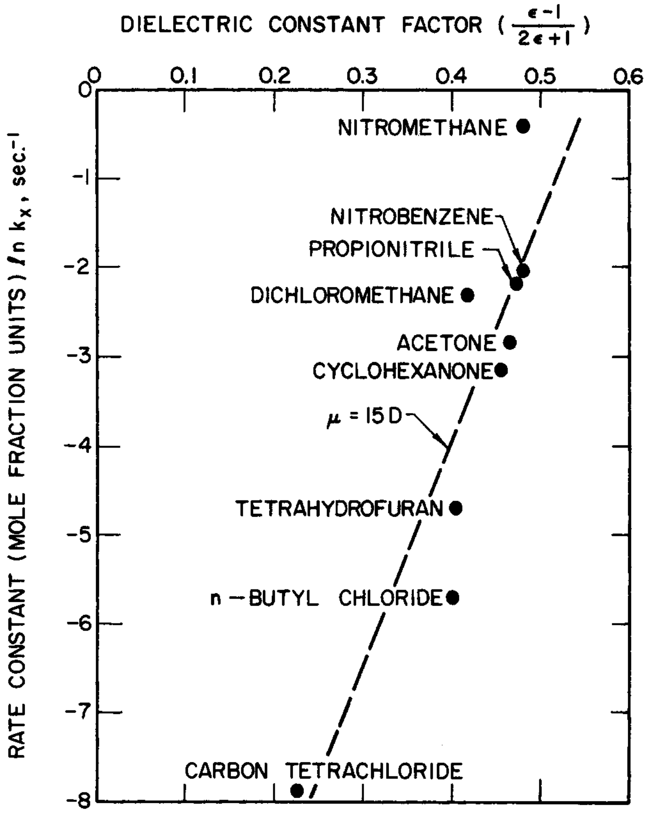

Fig. 5. Electrostatic approach to solvent effects on a Menschutkin reaction, tripropylamine-methyl iodide at 30°C (Eckert et al., 1974b).

Thus, this Kirkwood approach is known to be a terribly overly simplistic model, and it does in some cases (especially for associating solvents) give poor results, but it generally gives a reasonable, if not highly quantitative, picture of the polar interactions. Its predictions should be viewed cautiously, but no significantly better solution theory exists for polar reactions.

One additional area in which the Kirkwood model has proven especially useful is in the estimation of the dipole moment of reaction transition states from high pressure kinetic data. Since the rate constant is related to a Gibbs energy, its derivative with respect to pressure gives a volumetric term. Thus, the pressure derivative of Equation (12) gives the volume of activation in terms of the pressure derivative of the dielectric constant factor.

$$\Delta v^\ddagger = \Delta v_0^\ddagger - N^0\left[\frac{\mu_M^2}{r_M^3} - \frac{\mu_A^2}{r_A^3} - \frac{\mu_B^2}{r_B^3}\right] \frac{\partial}{\partial P}\left(\frac{\epsilon - 1}{2\epsilon + 1}\right) \qquad (13)$$

where $\Delta v_0^\ddagger$ is the activation volume in the absence of electrostriction. This expression of course includes all the basic assumptions in the simple Kirkwood model, some of which are admittedly not well satisfied. However, the advantage in applying Equation (13) is that any errors which are pressure-independent vanish. For example, although nonelectrostatic (that is, van der Waals) forces are generally not negligible as the Kirkwood model assumes, it appears that often the pressure derivative of this effect truly is negligible (Heydtmann et al., 1966; Hartmann et al., 1968). Thus if the activation volume is known in solvents with varying values of $\frac{\partial}{\partial P}\left(\frac{\epsilon - 1}{2\epsilon + 1}\right)$, then a linear plot should yield as slope the quantity inside the brackets, and taking r^3 as proportional to molar volume, a good estimate of μ_M is possible. As an example of this, the activation volume data for the tripropylamine-methyl iodide Menschutkin reaction are plotted according to Equation (13) in Figure 6, and from the slope a rough estimate of μ_M is possible, 13 Debye in this case, in reasonable agreement with the slope of the atmospheric pressure data for the same reaction (μ_M = 15 Debye from Figure 5).

Equation (13) represents a difference of effects and as such is rather insensitive. A more refined approach can be achieved by viewing the volumes of individual species rather than the difference of large numbers which a $\Delta v^\ddagger$ represents. Direct differentiation of Equation (11) for the *i*th species in a mixtures gives a linear expression for the partial molal volume $\overline{v}_i$ as

$$\overline{v}_i = \overline{v}_{0i} - N^0 \left[\frac{\mu_i^2}{r_i^3}\right] \frac{\partial}{\partial P}\left(\frac{\epsilon - 1}{2\epsilon + 1}\right) \qquad (14)$$

and since the partial molal volumes of reactants and products of a reaction may be determined individually (Heydtmann et al., 1966; McCabe et al., 1970; Grieger et al., 1971), in addition to the activation volume, Equation (14) can be combined with Equation (7) for the determination of individual dipole moments directly. One must recognize that there are many limiting assumptions in the Kirkwood approach, and the results are only semiquantitative. However, no better method currently exists. As an example of such an application the partial molal volumes for the transition states of five reactions of substituted 1,3-butadienes with maleic anhydride are shown in Figure 7 (McCabe and Eckert, 1974). From this plot, the estimated values of μ_M are given in Table 1, along with relative values of the rate constant and the activation energy.

Such results demonstrate that the transition state polarity can be characterized by both the position and properties of the substituent. 1-substitution leads to a larger μ_M than 2-substitution, most likely due to the effect of placing electron donating groups (CH_3, OCH_3) on the bond-forming carbon atom. The polar substituent effect, with respect to type, appears to be $OCH_3 > CH_3 > Cl$. With

TABLE 1. DIELS-ALDER REACTION OF SUBSTITUTED BUTADIENES WITH MALEIC ANHYDRIDE (MCCABE AND ECKERT, 1974)

Diene	μ_M (Debye)	$k \cdot 10^{4*}$ sec.$^{-1}$ at 35°C	E_a* kcal/mole
1-methoxybutadiene	7	264	10.0
1-methylbutadiene (piperylene)	4	52.2	11.0
2-methoxybutadiene	4	8.72	13.9
2-methylbutadiene (isoprene)	3	30.0	13.1
2-chlorobutadiene (chloroprene)	3	1.00	15.3

* Average values for several solvents.

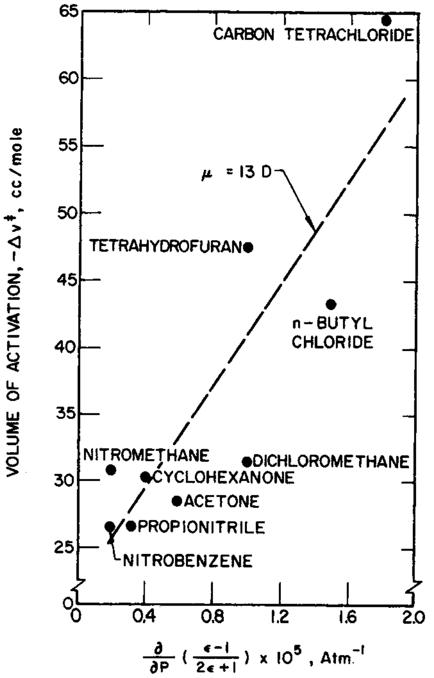

Fig. 6. Activation volumes in various solvents for the tripropylamine-methyl iodide reaction at 30°C (Eckert et al., 1974b).

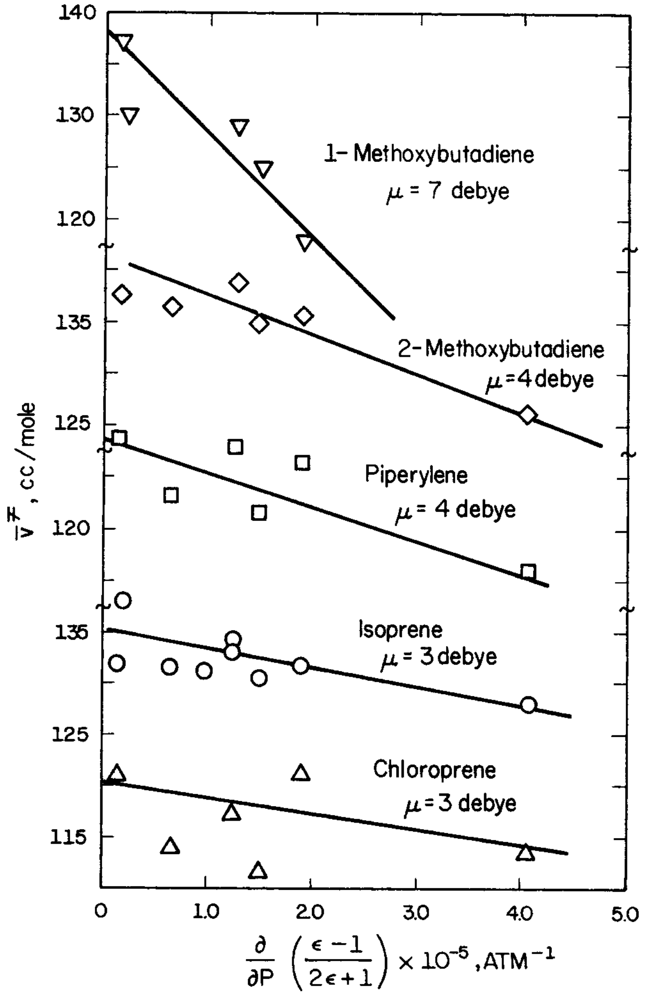

Fig. 7. Partial molal volumes in various solvents for the transition states of five Diels-Alder additions of substituted butadienes to maleic anhydride (McCabe and Eckert, 1974).

an electron deficient dienophile such as maleic anhydride, there is an electron flow from the electron-rich diene to the dienophile which takes place during the reaction. The strong electron donating character of the methoxy substituent in the 1-position thus has a large effect on the partial transfer of charge which occurs and accounts for the large transition state polarity and low activation energy observed with 1-methoxybutadiene. In comparison, chloroprene, with the electron withdrawing chlorine substituent, exhibits the highest activation energy, the lowest rate, and least polar transition state.

Table 1 also indicates that the rate is not proportional to transition state polarity but does correlate with the activation energy. This indicates that kinetic substituent effects should correlate better with the electronic rather than the polar characteristics of the substituent as one would expect for a nonpolar reaction. Thus a linear free energy method is quite applicable for treating kinetic substituent effects (see below).

Ionic Solutions

In addition to molecular reactions, there are many reactions of both industrial and theoretical importance which involve ionic reactants or reaction intermediates. Yet in spite of this importance, there is relatively little quantitative understanding of how to correlate and predict rate constants of reactions involving ions at reasonable concentrations due to the limitations of existing solution theories. Only in the very dilute region, where the rigorous electrostatics of the Debye-Hückel theory apply, are exact correlations available.

The original application of the Brønsted-Bjerrum relation, Equation (5), was in the variation of rate with ionic strength from the Debye-Hückel (1923) theory. Though the effect is rather small and limited to extremely dilute solutions, well under 0.01 M, the results are in excellent quantitative agreement with experiment (Livingston, 1930; Olson and Simonson, 1949). Also, the same theory was used to determine the solvent effect on ionic reactions, in the limit of infinite dilution only (Scatchard, 1932a). Many experimental studies have verified that the rate constant at infinite dilution, expressed as ln k_0', is a linear function of the reciprocal of the dielectric constant (LaMer and Kamner, 1935; Eagle and Warner, 1936; Davis and LaMer, 1942). An example of such a plot of solvent effects is presented in Figure 8.

Unfortunately, however, there is little practical interest in kinetics in the very dilute Debye-Hückel region. Moreover, this theory applies only to ion-ion reactions and predicts nothing about ion-molecule reactions. In most reactions involving ions in solution, the overwhelmingly dominant factor is the relative solvation of ionic reactants, transition state, and products, and virtually all other effects are second-order. For example, by including solvation by the method of Scatchard (1923b, 1936), Clements and Eckert (1971) were able to extend predictions of the salt effect on the rate of an ion-ion reaction in several solvents up to ionic strengths of about unity, as shown for the bromoacetate-thiosulfate reaction in water in Figure 9.

To show the importance of solvation effects as well as the difficulty in predicting or even understanding them, let us take ion-molecule reactions as an example. The rate will of course depend on the Gibbs energy difference between the reactants and the transition state; the problem of the solvent effect on reaction rates is therefore again the determination of the Gibbs energies, entropies and enthalpies of solvation of the reactants and of the transition state (Leffler and Grunwald, 1963; Wiberg, 1964; Arnett et al., 1965).

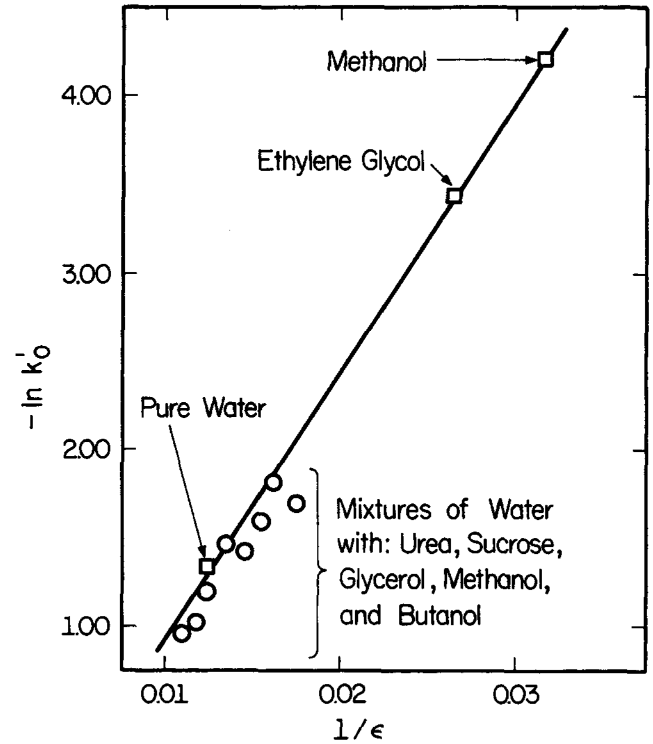

Fig. 8. Effect of solvent dielectric constant on the rate of an ionic reaction, bromoacetate-thiosulfate at 25°C (Clements and Eckert, 1971).

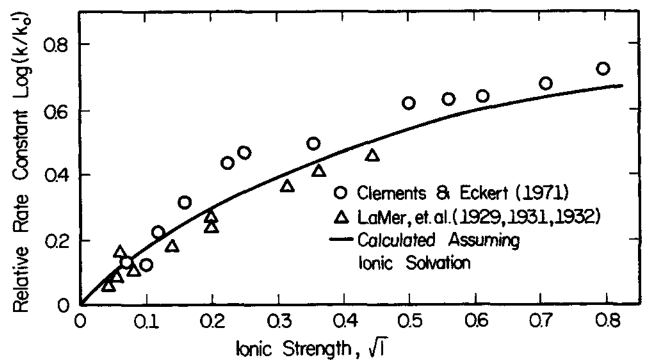

Fig. 9. Effect of ionic strength on the rate of an ionic reaction, bromoacetate-thiosulfate in water at 25°C.

For a bimolecular nucleophilic substitution reaction

$$Y^- + RX \rightarrow [Y^{\delta-} - - - R - - - X^{(1-\delta)-}]^{\ddagger} \rightarrow RY + X^-$$

the Brønsted-Bjerrum expression becomes

$$\frac{k_A}{k_0} = \frac{{}^0\gamma_Y{}^s - {}^0\gamma_{RX}{}^s}{{}^0\gamma_{\ddagger}{}^s} \qquad (15)$$

where the solvent activity coefficient ${}^0\gamma_i{}^s$ reflects the change in the standard chemical potential of species i on transfer from an arbitrary reference solvent 0 to another solvent s.

Most reactions of this type are considerably faster in dipolar aprotic solvents than in protic solvents. Protic solvents, such as water, methanol, formamide, and ammonia are hydrogen bond donors, while aprotic solvents such as dimethylformamide (DMF), acetone, hexamethylenephosphoramide, nitromethane, acetonitrile, and many others are highly polar but no more than weak hydrogen bond donors. The much faster rates of the bimolecular nucleophilic substitution reactions in dipolar aprotic solvents relative to protic solvents can be attributed to the enhanced solvation of the nucleophiles in the protic solvent or to greater solvation of the transition states in the dipolar aprotic solvents, since the solvation of the neutral molecule is considered to be relatively small. Parker (1969) determined in various solvents the activity coefficients of anions, neutral molecules and transition states relative to a reference solvent, methanol. In the determination of the solvent activity coefficient of anions, he used extrathermodynamic assumptions to divide the solvent activity coefficient of electrolytes into those of the individual ions. From the resulting values he concluded that the reactions are faster in dipolar aprotic solvents because the reactant anion Y^- is much more solvated by protic than by dipolar aprotic solvents, and this outweighs any effects due to transition state anion or neutral molecule solvation. The main source of the solvation of anions by protic solvents is hydrogen bonding since normal ions are strong hydrogen bond acceptors.

Another approach for interpreting the solvent effects on the rate of bimolecular substitution reactions was attempted by Haberfield et al. (1969). They determined the enthalpy of transfer of a transition state from one solvent to another by measuring the heats of solution of reagents in methanol and DMF and assuming that tetrabutylammonium ion and tetrabutylboride ion have the same enthalpies of transfer. The enthalpy of transfer of transition states from methanol to DMF was exothermic, and in most cases the enthalpies of transfer of the reactants were also exothermic. This means that both the transition state and the reactants are generally more solvated in DMF than methanol. However, the effect is much greater for the transition state than for the nucleophiles, which results in a lower enthalpy of activation in DMF than in methanol.

This solvent effect on the rate of a bimolecular reaction has been interpreted by two different approaches; one approach emphasizes the solvation of nucleophiles, and the other approach the solvation of the transition state as the primary factor in determining the reaction rate. Such solute-solvent interactions can also be studied in terms of the partial molal volume, which may be obtained from dilatometric measurements coupled with high pressure kinetic data.

Such measurements are available for both symmetric (Ershov et al., 1959, 1960) and unsymmetric (Byon, 1973) halide exchange reactions. For example, the results for the reversible reaction of *n*-propyl bromide with iodide ion are given in Table 2. The rates and equilibria vary by orders of magnitude between the aprotic acetone and acetonitrile and the highly protic alcohols and *N*-methyl formamide (NMF). Such results are due to the greater solvation of the bromide ion relative to that of the iodide ion in the protic solvents, while the solvation of the iodide ion is larger than that of the bromide ion in the aprotic solvents. This agrees with the conclusion of Parker (1969) that, in general, small ions are more solvated than large ions by hydrogen bonding in protic solvents, while large ions are more polarizable and so more solvated than small ions by aprotic solvents, which are more polarizable than protic solvents.

Table 2. Experimental Data for the Reaction of Iodide Ions with *n*-Propyl Bromide (Byon, 1973)

Solvent	Rate constant · 10^3 sec.$^{-1}$ at 50°C	Equilibrium constant	$\Delta v^\ddagger$ cc/mole	$\Delta \bar{v}_{rxn}$ cc/mole
Acetone	313	0.0286^a	$+4^a$	$+2.0^a$
Acetonitrile	64.9	0.0802^a	-6.2^a	-3.4^a
Ethanol	6.58	2.48^b	-18^b	-3.1^b
Methanol	4.80	2.26^b	-14^b	-2.5^b
NMF	10.0	1.057^b	-17^b	-3.4^b

[a] Data at 25°C.
[b] Data at 50°C.

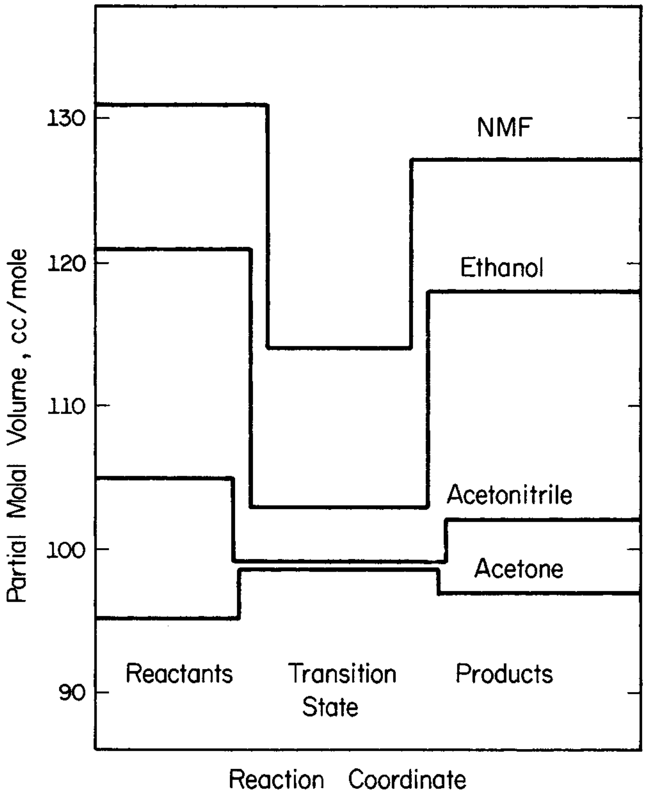

Fig. 10. Volume profile along the reaction coordinate (from high pressure kinetics and dilatometric data) for the halide exchange reaction propyl bromide-iodide ion (Byon, 1973).

This effect can be shown dramatically in a plot (Byon, 1973) of the partial molal volume for this reaction along the reaction coordinate (see Figure 10). Since the partial molal volumes of all the propyl halides are very close to their molar volumes, solvation of these species is negligible. All effects on rates and equilibria which are related to solvation or solvent-solute interactions are due to relative changes in the solvation of the halide anions and of the ionic reaction transition state.

Unfortunately, the large variations in solvation, reflected not only in the volume changes shown in Figure 10 but also in the rate data and calorimetric data, are not at this time a priori predictable. Some factors are understood qualitatively (Parker, 1962, 1969; Dack, 1970), and clearly some recent advances in modern solvation theory appear promising (Friedman, 1973), but a significantly better quantitative understanding of solvation phenomena will be required before it will be possible to make accurate and reliable predictions of solvent effects on reactions involving ionic species.

LINEAR FREE ENERGY RELATIONS

Another quite different approach exists for using a thermodynamic technique to make predictions about both the rates and equilibria of chemical reactions. For more than three decades, organic chemists have used the so-called "linear free energy relationships," which relate rates of reactions of like mechanisms to correlatable structural features of the reactants. Chemical engineers have not generally recognized the utility of such methods for making practical predictions in reaction design, especially for the effects of substituents and reagents on rates of related reactions.

The earliest example of a correlation of the linear free energy type was the Brønsted catalysis law relating catalytic activity to the strength of an acid or base in solution (Brønsted and Pedersen, 1924). In general, at constant temperature the logarithm of the kinetic rate constant k is proportional to $\Delta G^\ddagger$. For chemical equilibria the corresponding quantities are the equilibrium constant K and the standard Gibbs energy of reaction ΔG^0. The proportionality constant depends on the reaction series as well as on the substrate. Many reviews have appeared in the chemical literature (Hammett, 1940; Jaffe, 1953; Pal'm, 1961; Leffler and Grunwald, 1963; Wells, 1963; Ehrenson, 1964; Ritchie and Sager, 1964; Farcasiu, 1966; Wells, 1968; Shorter, 1969; Wong and Eckert, 1970b), and the subject has continued to receive considerable attention not only in organic chemistry but also in other branches of chemistry.

As above, this technique also reduces rate problems to equilibrium problems by working through the transition state theory and the accompanying Gibbs energy of activation (or the standard Gibbs energy of reaction for the case of equilibrium), which is of course the sum of an enthalpy contribution and an entropy contribution,

$$\Delta G^\ddagger = \Delta H^\ddagger - T\Delta S^\ddagger \tag{16}$$

The usual form of the linear free energy relationship is to take some given reaction as a model and assume that in a similar reaction (with some relatively minor change in either the solvent, the substrate, or the substituent group) one of the following three conditions is satisfied:

1. $\Delta H^\ddagger$ is unchanged.
2. $\Delta S^\ddagger$ is unchanged.
3. Although both $\Delta S^\ddagger$ and $\Delta H^\ddagger$ change, the variation in them is linearly related (Thorn, 1969).

Most of the correlations proposed have been two-parameter relationships between the above quantities. For example, one of the most frequent applications of linear free energy relationships is for the substituent effect. For these cases, the correlations are used to estimate the rate of reaction using as data the rate of another reaction of similar mechanism, involving a related molecule with similar structure, but differing by a substituent group not at the reaction site. For example, suppose one wanted to estimate the rate of esterification of *p*-chlorobenzoic acid by methanol. If the rate constant for the reaction of *m*-nitrobenzoic acid with methanol were known, this could then be used to make a very good quantitative estimate of the rate of the former reaction. To apply this type of correlation, organic reactions are categorized by mechanism, and treatments are available for various types of mechanisms (Gould, 1959; Hine, 1962; March, 1968). For instance, the classic Hammett correlation for the effect of meta- and para-substituents on aromatic side chain reactions is expressed as

$$\log\left(\frac{k}{k_0}\right) = \rho\sigma \tag{17}$$

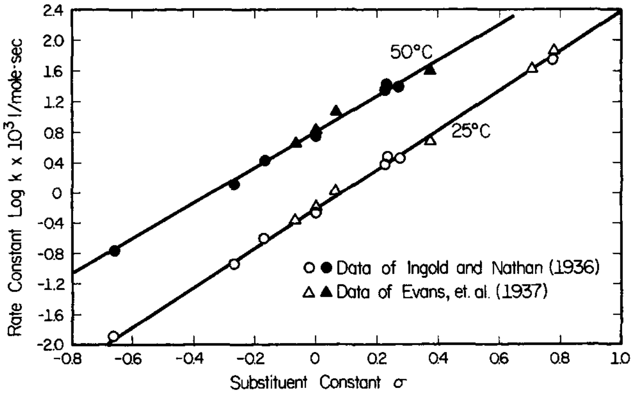

Fig. 11. The Hammett equation applied to the hydrolysis of substituted ethyl benzoates.

where k_0 represents the rate constant for some model reaction, σ represents the relative energetic effect of a given substituent, and ρ is a constant for each type of reaction, representing the sensitivity of that class of reactions to substituent effects. An example of the application of Equation (17) is the hydrolysis of substituted ethyl benzoates, shown in Figure 11 (Ingold and Nathan, 1936; Evans et al., 1937).

Linear free energy relationships have also been developed to correlate rates and equilibria of chemical reactions where the changes occur at the reaction site itself, due to the effect of reagents. The most common examples of this type are the rates of acid-catalyzed or base-catalyzed reactions such as dehydrations or hydrolyses. Here the rates will depend on the strength of the catalyst and thus on the ionization equilibrium in solution. This, of course, is a thermodynamic quantity and lends itself readily to correlations of the type discussed here. For example, Brønsted and Pedersen (1924) found that a simple equation related the rate constant for the acid-catalyzed decomposition of nitramide to the dissociation constant of the acid, K_A

$$k = aK_A{}^\alpha \tag{18}$$

where a and α are constants that depend upon the nature of the substrate, the class of acids, the solvent, and the temperature. The exponent α is known as the Brønsted slope or Brønsted coefficient. By selecting an acid as a standard, Equation (18) can be expressed as a Hammett-type equation

$$\log\left(\frac{k}{k_0}\right) = \alpha \log\left(\frac{K_A}{K_{A0}}\right) \tag{19}$$

In Figure 12, the Brønsted correlation for the acid-catalyzed dehydration of acetaldehyde hydrate (Bell and Higginson, 1949) is shown as an example for seventeen different acid catalysts.

Also, a number of linear free energy relationships have been proposed for kinetic solvent effects (see, for example, Dack, 1970). Probably the best-known example of such methods is that proposed by Grunwald and Winstein (1948) for solvolysis reactions. They proposed a parameter Y as a measure of the solvating power of various solvents based on the solvolysis of *t*-butyl chloride relative to that in 80% aqueous ethanol as the model reaction. Then the relative solvent effect for a reaction is given by

$$\log\left(\frac{k}{k_0}\right) = mY \tag{20}$$

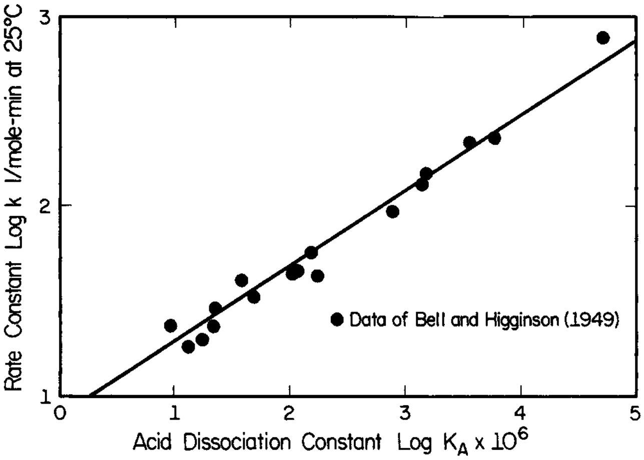

Fig. 12. The Brønsted correlation for the acid-catalyzed dehydration of acetaldehyde hydrate.

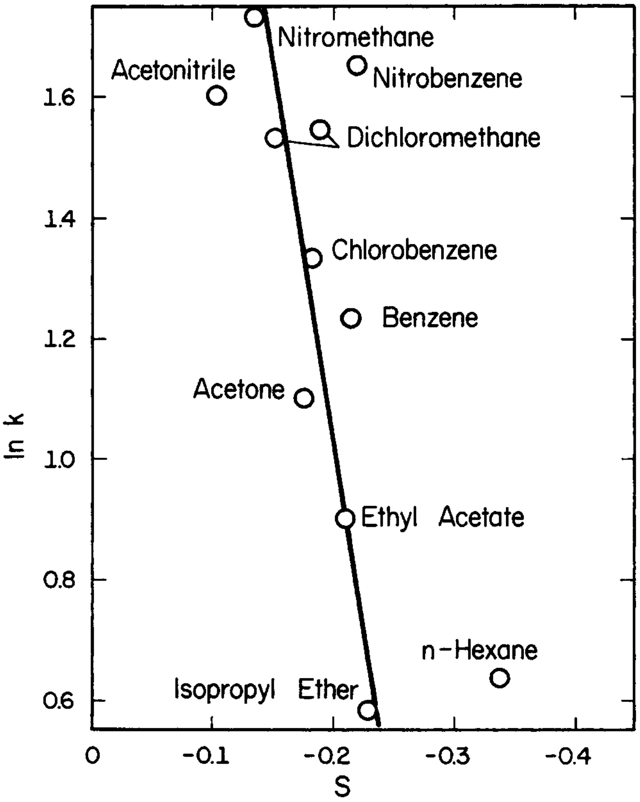

Fig. 13. The Brownstein correlation for the solvent effects on the isoprene-maleic anhydride reaction.

Wells (1963, 1968) has compiled values of m and Y for pure solvents and mixtures, and in general, m is near unity for S_N1 reactions, but appreciably smaller for S_N2 displacements. In any event, the mY method has given reasonable correlations for a number of solvolysis reactions (Fainberg and Winstein, 1956) for which the dielectric constant techniques do not give good results. However, for a given substrate no unique value of m is found which may be used for all solvents; rather the correlation of log k with Y generally yields a series of almost parallel lines, one for each binary solvent pair. Incomplete solubility of the substrate is a frequent limitation in using solvolysis as a model series in Equation (20) for application to other reaction series since many additional solvent parameters would have to be obtained indirectly. Other frequent complications are the effect of the leaving group, the nucleophilicity of the solvent, and the problem of internal return (Wells, 1963, 1968; Amis, 1966).

A similar form proposed by Brownstein (1960) relates the rate constant in any solvent k_s to that in ethanol k_e, in the form

$$\log\left(\frac{k_s}{k_e}\right) = SR \tag{21}$$

where S is a solvent parameter and R a reaction parameter. As might be expected, this gives good results for solvolysis data but rather poorer ones for most other rates and equilibrium data. As an example, this correlation has been applied to the kinetic data for the Diels-Alder reaction of isoprene-maleic anhydride in those solvents for which values of S are available (see Figure 13). The agreement is only fair, certainly not as good as that obtained from regular solution theory.

A somewhat different approach, most appealing in its simplicity, is the proposal of Drougard and Decroocq (1969) to relate solvent effects of any reaction to solvent effects on a standard reaction, the Menschutkin addition of methyl iodide to tri-n-propylamine in a series of organic solvents. This can be expressed in the form

$$\log\left(\frac{k}{k_0}\right) = \lambda\tau \tag{22}$$

where τ is a solvent parameter and λ a reaction parameter. Although the original authors did not express their results in this form, Wong and Eckert (1970b) have evaluated λ and τ by assuming λ equal to unity for the model reaction in benzene solution. This permits wide application of Equation (16) since rate data (Lassau and Jungers, 1969) are available for a large variety of solvents—78 pure solvents and many binary mixtures. Also, benzene is a common solvent used in many kinetic studies. For many reactions, especially the highly polar reactions similar to Menschutkin reactions, the correlation gives surprisingly good results (Drougard and Decroocq, 1969). An example of another reaction for which it is successful is shown in Figure 14 (Pincock, 1964). However, in some cases, such as for Diels-Alder reactions, it often does not work as well as other methods.

A good example of the engineering application of linear free energy relationships is the design equation for Diels-Alder additions to maleic anhydride developed by McCabe and Eckert (1974). This involves application of the Hammett Equation (17) along with regular solution theory to predict both solvent and substituent effects on the rate for this class of reactions. Although intended to correlate reactions and equilibria involving aromatic side chains, the Hammett equation has been found to be useful for Diels-Alder reactions (Brown and Okamoto, 1957; Charton, 1966; Takeno and Morita, 1970; Inukai and Kojima, 1971), probably due to the highly conjugated quasi-aromatic nature of the transition state.

Typical results for Diels-Alder reactions are shown in Figure 15 where values of σ indicate the electron withdrawing power of substituent groups on the diene. The rates are linear for all solvents investigated, and the Hammett plot lines have very nearly the same slope ρ, all about -3.5. This unusually large value of ρ is consistent with the strong rate enhancement from electron-donating substituents.

With very little error, well within the accuracy of most literature data, these lines may be considered parallel and the intercepts considered to be a solvent parameter d, thus yielding a general equation for both solvent and

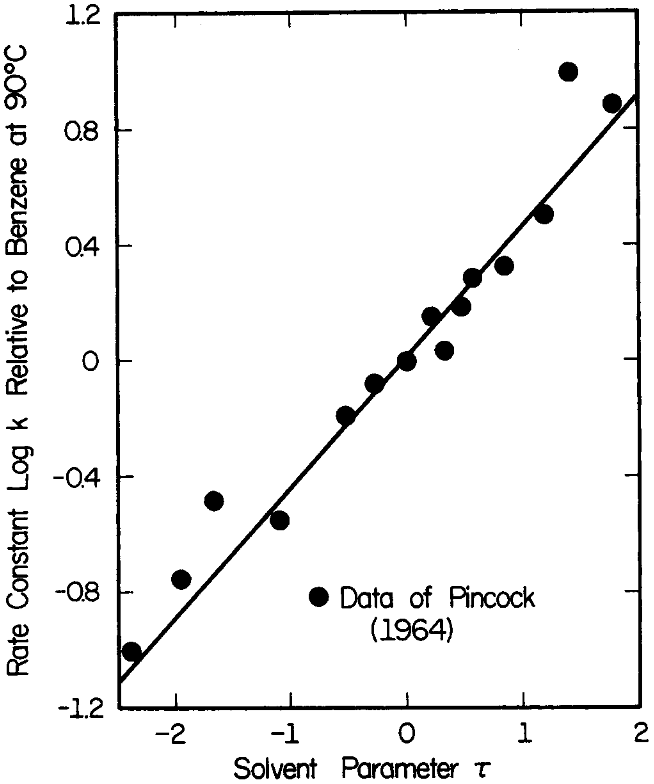

Fig. 14. The Drougard-Decroocq correlation for solvent effects on the ionic decomposition of t-butyl-peroxyformate catalyzed by pyridine.

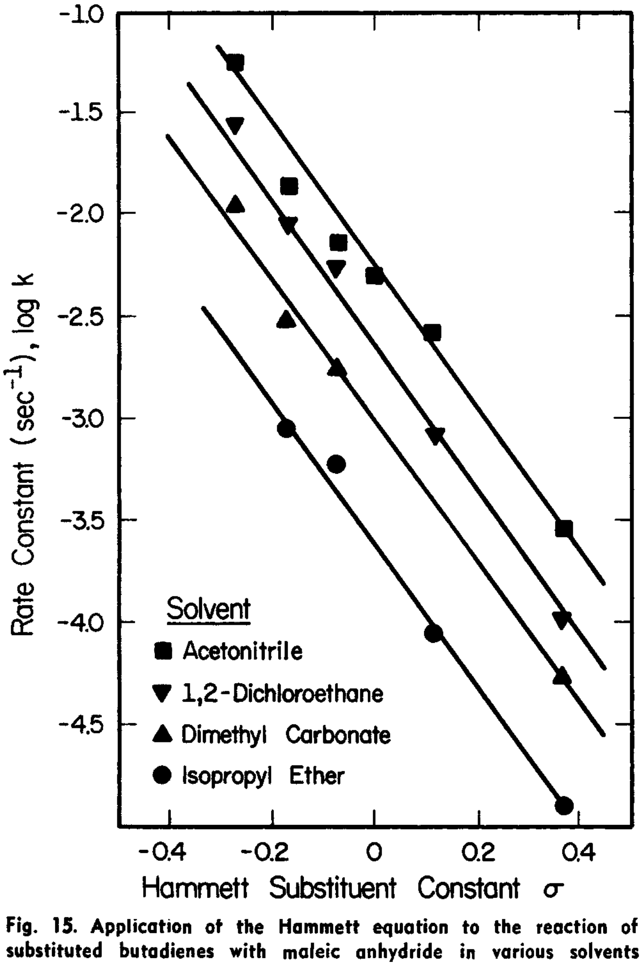

Fig. 15. Application of the Hammett equation to the reaction of substituted butadienes with maleic anhydride in various solvents (McCabe and Eckert, 1974).

substituent effects

$$\log\left(\frac{k}{k_0}\right) = \rho(\sigma - \sigma_0) + (d - d_0) \tag{23}$$

where the subscript 0 refers to the reference reaction in the reference solvent. As might be expected, the intercepts d correlate quite closely with the solubility parameter, and for hundreds of data over a very wide range of rates, solvents, and substituents, Equation (23) has been shown to predict rates to an accuracy of about ± 25%. Considering that the precision of the data is often no better than this, Equation (23) represents a remarkably good expression, with no adjustable parameters, for predicting a wide variety of rates. As such, it might prove useful for preliminary design work in optimizing rates, evaluating solvents or solvent mixtures, or predicting the relative rates of competing reactions. In addition, the same type of approach should be capable of extension to other reactions.

OTHER APPLICATIONS

In addition to the topics discussed above, there exist a few other areas in which applications of molecular thermodynamics or solution theory to problems in chemical kinetics have yielded useful results. Here three such topics will be reviewed: the effect of nonidealities on gas-phase reactions, the effect of a second-order phase transition (a critical point) on chemical reactions, and studies of catalyzed reactions.

Gas-Phase Reactions

For a chemical reaction in an ideal gas phase, the partial pressure is generally (and correctly) used as a measure of concentration. However, this does not mean that in a nonideal gas the partial pressure in the rate expression may simply be replaced by fugacities. Despite the fact that some chemical engineering kinetics books advocate such a procedure, it is clearly incorrect. Rigorous application of transition state theory (Eckert and Boudart, 1963) gives the gas-phase Brønsted-Bjerrum equation for a bimolecular reaction

$$\frac{k}{k_0} = \frac{\phi_A \phi_B z}{\phi_M} \tag{24}$$

where the ϕ's are fugacity coefficients and z is the usual compressibility factor of the gaseous mixture. Thus with Equation (24) the question of the variation of rate with thermodynamic conditions is quite similar to that in liquid solutions. One must have an adequate model of the transition state and a valid equation of state of the mixture.

This was first applied by Eckert and Boudart (1963) to the bimolecular pyrolysis of hydrogen iodide. The model of Wheeler et al. (1936) for the transition state was used with the virial equation to give results in excellent agreement with the experimental data of Kistiakowsky (1928). Subsequently, Mills and Eckert (1969) used a modified Redlich-Kwong equation of state (Redlich et al., 1965) to calculate gas phase solvent effects, and substantial if not dramatic effects were calculated, especially when the solvent gas was near its critical point. Similar solvent effects in dense gases on equilibrium controlled reactions were investigated by Sortland and Prausnitz (1965), using the same equations of state for the gas phase. They too found that for strongly interacting gas-phase solvent species at high density significant changes in equilibrium could be anticipated.

Such effects are not unexpected, as dense gases can often have excellent solvent properties, interacting strongly with dilute species. There has been one very detailed experimental investigation of this phenomenon which demonstrates this result most dramatically.

Simmons and Mason (1972a) investigated the gas phase 1,2-cycloaddition dimerization of chlorotrifluoroethylene at pressures up to 100 atm. They also made careful measurements of the *P-V-T* properties of the monomer and used the results to develop a four-term virial equation which gave results in excellent agreement with the experiments (see Figure 16). As might be expected, neither the Redlich-Kwong expression nor a truncated virial equation was adequate at the higher pressures. Although all experiments were at temperatures above the critical temperature of the reactant monomer, at higher pressures the density approached the critical density.

These results demonstrate that gas-phase nonidealities may be treated by the Brønsted-Bjerrum expression also, although in general the effects are small unless the gas is quite dense. Since the most nonideal species in most reactions is the very dilute transition state, it is quite important to use an equation of state that yields accurate values of the fugacity of dilute species. The virial equation is the only expression that includes theoretically valid mixing rules, so it is not surprising that it seems to yield the best results.

Reactions Near a Critical Point

The marked changes in curvature in the 120°C data of Simmons and Mason (1972a) shown in Figure 16 occur nearest to the critical condition of the reaction mixture. Further, they observed very abrupt changes in the partial molal volume of the transition state at the conditions nearest the critical, as it even became sharply negative. Since the transition state is a dimer of sorts, it is the less volatile species in a critical mixture and would be expected to have a small or negative partial molal volume, and Simmons and Mason (1972b) showed that this followed reasonably from a consideration of a van der Waals gas. Using the same model they did not predict any point of inflection in the rate constant plot but did predict that the rate plot should have an infinite slope at the critical point itself.

This raises the very interesting point of what does happen to a reaction near a second-order phase transition. It is well established that many transport properties and derivative thermodynamic properties have extreme maxima, even diverging at criticality. Although classical thermodynamics (that is, van der Waals equation) does not correctly predict scaling laws, the actual type of behavior is reasonably well established (Fisher, 1967; Heller, 1967; Rowlinson, 1969). Virtually, no data exist in the literature for precise reaction rates very near a vapor-liquid critical point, no doubt due to the enormous experimental difficulty of maintaining an exact (and changing) critical temperature and pressure in a reacting system of changing composition.

Recently, however, there has been some work done at the mathematically equivalent, but experimentally more accessible, liquid-liquid critical point. Snyder and Eckert (1973) measured the rate of the Diels-Alder reaction of maleic anhydride with isoprene in the vicinity of the upper critical solution points of two liquid-liquid mixtures and also the rate of the Menschutkin reaction of ethyl iodide-triethylamine near a lower critical solution point. The reactions were slow and could be held at the exact critical point (within ± 0.01°C) by two methods—precalibration and laser-light scattering.

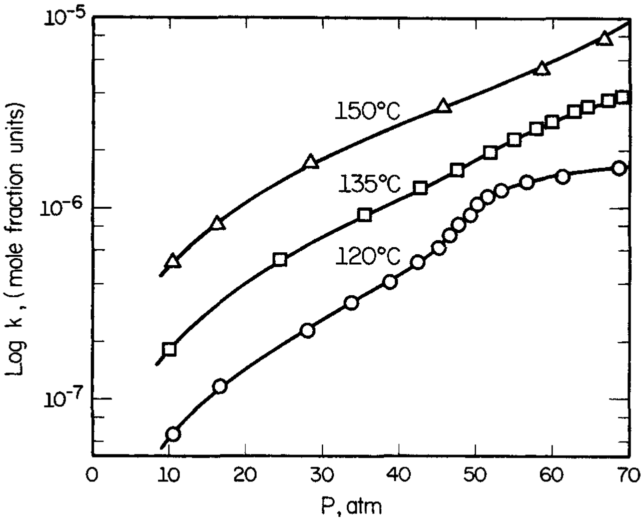

Fig. 16. Variation of the rate constant with pressure for the cyclic dimerization of chlorotrifluoroethylene (Simmons and Mason, 1972a).

The Diels-Alder cycloaddition was found to proceed as much as 30 to 40% faster at the upper critical solution points of the hexane-nitrobenzene and hexane-nitroethane solvent systems, while the Menschutkin reaction appeared to have a somewhat slower rate than expected at the lower critical solution point of the triethylamine-water system. The results are explained in terms of sharp changes in the activity coefficient ratio in Equation (5) for dilute solutes near a critical point and have been explained qualitatively by a decorated lattice-gas model of the critical state (Wheeler, 1972).

In a similar study Breuer et al. (1973), investigated the rate of a hydrolysis reaction near the liquid-liquid critical point of the water-2,6-lutidine system. They observed no anomalous behavior in this system, but from an irreversible thermodynamics analysis they conclude that unusual effects could occur near critical points, but for a given system no a priori calculation of them may be made. While these investigations may lead to further understanding of the nature of the critical state they probably have little practical application at the present time since the effects are at most rather small (20 to 40%) and the precise temperature control required would be extremely difficult to achieve on an industrial scale. One might speculate that the effects would be markedly larger at a vapor-liquid critical point, but the problem of control would be even more severe.

Catalyzed Reactions

Thermodynamic approaches have been used relatively little on studies of solvent effects on catalyzed reactions, probably primarily because variations in catalyst activity and/or transport phenomena are generally greater than the solvent effects. Eckert (1967) has given a very rough correlation for the competitive hydrogenation of cyclohexene and acetone on nickel (Jungers et al., 1958) in various solvents, emphasizing equilibria between adsorbed and bulk species. Recently, Boydell (1973) has completed an extensive study of both the kinetics and solution nonideality for a series of catalyzed esterification and transesterification reactions. From both vapor-liquid equilibria and calorimetric data he is able to give quite a good thermodynamic description of the reacting system in terms of the excellent Wilson (1964) equation and thus correlate the kinetic results.

A very different application of thermodynamics to a catalyzed reaction has been reported by Poling and Eckert (1972), who used high pressure kinetics together with solution data to elucidate the mechanism of catalysis of a homogeneously catalyzed reaction. They investigated the $AlCl_3$ catalyzed Diels-Alder reaction, and although there have been numerous examples of the catalysis of this cycloaddition reaction by Lewis acids (see, for example, Andreev and Andreeva, 1970 or Seltzer, 1968), there has been considerable controversy about the exact mechanism of the reaction, as both one-step and two-step schemes have been proposed, with the evidence being inconclusive on either side. The recent work clarified the mechanism by comparing the activation volumes and partial molal volumes for both the uncatalyzed and $AlCl_3$ catalyzed reaction of 2,3-dimethyl butadiene with *n*-butyl acrylate. The results, shown in Figure 17, demonstrate conclusively that both reactions proceed by the same mechanism, the concerted four-center cycloaddition, as the volume profiles are essentially identical for the catalyzed and uncatalyzed reactions, except that the partial molal volumes for the catalyzed reaction are shifted upwards by approximately 73 cc/mole due to the presence of the $AlCl_3$. The catalysis involves complexing of the $AlCl_3$ with the carbonyl oxygen, withdrawing electrons from the conjugated system and rendering the dienophile more reactive. Moreover, the $AlCl_3$ is complexed with the carbonyl oxygen in both the transition state and product as well but equilibrates rapidly with reactant so that at low conversion the catalyst is completely active. The role of the Lewis acid catalyst in this reaction is not to change the mechanism but merely to make the dienophile more reactive by rendering it more electron-deficient. Thus, high pressure kinetics and thermodynamic analysis provide a technique for finding out the exact molecular details of how this catalyzed reaction proceeds.

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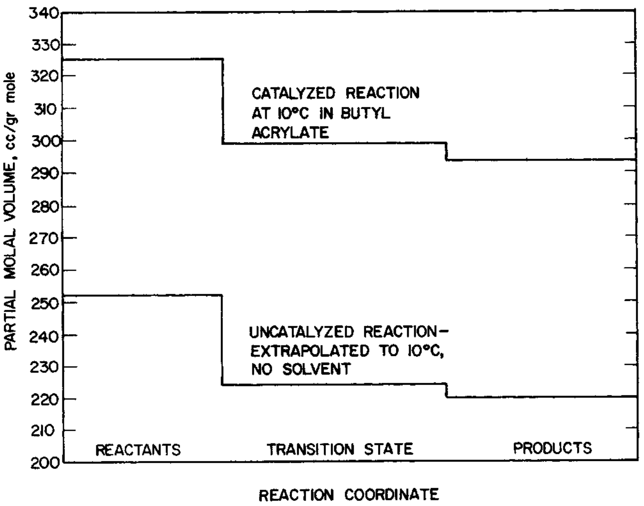

Fig. 17. The catalyzed and uncatalyzed cycloaddition of 2,3-dimethylbutadiene to butyl acrylate; volume profile along the reaction coordinate from high pressure kinetics and dilatometric data (Poling and Eckert, 1972).

NOTATION

A = reactant species
a = proportionality constant in Brønsted catalysis equation
B = reactant species
d = empirical solvent parameter
E_t = empirical solvent parameter
G = Gibbs energy
H = enthalpy
h = Planck's constant
I = ionic strength
K = equilibrium constant
k = rate constant
k_0 = rate constant, reference system
k_0' = rate constant at zero ionic strength
k_B = Boltzmann's constant
M = transition state
m = parameter in Grunwald-Winstein correlation
N^0 = Avagadro's number
P = pressure
R = gas constant, also constant in Brownstein correlation
r = radius
S = entropy, also constant in Brownstein correlation
T = temperature
v = molar volume
$\bar{v}$ = partial molal volume
$\Delta v^\ddagger$ = activation volume
Δv_{rxn} = volume change on reaction
Y = parameter in Grunwald-Winstein correlation
Z = empirical solvent parameter
z = compressibility factor

Greek Letters

α = Brønsted slope
γ = activity coefficient
δ = solubility parameter
ϵ = dielectric constant
κ = transmission coefficient
λ = reaction parameter in Drougard-Decroocq equation
μ = dipole moment
ρ = reaction parameter in Hammett equation
σ = substituent parameter in Hammett equation
τ = solvent parameter in Drougard-Decroocq equation
ϕ = fugacity coefficient

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Analytical Formulas for Disk Filters

Formulas relating the rate of filtrate flow in a sectioned disk filter to the process variables are derived. It is demonstrated that an optimum inner radius which yields a maximum flow of filtrate can be chosen. Flow rates for drum and disk filters operating under similar conditions are compared.

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SCOPE

Disk filters consist of a series of thin disks revolving on a common shaft and partially submerged in a slurry. They are useful where rapid cake buildup is possible under vacuum and washing is not required. Up to the present time, no mathematical analysis of the mechanism of a disk filter has been published. Empirical design has completely dominated the field.

A disk filter can be analyzed in a manner analogous to that used for a drum filter. However, the disk filter collects solids at variable distance from the axis of rotation, whereas all of the solids are deposited at a constant radius on the drum. An additional integration in the radial direction is necessary to obtain equations for the disc filter.

In the past, it has been customary (Nickolaus and Dahlstrom, 1956; Rushton et al. 1969; Tiller, 1972) to treat continuous rotary drums as though there were no internal sectioning and the vacuum were effective over the entire angle of submergence. Following the methodology of Risbud (1973), derivations in this paper account for sectioning and thus are more realistic than previous theoretical analyses of similar problems.

The objective of this study is to present formulas which can be used for the analysis and design of disk filters. The derivations are restricted to the case of constant specific filtration resistance and constant flow rate through the cake. While a more sophisticated analysis could have been made involving variable flow resistance and variable flow rate (Tiller and Shirato, 1964), the proposed equations represent a reasonable first step toward substituting theory for empiricism.

CONCLUSIONS AND SIGNIFICANCE

Design of disk filters has been entirely empirical in the past as no analytic equations have been available. A series of formulas which can be used for predicting behavior of disk filters under the limitation of constant specific resistance have been derived. (These formulas do not imply incompressible cakes.)

Equation (19) provides a double integral for numerical calculation of the flow rate as a function of number of sections, rotational speed, medium resistance, cake resistance, hydrostatic head, submergence angles, inner and outer radii of the disks, and slurry properties. Simplified formulas are presented for negligible medium resistance and hydrostatic head and also infinite sectioning.

Comparison of formulas for disk and drum filters indicate that the former may be expected to give two to three times the flow rate of the latter for equal floor space and the same radii.

Previously, it has not been recognized that there is an optimum value of the inner radius which yields a maximum flow rate. The optimum ratio of inner to outer radii varies from about 0.6 to 0.75 and is independent of the properties of the slurry.