

has been presented for the existence of a trihydrated hydroxyl ion, whatever its lifetime and its exact configuration might be. Concentrated solutions of the hydroxides can thus be pictured as solutions of a hydrated hydroxide ion in free water. Although more refined theories regarding the structure of free water and aqueous solutions in general<sup>39-41</sup> may require some

modification in this model, it seems to account well enough for the properties discussed. An attempt to interpret rate data in these solutions by this model will therefore be presented in the following paper.<sup>8</sup>

(39) A. Eucken, *Z. Elektrochem.*, **52**, 255 (1948).

(40) W. Kangro, *Z. physik. Chem. (Frankfurt)*, **32**, 273 (1962).

(41) G. Némethy and H. A. Scheraga, *J. Chem. Phys.*, **36**, 3382 (1962).

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## The Effect of Ionic Hydration on Rate and Equilibrium in Concentrated Alkaline Solutions. II. The Kinetics of Base-Catalyzed Reactions in Concentrated Hydroxide Solutions<sup>1a</sup>

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Equations are derived which relate the observed rate constants of various base-catalyzed reactions with the indicator basicity of the solution as expressed by the  $H_-$  function. Assuming that the hydroxide ion is a trihydrated species, the rate equations were expressed also in terms of  $C_{OH^-}$ , the stoichiometric concentration of hydroxide ion, and of  $C_{H_2O}$ , the concentration of "free" water. It is concluded that the rate dependence in concentrated alkaline solution primarily indicates the change in the number of water molecules involved in the formation of the transition state. In the cases of nucleophilic substitution by a nonhydrated hydroxide, as well as in a proton abstraction reaction, a rate law obeying  $C_{OH^-}/C_{H_2O}$  is derived. These equations have been verified experimentally in the cases of hydrolysis of ethyl iodide, a typical  $SN_2$  reaction, and in the breakdown of serine phosphate in basic solution which has been shown to proceed by a mechanism involving proton abstraction.

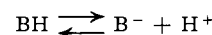
The kinetic behavior of acid-catalyzed reactions in concentrated acid solution has been the subject of increasing interest for some time. Considerable data have been accumulated,<sup>2</sup> the interpretation of which has been the subject of discussion ever since Zucker and Hammett advanced their original hypothesis.<sup>3</sup> This hypothesis states that reactions whose rate follows the appropriate indicator acidity function operate through a unimolecular conversion of the active species, while reactions whose rate tends to follow the stoichiometric acidity of the solution operate by a bimolecular reaction of the active species with solvent water.

Little attention has been paid in the past to reactions in concentrated basic solutions. Only recently have rate measurements been conducted in concentrated aqueous KOH and NaOH<sup>4a,b</sup> and in alkaline methanol.<sup>5</sup> In each of these reactions a correlation of the rate with the  $H_-$  acidity function was found. It was therefore of interest to evaluate what information on the mechanism can be obtained from this relationship in more general cases.

A formal derivation based on an approach analogous to that employed in the interpretation of acid-catalyzed reactions can be presented for each type of base-catalyzed mechanism correlating the observed rate constant with the  $H_-$  acidity function. Taking into consideration that the hydroxide ion is strongly hydrated, most probably as a trihydrated ion, it was shown<sup>6</sup> in analogy to the approach used by Bascombe and Bell<sup>7</sup> that the  $H_-$  function may be expressed in terms of concentrations of hydroxide ions and of "free" water. Consequently, the rate equations of base-catalyzed reactions in con-

centrated alkaline solutions may also be expressed in terms of hydrated hydroxide ion concentration and of "free water." The use of this form of the equations can lead to a better understanding of the mechanism of base-catalyzed reactions. Also it does point out the limitations of the use of the equations, as will be considered in the discussion.

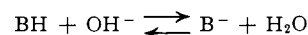
**Mechanisms of Base-Catalyzed Reactions. Definitions.**—The notation used here is that employed by Long and Paul<sup>2a</sup> in their discussion of acid-catalyzed reactions;  $H_-$  is defined<sup>2b</sup> by  $H_- = -\log (C_{BH}/C_B^-)K_{BH}$  where  $K_{BH}$  is the acid dissociation constant of an indicator dissociating according to



$h_-$  is defined<sup>2b</sup> by the equation  $H_- = -\log h_-$ ; hence

$$h_- = \frac{C_{BH}}{C_B^-} K_{BH} = \frac{f_{B^-}}{f_{BH}} a_{H^+} = \frac{f_{B^-} a_{H_2O}}{f_{BHA_{OH^-}}} K_w$$

In alkaline solutions the dissociation of BH can also be described by



Consequently one can define  $b_-$  as

$$b_- = K_w/h_- = f_{BHA_{OH^-}}/f_{B^-} a_{H_2O} \quad (1)$$

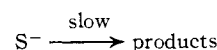
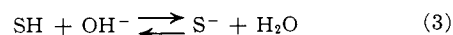
This quantity has the advantage that it is directly related to the  $OH^-$  activity and in dilute alkaline solutions, where the activity coefficients approach unity, it approaches  $C_{OH^-}$ . The function  $b_-$  is thus the alkaline counterpart of  $h_0$  in acid solution. A consequence of eq. 1 is the relation

$$f_{BH}/f_{B^-} = b_-(a_{H_2O}/a_{OH^-}) \quad (2)$$

which will be used in the discussion which follows.

Each of the main types of reaction which can occur in alkaline solution will now be dealt with in more detail.

**Case I.**—The substrate SH is in rapid pre-equilibrium with its conjugated base  $S^-$ ;  $S^-$  is then converted to the products of the reaction in a unimolecular rate-determining step



(1) (a) Presented before the 140th National Meeting of the American Chemistry Society, Chicago, Ill., September, 1961; (b) The Research Laboratories of the Israel Atomic Commission, Rehovoth, Israel.

(2) (a) F. A. Long and M. A. Paul, *Chem. Rev.*, **57**, 935 (1957); (b) M. A. Paul and F. A. Long, *ibid.*, **57**, 1 (1957); (c) H. Zollinger, *Ann. Rev. Phys. Chem.*, **13**, 391 (1962).

(3) L. Zucker and L. P. Hammett, *J. Am. Chem. Soc.*, **61**, 2791 (1939).

(4) (a) M. Anbar and G. Yagil, *ibid.*, **84**, 1790 (1962); (b) G. Yagil and M. Anbar, *ibid.*, **84**, 1797 (1962).

(5) M. F. L. Allison, C. Bumford, and J. H. Ridd, *Chem. Ind. (London)*, 718 (1958).

(6) G. Yagil and M. Anbar, *J. Am. Chem. Soc.*, **85**, 2376 (1963), to be referred to as Part I.

(7) (a) K. N. Bascombe and R. P. Bell, *Discussions Faraday Soc.*, **24**, 158 (1957); (b) R. P. Bell, "The Proton in Chemistry," Cornell Univ. Press, Ithaca, N. Y., 1959, p. 81 ff.

The concentration of the active species  $S^-$  is given by

$$C_{S^-} = C_{SH} \frac{K_{SH}}{K_w} \frac{f_{SH}}{f_{S^-}} \frac{a_{OH^-}}{a_{H_2O}} \quad (4)$$

In the following treatment only cases where  $C_{S^-} \ll C_{SH}$  will be considered, so that  $C_{SH}$  will also represent the *total* concentration of the substrate. In mechanism I, the rate-determining step is the unimolecular decomposition of  $S^-$ , and hence

$$-\frac{dC_{SH}}{dt} = k' \frac{f_{S^-}}{f^*} C_{S^-} = k' \frac{K_{SH}}{K_w} \frac{f_{SH}}{f^*} \frac{a_{OH^-}}{a_{H_2O}} C_{SH} \quad (5)$$

where  $f^*$  is the activity coefficient of the transition state. If the assumptions usually made in acid solution<sup>2a,3</sup> also hold in this case, then

$$f_{SH}/f^* = f_{BH}/f_{B^-} \quad (6)$$

and inserting eq. 5 into 4 and using relation 2 one gets

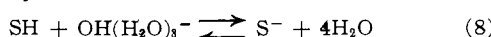
$$-\frac{dC_{SH}}{dt} = k' \frac{K_{SH}}{K_w} b_- C_{SH}$$

so that

$$k_{obs} = -\frac{1}{C_{SH}} \frac{dC_{SH}}{dt} = \text{const. } b_- \quad (7)$$

A first-order dependence of the rate on  $b_-$  is thus to be expected when a compound breaks down *via* the unimolecular reaction of its conjugate base, in complete analogy with the Zucker-Hammett hypothesis for the acid region.

As outlined in part I<sup>6</sup> the fast pre-equilibrium (3) may be rewritten taking into consideration the hydration of the hydroxide ion



If, as was suggested,<sup>6</sup> the activity coefficient ratio is close to unity and can be neglected, then the concentration of the active species  $S^-$  is

$$C_{S^-} = K_o \frac{C_{OH^-}}{C_{H_2O}^4} C_{SH}$$

$C_{H_2O}$  denoting the concentration of free water and  $K_o$  the equilibrium constant of eq. 8. It was also shown in part I that the "basicity function"  $b_-$  is given by a similar expression (*cf.* eq. 8 part I and eq. 1 above)

$$b_- = C_{OH^-}/C_{H_2O}^4 \quad (9)$$

The rate of the reaction can therefore be expected to follow  $b_-$

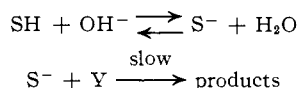
$$-\frac{dC_{SH}}{dt} = k'' C_{S^-} = k'' K_o C_{SH} (C_{OH^-}/C_{H_2O}^4)$$

$$k_{obs} = -\frac{1}{C_{SH}} \frac{dC_{SH}}{dt} = \text{const. } \frac{C_{OH^-}}{C_{H_2O}^4} = \text{const. } b_- \quad (10)$$

This means that the reaction, besides being first order in hydroxide concentration, exhibits inverse fourth-order dependence on free water concentration, as a result of the loss of four water molecules by the reactants (*i.e.*, as the  $OH(H_2O)_3^-$  ion plus a proton from the substrate), in forming the transition state. It has been assumed here that the activity coefficient ratio ( $f_{S^-}/f^*$ ) is unity ( $f'$  denotes activity coefficient of *hydrated* species in *free* water solution).

An example of a reaction following mechanism I is the alkaline decomposition of chloroform<sup>8</sup> which has been shown<sup>5</sup> to follow  $H_-$  in concentrated solutions of methoxide in methanol.

**Case II.**—The substrate SH is in rapid pre-equilibrium with its conjugate base  $S^-$ , as in mechanism I, but  $S^-$  is subsequently converted to the products in a *bimolecular* rate-determining interaction with another reactant Y



the rate expression in this case is

$$-\frac{dC_{SH}}{dt} = k' \frac{f_Y f_{S^-}}{f^*} C_Y C_{S^-} \quad (11)$$

Introducing the value of  $C_{S^-}$  from eq. 4 one obtains

$$-\frac{dC_{SH}}{dt} = k' \frac{K_{SH}}{K_w} \frac{f_Y f_{SH}}{f^*} \frac{a_{OH^-}}{a_{H_2O}} C_Y C_{SH} \quad (12)$$

An assumption may now be made that

$$f_Y f_{SH}/f^* = f_{BH}/f_{B^-} \quad (13)$$

This assumption may be justified if the Y component in the transition state interacts with the environment in a similar manner to its interaction in the free state;  $f_{S^-} = f^*/f_Y$  is then the coefficient of the "B-like" part of the transition state; eq. 13 is therefore the bimolecular counterpart of the usual assumption (6). Using this assumption and eq. 2 one may derive from eq. 12

$$k_{obs} = -\frac{1}{C_Y C_{SH}} \frac{dC_{SH}}{dt} = \text{const. } b_- \quad (14)$$

Applying the actual pre-equilibrium, eq. 8, leads without any further assumptions to the same result

$$-\frac{dC_{SH}}{dt} = k'' C_{S^-} C_Y = k'' K_o (C_{OH^-}/C_{H_2O}^4) C_Y C_{SH} \quad (15)$$

which gives

$$k_{obs} = \text{const. } (C_{OH^-}/C_{H_2O}^4) = \text{const. } b_- \quad (16)$$

which is identical with eq. 14. Thus both cases I and II show a similar dependence on  $b_-$  (or on *free* water concentration).

The reaction between chloramide ( $NHCl^-$ ) and ammonia in aqueous solutions to form hydrazine is an example of case II and this reaction was indeed found to be first order in  $b_-$ <sup>4b</sup> (slope 0.90).

**Case IIa.**—A case of special importance occurs when Y is a molecule of the solvent, *i.e.*, a water molecule, and the reaction is pseudo-first order. This mechanism, as well as mechanism I, has previously been discussed by Ridd.<sup>5</sup> Two methods of deriving the rate expression for this case can lead to two different conclusions; inserting  $H_2O$  instead of Y in eq. 12 yields

$$-\frac{dC_{SH}}{dt} = k' \frac{K_{SH}}{K_w} \frac{f_{H_2O} f_{SH}}{f^*} \frac{a_{OH^-}}{a_{H_2O}} C_{H_2O} C_{SH} \quad (17)$$

Equation 17 may be interpreted as indicating that the rate is proportional to  $C_{OH^-}$  as predicted by the Zucker-Hammett criteria for the A-2 cases. However, mechanism IIa can also be considered a special case of the general bimolecular mechanism II. If assumption 13 holds

$$f_{H_2O} f_{SH}/f^* = f_{BH}/f_{B^-} \quad (18)$$

then the observed rate constant will be

$$k_{obs} = -\frac{1}{C_{SH}} \frac{dC_{SH}}{dt} = \text{const. } b_- C_{H_2O} \quad (19)$$

The symbol  $C_{H_2O}$ , which has been used above to notate *free* water, has been retained here, since it is reasonable to expect the free water molecules to be the kinetically active ones. The same result is obtained using the hydration approach and applying eq. 15

$$k_{obs} = \text{const. } \frac{C_{OH^-}}{C_{H_2O}^3} = \text{const. } b_- C_{H_2O} \quad (20)$$

where again only the free water is considered to be the kinetically active species. It follows that the rate of a reaction which proceeds by bimolecular interaction of  $S^-$  with a free water molecule should be inversely proportional to the third power of free water concentration and directly proportional to the first power of  $C_{OH^-}$ . Unfortunately, no experimental data for a reaction known to proceed by mechanism IIa are available at present.

**Case III.**—A *proton abstraction* by a hydroxide ion from the nonionized substrate SH in the rate-deter-

(8) J. Hine, in "Physical Organic Chemistry," 2nd Ed., McGraw-Hill Book Co., Inc., New York, N. Y., 1962, p. 484.

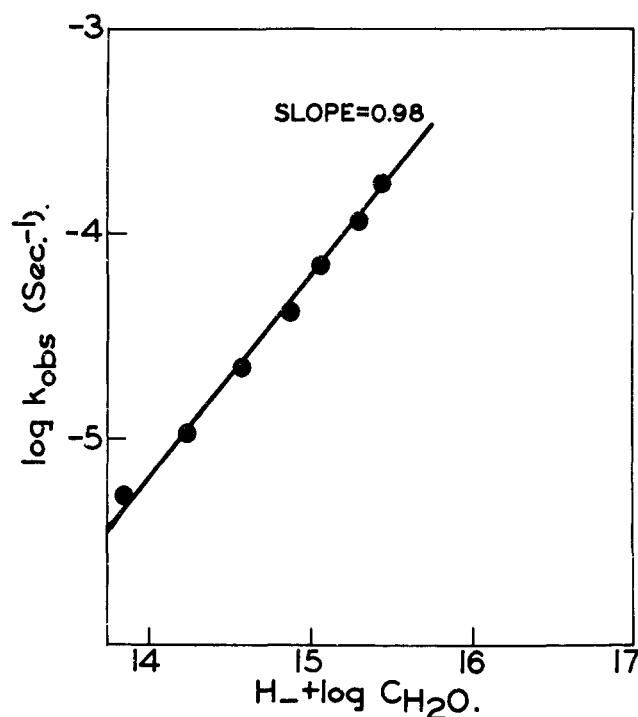
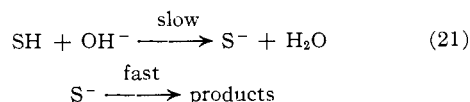


Fig. 1.—The rate of elimination of *d,l*-serine phosphate in alkaline solution as a function of  $H_- + \log C_{H_2O}$ .

mining step, such as the base-catalyzed elimination reaction of alkyl halides, forms olefins



The breaking of the SH bond may be simultaneous with, or prior to, the final breakdown to products. Since the breaking of the bond to hydrogen is the rate-determining step, these reactions are expected to exhibit a substantial primary kinetic isotope effect when this hydrogen is replaced by deuterium, and also to be general base catalyzed.

The rate expression in this case will be

$$-\frac{dC_{\text{SH}}}{dt} = k' \frac{f_{\text{OH}^-} f_{\text{SH}}}{f^*} C_{\text{OH}^-} C_{\text{SH}} \quad (22)$$

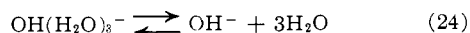
Multiplying 22 by  $a_{H_2O} = C_{H_2O}/C_{H_2O}$  and dividing by the same, one obtains the expression

$$-\frac{dC_{\text{SH}}}{dt} = k' \frac{f_{H_2O} f_{\text{SH}}}{f^*} \frac{a_{\text{OH}^-}}{a_{H_2O}} C_{H_2O} C_{\text{SH}} \quad (23)$$

which is identical with eq. 17, except for the values of the constants. The observed rate constant will therefore behave in the same way as in the previous case, *i.e.*, follow eq. 19 and 20. This similarity is not surprising in view of the similar composition of the transition states in both cases. When the conjugate base  $\text{S}^-$  undergoes a slow reaction with water, the transition

state can be depicted as  $\left[ \text{S} \cdot \text{O} \begin{array}{l} \text{H} \\ \text{H} \end{array} \right]^-$ , whereas on proton abstraction a transition state of the type  $[\text{S} \cdot \text{H} \cdot \text{OH}]^-$  is formed.

The same result is obtained if one assumes that the pre-equilibrium

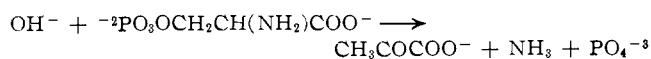


precedes the rate-determining step, and that the non-hydrated hydroxide ion is the reactive species; this leads to

$$k_{\text{obs}} = \text{const.} (C_{\text{OH}^-}/C_{H_2O}^3) = \text{const.} b_- C_{H_2O} \quad (25)$$

which is identical with eq. 20. The existence of such a pre-equilibrium is supported by the extremely high basicities of hydroxides dissolved in nonaqueous solvents like dimethyl sulfoxide and sulfolane.<sup>9</sup>

An example of mechanism III is the elimination reaction of *dl*-serine phosphate to form pyruvate, phosphate, and ammonia<sup>10</sup>



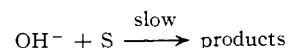
The replacement of the hydrogen on the  $\alpha$ -carbon by deuterium gives  $k_{\text{H}}/k_{\text{D}} \sim 2$  (at 100°) which demonstrates that the  $\alpha$ -proton is removed in the rate-determining step followed by rapid breakdown to products and that the reaction proceeds *via* mechanism III. A full discussion of the mechanism proposed, as well as the experimental procedure employed to follow the rate of the elimination, is given elsewhere.<sup>10</sup> The observed rate constant  $k_{\text{obs}}$  for the reaction as a function of NaOH concentration is given in Table I. These values of  $k_{\text{obs}}$  were computed from plots of the logarithm of substrate concentrations *vs.* time (with hydroxide in large excess). The values of  $H_-$  and  $H_- + \log C_{H_2O}$  at each concentration of base are presented in columns 3 and 4. The values of  $H_-$  are those measured by Schwarzenbach and Sulzberger,<sup>11</sup> corrected as stated in the previous paper.<sup>6</sup> Objections may be raised against the use of these values but, as mentioned in part I, the fact that several reactions show an almost ideal linear relationship with these values support their validity: Values of  $C_{H_2O}$ , free water concentration, were computed as given in Table I of part I and were thus expressed in  $M/55.5$ . A plot of the observed rate constants *vs.*  $H_- + \log C_{H_2O}$  is shown in Fig. 1 from which it is evident that the observed rate constants follow  $H_- + \log C_{H_2O}$  (slope = 0.98) in agreement with eq. 25. It should be noted that the empirical values of Schwarzenbach were measured at 20° whereas the kinetics of elimination were carried out at 100°, but the effect of temperature on acidity functions seems to be small (see ref. 2a, p. 945). In any case it is evident that the rate of this elimination reaction is much closer related to  $b_-$  than to the stoichiometric concentration of the hydroxide ion.

TABLE I

THE RATE OF HYDROLYSIS OF *dl*-SERINE PHOSPHATE IN AQUEOUS SODIUM HYDROXIDE SOLUTION

NaOH concn., M	$10^4 k_{\text{obs}}$ , sec. <sup>-1</sup>	$H_-$ (at 20°)	$H_- + \log C_{H_2O}$ (at 20°)
0.50	0.054	13.86	13.85
1.00	.104	14.26	14.24
2.00	.220	14.62	14.57
3.00	.415	14.95	14.87
4.00	.700	15.18	15.07
5.00	1.15	15.44	15.50
6.00	1.80	15.64	15.45

**Case IV.**—A nucleophilic attack on the substrate by  $\text{OH}^-$  in the rate determining step. The transition state in this case is rather similar to that of proton abstraction by  $\text{OH}^-$  (mechanism III), the difference being only in the point of attack.



The rate expression has consequently the same form as eq. 25 and can be derived in the same way. The ob-

(9) (a) D. J. Cram, B. Rickborn, C. A. Kingsbury, and P. Haberfeld, *J. Am. Chem. Soc.*, **83**, 3678 (1961); (b) C. H. Langford and R. L. Burwell, *ibid.*, **82**, 1503 (1960); (c) R. Stewart and J. P. O'Donnell, *ibid.*, **84**, 493 (1962).

(10) D. Samuel and B. L. Silver, *J. Chem. Soc.*, in press.

(11) G. Schwarzenbach and R. Sulzberger, *Helv. Chim. Acta*, **27**, 348 (1944).

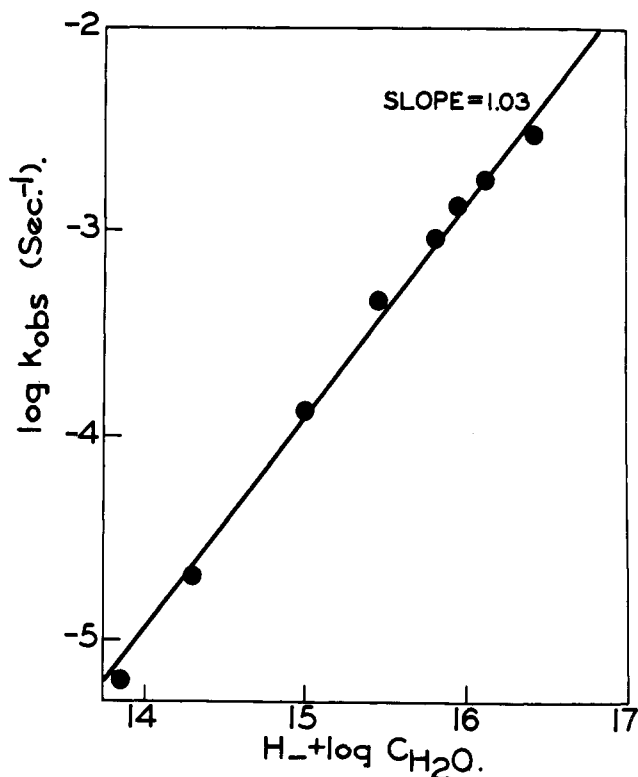
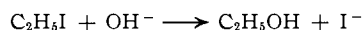


Fig. 2.—The rate of hydrolysis of ethyl iodide as a function of  $H_- + \log C_{H_2O}$ .

served rate constant is again expected to be proportional to  $b_-$ ;  $C_{H_2O}$ , or  $\log k_{obs} = H_- + \log C_{H_2O} + A$ ;  $A = \log \text{const.}$  A well established  $S_N2$  substitution was chosen to verify this conclusion, namely, the hydrolysis of ethyl iodide in aqueous alkaline solution<sup>12</sup>



The kinetics of this reaction have been measured using EtI labeled with  $I^{131}$  at  $10^{-3} M$  concentration owing to the low solubility of EtI in concentrated alkaline solutions. The experimental details used for following the rate of deiodination are described elsewhere.<sup>13</sup> The extent of EtI undergoing elimination to ethylene under our experimental conditions was determined by the addition of bromine labeled with  $Br^{82}$  at a known specific activity and extracting  $C_2H_4Br^{82}$  with ethylene bromide. The yield of olefin did not exceed 5% of the over-all deiodination products. The results are summarized in Table II and plotted *vs.*  $H$  and  $H_- + \log$

TABLE II

THE RATE OF THE HYDROLYSIS OF ETHYL IODIDE IN POTASSIUM HYDROXIDE SOLUTION

Initial concn. of EtI 0.001 M;  $T = 20^\circ$ .

KOH, M	$10^4 k_{obs}$ , sec. <sup>-1</sup>	$H_-$	$H_- + \log C_{H_2O}$
0.50	0.064	13.87	13.85
1.09	0.21	14.34	14.30
2.97	1.36	15.10	15.00
4.60	4.65	15.62	15.45
5.86	9.4	16.01	15.80
6.40	13.3	16.19	15.95
7.00	18.6	16.41	16.12
8.00	31.0	16.75	16.42

$C_{H_2O}$  in Fig. 2. The specific rate of hydrolysis of EtI follows  $b_- C_{H_2O}$  (slope 1.03) more closely than  $b_-$  (slope 0.93), as predicted by eq. 25.

(12) (a) G. H. Grant and C. H. Hinshelwood, *J. Chem. Soc.*, 258 (1933);  
(b) I. Dostrovsky, E. D. Hughes, and C. K. Ingold, *ibid.*, 173 (1946).  
(13) M. Anbar and M. Bobtelsky, to be published.

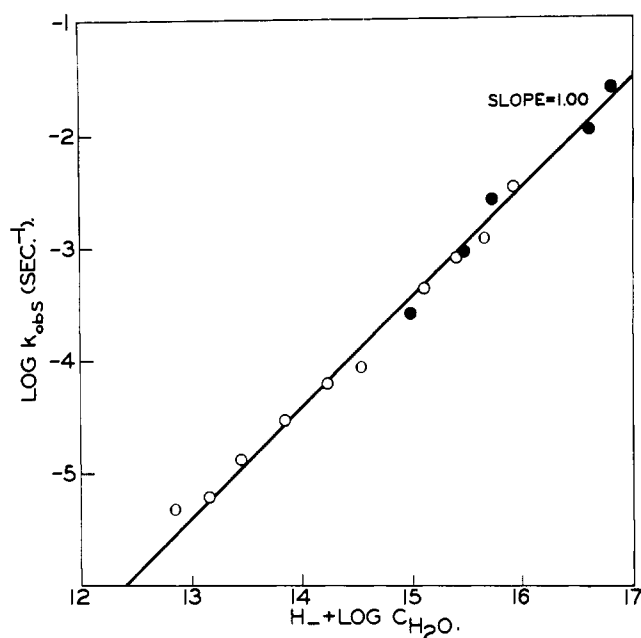
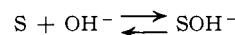


Fig. 3.—The rate of hydrolysis of chloramine in KOH (●) and NaOH (○) solutions as functions of  $H_- + \log C_{H_2O}$ .

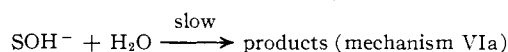
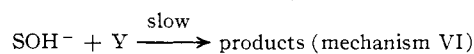
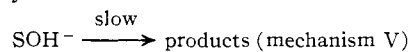
Another reaction investigated in the concentrated alkaline region is the hydrolysis of chloramine, and of dimethylchloramine,<sup>14</sup> and it was suggested that this reaction proceeds by nucleophilic attack by  $OH^-$ . Plotting  $\log k_{obs}$  *vs.*  $H_-$  gave a slope of 0.92, whereas when  $\log k_{obs}$  was plotted *vs.*  $H_- + \log C_{H_2O}$  (Fig. 3) a straight line of slope 1.00 was obtained as suggested by eq. 25 for substitution by  $OH^-$ . It should be noted that considerable experimental precision is required in order to make a decision whether  $H_-$  or  $H_- + \log C_{H_2O}$  ( $C_{OH^-}/C_{H_2O}^4$  or  $C_{OH^-}/C_{H_2O}^3$ ) are followed. It can be concluded, however, that the reactions discussed are not proportional to the stoichiometric hydroxide ion concentration.

Taking the model of the hydrated hydroxide ion, these results support the suggestion that the active nucleophilic species in this mechanism (as in III) is the *nonhydrated*  $OH^-$  ion, and that a pre-equilibrium (eq. 24) precedes the rate-determining step; this scheme leads directly to eq. 25. It has been found that in many nucleophilic substitutions the hydroxide anion is a weaker nucleophile than many much weaker bases<sup>14</sup>; this also supports the idea that the prevailing form of hydroxide in aqueous solutions, the hydrated anion, is not the most reactive form of this species.

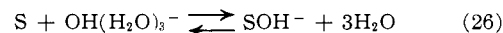
**Cases V, VI, and VIa.**—Rapid reversible hydroxide additions precede the rate-determining step



followed by



Mechanisms of this type are prevalent in base-catalyzed reactions in which carbonyl groups are involved, *e.g.*, ester hydrolysis.<sup>15,16</sup> When the hydration of the  $OH^-$  ion is considered, the pre-equilibrium to the addition may be presented by



(14) J. O. Edwards and R. G. Pearson, *J. Am. Chem. Soc.*, **84**, 16 (1962).  
(15) Cf. C. K. Ingold, "Structure and Mechanism in Organic Chemistry," Cornell Univ. Press, Ithaca, N. Y., 1953, p. 754.  
(16) M. L. Bender, *Chem. Rev.*, **60**, 53 (1960).

The concentration of  $\text{SOH}^-$  would best be given by the appropriate indicator function; following the argument of part I, it can also be given by

$$C_{\text{SOH}^-} = K_c' (C_{\text{OH}^-} / C_{\text{H}_2\text{O}^3}) C_S \quad (27)$$

$K_c'$  being the pseudo-equilibrium constant of eq. 26. This leads to the following expressions for the observed rate constants:

Mechanism V, unimolecular breakdown of  $\text{SOH}^-$ , giving

$$k_{\text{obs}} = - \frac{1}{C_S} \frac{dC_S}{dt} = \text{const.} \frac{C_{\text{OH}^-}}{C_{\text{H}_2\text{O}^3}} = \text{const.} b_- C_{\text{H}_2\text{O}} \quad (28)$$

mechanism VI, bimolecular reaction of  $\text{SOH}^-$  with a second reactant Y, giving

$$k_{\text{obs}} = - \frac{1}{C_Y C_S} \frac{dC_S}{dt} = \text{const.} \frac{C_{\text{OH}^-}}{C_{\text{H}_2\text{O}^3}} = \text{const.} b_- C_{\text{H}_2\text{O}} \quad (29)$$

and mechanism VIa, a special case where  $Y = \text{H}_2\text{O}$ , giving

$$k_{\text{obs}} = - \frac{1}{C_S} \frac{dC_S}{dt} = \text{const.} \frac{C_{\text{OH}^-}}{C_{\text{H}_2\text{O}^2}} = \text{const.} b_- C_{\text{H}_2\text{O}}^2 \quad (30)$$

Reactions proceeding by mechanisms V and VI will thus follow  $b_- C_{\text{H}_2\text{O}}$  or  $C_{\text{OH}^-} / C_{\text{H}_2\text{O}^3}$  as in mechanisms IIa, III, and IV. Mechanism VIa, however, differs from all others by being dependent on the *second* power of the free water concentration (eq. 30). Here again it is to be regretted that no experimental rate data regarding these three mechanisms in sufficiently concentrated alkaline solutions are available at present.

### Discussion

When the formulation of rate expressions of the eight types of base-catalyzed reactions presented is summarized, these may be divided into three groups: The rate of reaction in cases I and II can be expected to follow  $H_-$  or  $C_{\text{OH}^-} / C_{\text{H}_2\text{O}^4}$ ; in cases IIa, III, IV, V, and VI to follow  $H_- + \log C_{\text{H}_2\text{O}}$  or  $C_{\text{OH}^-} / C_{\text{H}_2\text{O}^3}$ ; and in case VIa to follow  $H_- + 2 \log C_{\text{H}_2\text{O}}$  or  $C_{\text{OH}^-} / C_{\text{H}_2\text{O}}$ . In principle this enables a distinction to be made between these three groups. However, several complicating factors have to be considered which might make such a distinction difficult in practice. The main factor among these is the possibility that water is involved in formation of the transition state in additional ways than those considered above. Consider, for example, a case where a conjugate base decomposes unimolecularly (*i.e.*, case I), but the transition state is stabilized due to hydration by four water molecules, which were not in the hydration shell of the substrate SH. Following the approach discussed above, *no* dependence on the concentration of water is expected in this case

$$- \frac{dC_{\text{SH}}}{dt} = k'' C_S C_{\text{H}_2\text{O}^4} = k'' K_c C_{\text{SH}} \frac{C_{\text{OH}^-}}{C_{\text{H}_2\text{O}^4}} C_{\text{H}_2\text{O}^4} = k'' K_c C_{\text{OH}^-} C_S$$

Another example is the case where the *hydrated* hydroxide and not the bare ion is reactive, in either elimination or nucleophilic substitution, leading again to rates proportional to  $C_{\text{OH}^-}$  and zero order in  $C_{\text{H}_2\text{O}}$ .

It may be concluded, therefore, that *the dependence of the rate of a base-catalyzed reaction on  $C_{\text{H}_2\text{O}}$  (or on  $H_-$ ) can only be used as an indication of the number of water molecules involved in the formation of the transition state and any further conclusion on the mechanism of reaction requires independent information* about the role of water in the reaction. The best way to evaluate the rate of a base-catalyzed reaction is to plot  $k_{\text{obs}} / C_{\text{OH}^-}$  vs.  $\log C_{\text{H}_2\text{O}}$ ; the slope of such a plot indicates the probable number of water molecules involved. It does not indicate with which reactant or intermediate these water molecules were associated, nor in which steps they were lost or gained.

This is of course true only if the following further effects are not interfering: (a) The existence of an appreciable amount of substrate in the base form; if the  $pK$  is known, this can readily be corrected. (b) Above 7 molar hydroxide, an increasing proportion of the hydroxide ion can no longer be in the trihydrated form. (c) Bound water too may be active, *e.g.*, as nucleophile; in such a case it is necessary to introduce additional rate terms. (d) Cation catalysis: if the cation has a catalytic effect in a base-catalyzed reaction one may expect an apparent higher order in  $C_{\text{OH}^-}$  superimposed on the  $C_{\text{H}_2\text{O}}$  dependence. (e) Additional "salt effects"—positive or negative—can be expected when changes of total charge or dipole occur in the formation of the transition state, *e.g.*, when a positively charged reactant is involved. In view of these limitations, it is of interest that all five base-catalyzed reactions discussed in this paper follow rate expressions of simple dependence on  $C_{\text{OH}^-}$  and  $C_{\text{H}_2\text{O}}$ . It seems that water is actually involved in these reactions only as hydration water of the  $\text{OH}^-$  ion, which becomes free upon formation of the transition state.

In the final analysis, it seems that only further experimentation will decide whether the simple model picturing concentrated hydroxide solutions as containing only two kinds of water, bound and free, is sufficient to describe rate processes in these media adequately. The experimental data indicate, however, that a consideration of the role of water is essential in evaluating rate data in concentrated alkaline solutions.

**A Comment Regarding Acid Solutions.**—The conclusions regarding alkaline solutions are also applicable to concentrated acid solutions. The interpretation of rate data in these solutions has been regarded as unsatisfactory for some time<sup>17-21</sup> and two of the main features of the present treatment have already been suggested for concentrated acid solutions: Leisten<sup>22</sup> proposed the use of the concept of free water based on the suggestions of Bascombe and Bell, and an assumption analogous to eq. 13 has been suggested by Koskikallio and Whalley.<sup>23</sup> While this paper was in preparation, a detailed theory of reaction rates in concentrated acid solutions was published by Bunnett.<sup>24</sup> His approach resembles the present one in that order of reaction in water is suggested as the rate-determining factor. However, Bunnett's treatment differs from ours in suggesting that the thermodynamic *activity* of water be used as a measure of its kinetic activity instead of the *concentration* of free water, while we have preferred to include the activity coefficient in expression 18. However, when using water activity, reactions whose rate follows  $h_0$  are no more linear in  $a_{\text{H}_2\text{O}}$ , and the approximate order in water ( $w^*$ ) is often as high as  $-7$  or  $-8$ . This seems an unlikely large number of water molecules, and is out of line with Bascombe and Bell's picture, which requires that about four water molecules are involved. It seems, therefore, that the use of free water concentration will also prove helpful in interpreting rate measurements in acid solutions and provide a closer indication to the actual number of water molecules involved in these reactions.

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