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Effects of Ureas on the Kinetics of Acid Hydrolysis of 1-Phenylethyl Nitrite in Water and in Aqueous Micellar Solutions of Sodium Dodecyl Sulfate

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The acid ionization constants of urea, methylurea (MEU), ethylurea (ETU), *N*,*N*-dimethylurea (DMU) and tetramethylurea (TMU) have been determined in water and in aqueous micellar solutions of sodium dodecyl sulfate (SDS) from the kinetic study of the acid hydrolysis of 1-phenylethyl nitrite (PEN). The acid hydrolysis in water was inhibited by the addition of urea, as a result of its protonation which decreases the effective H^+ concentration. Analysis of the solvent isotope effects and the activation parameters shows that urea does not modify the solvation of the substrates. MEU and ETU show a similar behaviour to that found with urea, but TMU behaves slightly differently, because TMU decreases the relative permittivity of the water mixtures.

The rate of acid hydrolysis in aqueous SDS goes through a maximum with increasing SDS concentration. The addition of urea and urea derivatives strongly diminishes the value of the attained maximum in such a way that at high concentrations of ureas the observed rate constant continually decreases as the SDS concentration increases. The kinetic data were analysed by using the simple pseudophase ion-exchange (PPIE) model. The results are explained under the assumption that ureas do not modify the structure of the micelles, at least not for the quantities used in this work, except in the case of TMU and DMU. These two ureas increase the solubility of PEN in water and strongly decrease the value of the rate constant achieved in the micellar phase.

In recent years, several studies have attempted to test the effect of urea on the properties of aqueous micellar,[1–3] protein[4] or cyclodextrin[5] solutions and microemulsions.[6] The addition of urea to water increases both the relative permittivity[7] and the surface tension. On the other hand, urea is a weak base.[8]

The effect of urea on hydrophobic interactions is seen in a variety of experimental situations. The low solubility of any non-polar species in water is viewed as a consequence of the high relative permittivity or the large cohesive energy density of water. Nevertheless, urea enhances the solubility of such species in water,[9,10] despite the fact that its addition increases the relative permittivity of aqueous solutions. The effect of urea upon water structure is a controversial matter.[11–14] Traditionally, the effects of urea on protein denaturation and hydrocarbon solubilization have been attributed to its disrupting the long-range order of water. More recently, urea is thought to replace some of the water molecules that solvate the hydrophobic groups of the solutes.

The properties of micellar solutions depend on the balance between 'hydrophobic' and 'hydrophilic' interactions. In aqueous solutions this balance can be altered in a number of ways, which include (a) addition of certain substances that can be solubilized into the micelle;[15] (b) counterion complexation, particularly for ionic surfactants; and (c) modification of the properties of the aqueous solvent through the addition of electrolytes, such as simple salts, or non-electrolytes, such as urea and its derivatives.[16]

De Lisi and co-workers[1] found that the addition of urea to micellar solutions leads to an increase of both the critical micelle concentration (c.m.c.) and the degree of counterion dissociation (α) of dodecyl trimethylammonium bromide (DTAB). The authors suggest that urea does not penetrate the micellar surface, thus attributing the effect of added urea to the change in physicochemical properties of the solvent mixture. On the other hand, the results obtained by Baglioni *et al.*[3] from studying the action of urea on SDS or DTAB micellar solutions by the use of aminoxyl spin probes suggest that urea is dissolved at the micellar surface resulting in a decrease of the polarity and a strong increase in the microviscosity of the micellar interface. Other studies of fluorescence measurements support the direct mechanism of urea action, by which organic molecules in hydrophobic surfaces are replaced by urea in aqueous solutions of SDS and Triton-X micelles,[17] cyclodextrins[6] and protein[5] surfaces. These experimental studies suggest that urea weakens the hydrophobic interaction, which plays a dominant role in the formation of higher-order structures of folded globular proteins and molecular assemblies, such as micelles.

One might expect that the presence of urea would also have some influence on the kinetics of chemical reactions taking place in aqueous solutions of micelle-forming surfactants, or that kinetic studies might afford some additional information on urea action. To our knowledge, very few studies have been dedicated to investigating this kinetic aspect.[18] Therefore, we have examined the influence of urea and some alkylureas on the structure and properties of SDS micelles by studying a well known reaction, the acid hydrolysis of PEN, in aqueous and micellar media.

Experimental

Materials

PEN was synthesized by treating 1-phenylethanol with sodium nitrite in aqueous sulfuric acid,[19,20] purified by fractional distillation, and stored at low temperatures over molecular sieves (3 Å) to prevent its hydrolysis. Ureas (Merck or Aldrich) of the maximum purity commercially available were recrystallized twice from hot ethanol and were stored under desiccating conditions. Urea solutions were prepared daily immediately before measurements were taken.[21] SDS of the highest available purity was supplied by Sigma, and D_2O (99.77%) was supplied by the Spanish Nuclear Energy Board. All other reagents were of the maximum purity commercially available and were used as received.

Conductance

Conductance measurements were taken in a Radiometer conductimeter using solutions prepared with doubly distilled water. The cell constant (1.024 ± 0.003 cm^{-1}) was determined at 25 °C by measuring the conductance of dilute solutions of aqueous KCl and using the equation of Lind *et al.*[22] The solutions were thermostatted in the conductivity cell at 25 ± 0.1 °C, equipped with a magnetic stirring device.

Kinetics

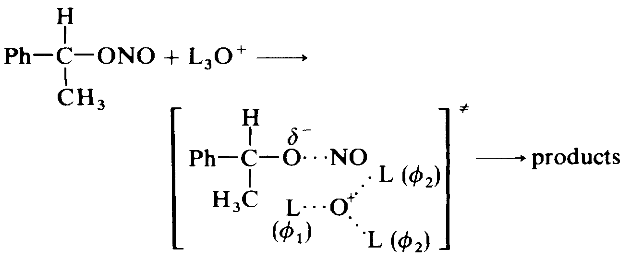
The kinetics of the reaction were studied by using a Hi-Tech stopped-flow apparatus, noting the decreasing absorbance due to the consumption of alkyl nitrite in the 245–250 nm region, except in the experiments carried out in the presence of tetramethylurea, in which case the reaction was followed at 260 nm. The kinetic procedure has been described elsewhere.[20]

The products of the acid hydrolysis of alkyl nitrites are the equivalent alcohol and nitrous acid, the latter being an effective nitrosating agent of amines,[23] amides and ureas.[24] This subsequent reaction was observed in the case of methyl- and ethyl-urea; nevertheless, this reaction never interfered with the acid hydrolysis of the alkyl nitrite, a very fast process lasting, in some of the experimental conditions of the present work, only a fraction of a second.

Results and Discussion

Reaction in Water

The acid hydrolysis of several alkyl nitrites in aqueous medium has been studied previously.[19] The results suggest that the reaction takes place through a concerted mechanism that involves both proton transfer and the breaking of the N—O bond, the latter facilitated by the presence of electron-withdrawing groups. However, there is no evidence for perfect synchronization between the proton transfer and N—O bond breakage; rather the results suggest an imbalanced transition state in which a slight negative charge develops on the O atom owing to the presence of electron-withdrawing groups, thus explaining the experimental observations, including the isotope effects and general acid catalytic behaviour (see Scheme 1).



Scheme 1

In order to study the effects of ureas on the properties of SDS micelles, the kinetics of a well known reaction were investigated. Before the results obtained in a micellar medium could be evaluated, it was necessary to analyse the influence of ureas on this reaction in water. Accordingly, we studied the hydrolysis of PEN by hydrochloric acid in light and heavy water, and in the absence and presence of urea.

Fig. 1 shows the influence of acidity in water and in D_2O due to the absence and presence of 1.0 mol dm^{-3} of urea. In the four cases, the influence of acidity (controlled with HCl) leads to the linear relationship:

$$k_0 = k_2[L^+]; \quad L^+ = K^+ \text{ or } D^+ \tag{1}$$

Table 1 lists the values for k_2 obtained by studying the influence of acidity under the various experimental conditions. The solvent isotope effect calculated in the absence of urea, $k_2^H/k_2^D = 0.89$, can be understood in terms of fractionation factor theory.[25] Given that L_3O^+ is the proton donor, then $k_2^H/k_2^D = l^3/(\phi_1\phi_2^2)$, where l (=0.69) is the fractionation factor of the three sites in the reactant L_3O^+; ϕ_1 corresponds to the hydrogen being transferred in the transition state, and ϕ_2 is the fractionation factor of the other two hydrogens, which are in an intermediate state between that of L_3O^+ and that of water ($\phi_2 = 1$) (see Scheme 1). The fractionation factor of RONO is assumed to be unity, as is usual for uncharged substrates.[26]

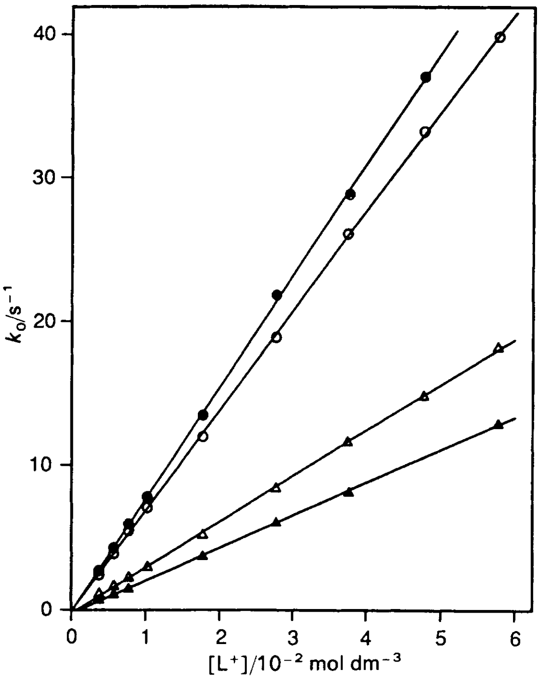


Fig. 1 Influence of acidity upon the pseudo-first-order rate constant for the acid hydrolysis of PEN in water (open symbols) and in D_2O (solid symbols): (○, ●) in the absence and (△, ▲) in the presence of 1.0 mol dm^{-3} urea

Assuming that the fractional extent of hydrogen transfer in the transition state measures the same properties of the transition state as the Brønsted exponent,[25] then $\phi_2 = l^{1-\alpha}$, and, in agreement with our previous study, a value of $\phi_2 \approx 0.92$ could be determined starting from the experimental value of the isotope effect. This value conforms perfectly to the form of the postulated transition state of our system: a concerted process ($0.4 < \phi_1 < 0.7$)[27] with an imbalanced transition state ($\alpha > 0.5$, $\alpha \approx 0.66$ in our case) in which an important negative charge has been developed on the O atom prior to proton transfer.

The isotope effect achieved in the presence of urea indicates that urea participates in the reaction process more than simply modifying the properties of the solvent. As shown in Fig. 2, the acid hydrolysis of PEN depends significantly on the relative permittivity of the medium, reducing the observed

Table 1 Second-order rate constants obtained in the acid hydrolysis of PEN in different media

medium	k_2/dm^3 mol^{-1} s^{-1}
H_2O	695 ± 3[a]
D_2O	777 ± 5[a]
H_2O–1.0 mol dm^{-3} urea	316 ± 2[b,c]
D_2O–1.0 mol dm^{-3} urea	226 ± 2[b,c]
H_2O–0.5 mol dm^{-3} DMU	369 ± 2[c]
H_2O–0.5 mol dm^{-3} TMU	385 ± 7[c]

[a] $k_2^H/k_2^D = 0.89$. [b] $k_1^H/k_1^D = 1.40$. [c] Corresponds to the value of k_1 in eqn. (2).

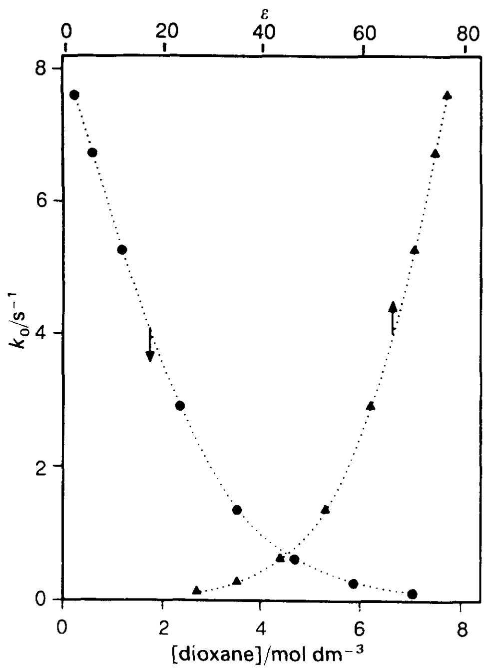


Fig. 2 Variation of the pseudo-first-order rate constant corresponding to the acid ([HCl] = 1.27×10^{-2} mol dm^{-3}) hydrolysis of PEN with (●) dioxane concentration and (▲) relative permittivity of the medium, determined following ref. 40. The dotted lines are to guide the eye.

rate constant with ε, the polarity of the medium. Analysis of the influence of urea concentration at constant [H$^+$] leads to a similar dependence of k_0 on [urea] [see Fig. 3(*a*)]. However, this behaviour cannot be attributed to a decrease in the polarity of the reaction medium, because urea raises slightly the relative permittivity of water ($\Delta\varepsilon/\Delta$[urea] $\approx$ 3 dm^3 mol^{-1} at 25 °C).[7] Both the extent of the isotope effect determined in the presence of urea as well as the variation of k_0 with [urea] can be explained by considering the basic character of urea, and this is shown in Scheme 2.

$$\mathrm{UH^+ \rightleftarrows U + H^+}, \qquad K_{\mathrm{UH}}$$

$$\mathrm{RONO + H_3O^+ \rightarrow ROH + HNO_2 + H^+}, \qquad k_2$$

Scheme 2

Here, UH$^+$ represents the protonated form of urea, U the unprotonated form and K_{UH} is the acid ionization constant of UH$^+$. From Scheme 2, and taking into account that $[\mathrm{H^+}]_t = [\mathrm{H^+}] + [\mathrm{UH^+}]$ and $[\mathrm{U}]_t = [\mathrm{U}]$ (since in every case $[\mathrm{U}] > 10[\mathrm{H^+}]$), one may obtain the following expression for the observed rate constant, k_0:

$$k_0 = \frac{k_2 K_{\mathrm{UH}}}{K_{\mathrm{UH}} + [\mathrm{U}]} [\mathrm{H^+}]_t = k_1[\mathrm{H^+}]_t \qquad (2)$$

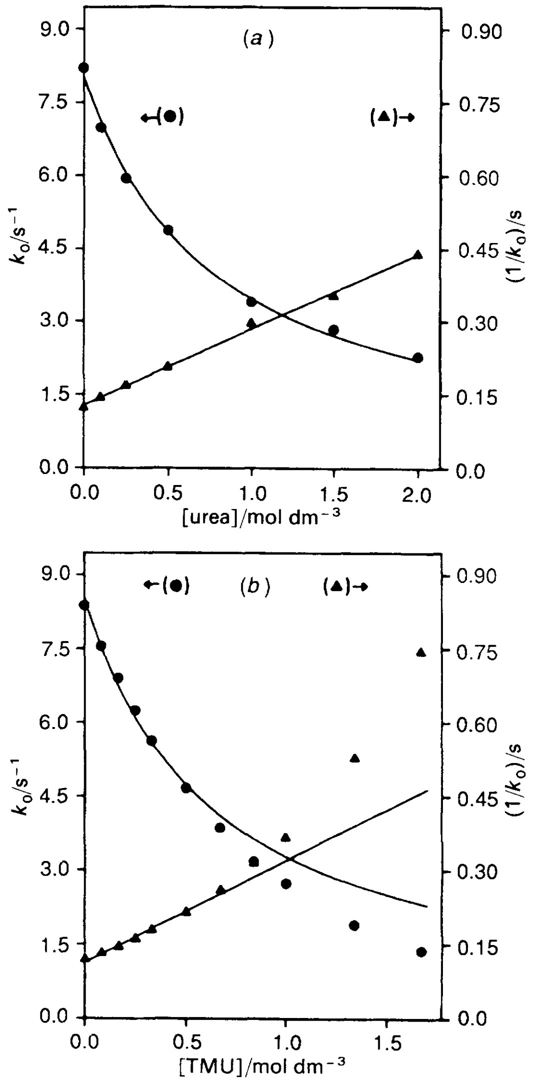


Fig. 3 Influence of (*a*) urea and (*b*) tetramethylurea concentration upon the pseudo-first-order rate constant for the acid hydrolysis of PEN in water at [HCl] = 0.0127 mol dm^{-3} (●). Plot of the data in the form of $1/k_0$ *vs.* [urea], eqn. (2) (▲). The solid lines are the fit to eqn. (2).

This equation predicts a linear dependence of k_0 on the total concentration of H$^+$ in the absence of urea or at constant urea concentration, as the experimental results show (see Fig. 1), and an inverse variation with urea concentration at constant $[\mathrm{H^+}]_t$ (*i.e.* $1/k_0$ *vs.* [urea] is linear). As can be seen from the data in Fig. 3(*a*), this is the result observed experimentally. The values of k_2 and of K_{UH} obtained for the acid

Table 2 Bimolecular rate constants (k_2) and the acid ionization constants (K_{UH}) of ureas obtained in the acid hydrolysis ([HCl] = 1.27×10^{-2} mol dm^{-3}) of PEN in aqueous medium in the presence of ureas determined by fitting eqn. (2) to the experimental data

urea	k_0/s^{-1}	k_2/dm^3 mol^{-1} s^{-1}	K_{UH}/mol dm^{-3}	range[a]
—	8.2	695	—	—
urea	8.1 ± 0.4[b]	639 (737)[c]	0.75 ± 0.03	0–2
urea	—	—	1.12,[d] 0.79,[e] 0.63[f]	—
methylurea	8.3 ± 0.3[b]	653	1.30 ± 0.03	0–0.5
ethylurea	8.2 ± 0.1[b]	646	0.64 ± 0.07	0–0.5
DMU	8.5 ± 0.5[b]	670 (676)[c]	0.60 ± 0.03	0–0.8
TMU	8.5 ± 0.7[b]	670 (692)[c]	0.63 ± 0.04	0–0.6

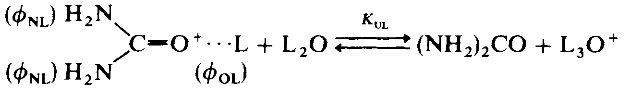
[a] Range of [urea], in mol dm^{-3}, used to obtain the kinetic data. [b] Determined from fitting eqn. (2) to the experimental data obtained at [HCl] = 0.0127 mol dm^{-3}. [c] Values obtained from the study of the influence of the acidity at fixed [urea]: [urea] = 1.0, [DMU] = 0.50 and [TMU] = 0.50 mol dm^{-3}. [d] Ref. 39(*a*). [e] Ref. 39(*b*). [f] Ref. 39(*c*).

hydrolysis of PEN in aqueous media in the presence of urea and alkylureas are displayed in Table 2. The determined value for the acidity constant of the protonated urea agrees perfectly with those of other studies, which are also shown in Table 2. In the case of alkylureas, we failed to find relevant data in the literature; nevertheless, the values found in this study, by all appearances, seem quite reasonable. Tetramethylurea has a relative permittivity of $\varepsilon = 23.6$ at 25 °C.[28] Thus, the addition of TMU to water decreases the relative permittivity of the corresponding mixtures. This effect was reflected in the kinetic results obtained in the presence of TMU. As we can see in Fig. 3(*b*) the plot of $1/k_0$ *vs.* [TMU] does not show a linear dependence in the whole interval of [TMU] studied, because, in this case, two factors make k_0 decrease as [TMU] increases: the decrease in the relative permittivity of the medium and the reduction of the effective $[H^+]$. A similar behaviour is observed with DMU, which can also be attributed to the effect of the relative permittivity.

Meanwhile, the value obtained experimentally for the isotope effect in the presence of urea and according to eqn. (2), can be given by:

$$\frac{k_1^H}{k_1^D} = \left(\frac{k_2^H}{k_2^D}\right)\left(\frac{K_{UH}^H}{K_{UH}^D}\right)\left(\frac{K_{UH}^D/K_{UH}^H + [U]/K_{UH}^H}{1 + [U]/K_{UH}^H}\right) \qquad (3)$$

In the absence of urea, k_2^H/k_2^D was 0.89; K_{UH}^D/K_{UH}^H can be estimated [according to reaction (I), below] from the corresponding fractionation factors of the hydrogenic position as $l^3\phi_{urea}/\phi_{OL}(\phi_{NL})^2$, in which ϕ_{OL} is the fractionation factor of the L^+ bound to the O atom of the urea molecule, and in which ϕ_{NL} corresponds to the L^+ bound to the N atom of the urea molecule. Taking the values of such fractionation factors from ref. 29, one can determine a value of 0.51 for the ratio of the acid ionization constants of urea in D_2O and in water. With these values, the result $k_1^H/k_1^D = 1.37$, in the presence of urea, agrees perfectly with the value found experimentally and included in Table 1.



$$\begin{matrix}(\phi_{NL})\,H_2N \\ (\phi_{NL})\,H_2N\end{matrix}\!\!>\!C{=}\underset{(\phi_{OL})}{O^+\cdots L} + L_2O \xrightleftharpoons{K_{UL}} (NH_2)_2CO + L_3O^+$$

(I)

Influence of Temperature

The reaction was studied at 20, 25, 30 and 35 °C in the absence and presence of 1 mol dm^{-3} of urea. The thermodynamic results corresponding to the reaction in water were determined from the values of k_2^w obtained at the above-mentioned temperatures and are listed in Table 3, together with the activation parameters for the reaction carried out in an aqueous SDS micellar medium and determined from the values of k_m obtained at 25 and 35 °C (see below).

We can observe that the presence of urea has no effect on the degree of $\Delta H^‡$, meaning that, from a mechanistic point of view and in terms of what is understood by the stabilization of reactants or the transition state, urea stabilizes the reactants and the transition state to a similar extent.

It is difficult to express the significance of the standard entropy of activation, but it has been described vaguely as a probability or geometrical term. For bimolecular reactions which go through a highly concerted mechanism (or if the breaking and forming of bonds are synchronous), negative entropies of activation are often found. The $\Delta S^‡$ values found for our system, could indicate an important asymmetry of the transition state. A similar conclusion was reached in interpreting the observed solvent isotope effects. The acid hydrolysis of alkyl nitrites must be a case of solvent-induced bimolecular substitutions.[30] The results also signify that the presence of urea does not affect the entropy of activation, which may imply that the urea does not modify the solvation state of the reactants or of the activated complex. This conclusion would not favour a direct mechanism of urea action, according to which urea replaces some water molecules that solvate the hydrophobic substrates.

Table 3 Activation parameters for the acid hydrolysis of PEN in water and in SDS micelles both in the absence and presence of urea calculated from the influence of temperature upon k_2^w and k_m

medium	$\Delta H^‡$/kJ mol^{-1}	$\Delta S^‡$/J mol^{-1} K^{-1}
H_2O	64.0	23.8
urea[a]	66.3	24.1
SDS	52.7	−36.4
SDS–urea[a]	61.7	−19.2

[a] [urea] = 1.0 mol dm^{-3}.

Reaction in SDS Micellar Medium

Conductance Measurements

Conductivity measurements were performed in the absence and presence of urea and alkylureas. The results show that the degree of ionization, α, of SDS micelles is hardly changed in the presence of urea, methylurea and ethylurea at concentrations $\leqslant 1$ mol dm^{-3}. DMU and TMU produce a higher effect on α and the c.m.c. values; *e.g.* when [TMU] = 0.50 mol dm^{-3}, $\alpha = 0.65$ and c.m.c. = 1.07×10^{-2} mol dm^{-3}, and when [TMU] = 1.0 mol dm^{-3}, $\alpha = 0.76$ and c.m.c. = 1.40×10^{-2} mol dm^{-3}. (The same values for pure SDS solutions are reported in the literature as $\alpha = 0.25$ and c.m.c. = 8.2×10^{-3} mol dm^{-3}.) Fig. 4 shows the corresponding results together with the variation of electrical conductance with [SDS] in the presence of fixed concentrations of H^+ ([HCl] = 6.0×10^{-3} mol dm^{-3}) with and without ethylurea ([ETU] = 0.50 mol dm^{-3}). As can be seen, at SDS concentrations higher than its c.m.c. value, the specific conductivity first decreases, both in the presence and in the absence of ureas, and then increases again. This is due to H^+ bonding at the micelle on exchange with Na^+, and because of the lower mobility of Na^+ with respect to H^+, the observed conductivity decreases. The conductance of the solution mixture, in mS cm^{-1}, at [SDS] > c.m.c. can be expressed by:[31]

$$\kappa_{mix} = \kappa_0 + (\lambda_{Na} - \lambda_H)[H^+]_m + \alpha(\lambda_{Na} + \lambda_{mic})[SDS]_m \qquad (4)$$

where the λ_i are equivalent conductivities, ionic charges have been omitted in the subscripts for the sake of simplicity; $\kappa_0 = (\lambda_H + \lambda_{Cl})[HCl] + (\lambda_{Na} + \lambda_{SD})$c.m.c.; $[SDS]_m$ is the micellized surfactant concentration, and $[H^+]_m$ represents the H^+ concentration bound to the micellar surface, which can be determined as it is frequently reported in the literature,[32] and for low acid concentrations nearly 80% of hydrogen ions are incorporated into the micelle.[31] At low [SDS], the second (and negative) term predominates over the third, and the conductance decreases.

The presence of ureas allows for two observations: first, the experimental conductivity is lower in the presence of ureas, which is attributable to UH^+ formation; secondly, the extent of reduction of the observed conductivity on increasing the SDS concentration above the c.m.c. is less in the presence of ureas, which can be explained if we assume that UH^+ associates at the SDS micellar surface, thus displacing the H^+ or

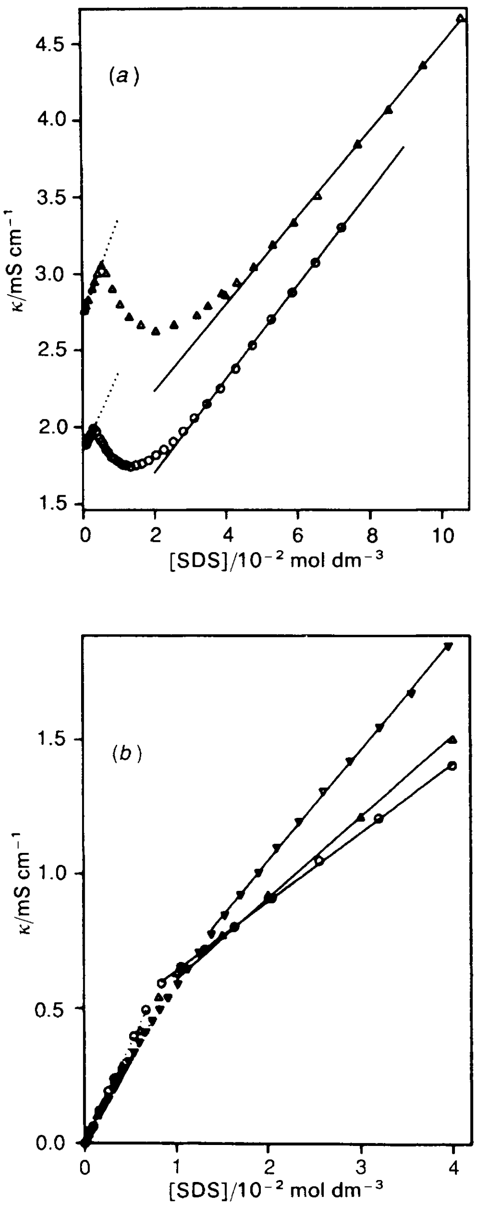


Fig. 4 Specific conductivity *vs.* SDS concentration (*a*) in the presence of [HCl] = 6.0 × 10^{-3} mol dm^{-3} with (△) no added ureas, (○) [ETU] = 0.50 mol dm^{-3}, and (*b*) in the absence of HCl with (▽) [TMU] = 0.50 mol dm^{-3}, (△) [urea] = 1.0 mol dm^{-3} and (○) no added ureas

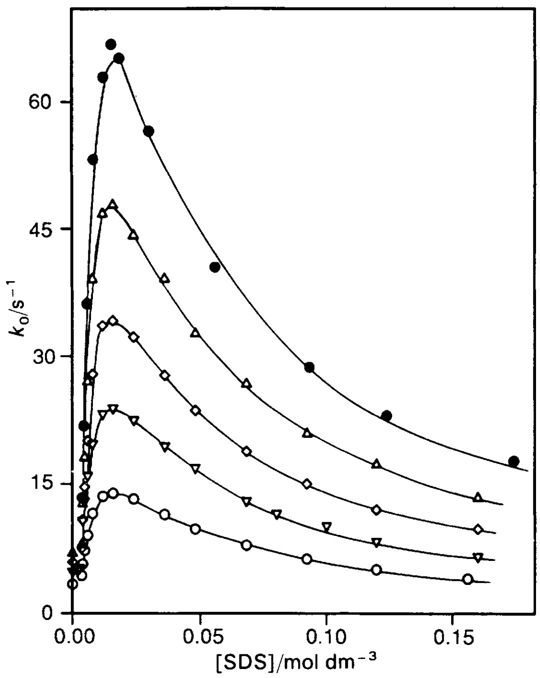


Fig. 5 Influence of SDS concentration on the acid hydrolysis of PEN at [urea] = 0.0 (●); 0.10 (△); 0.25 (◇); 0.50 (▽) and 1.0 mol dm^{-3} (○)

Na$^+$ ions, that is, K_{UH} decreases in the presence of SDS micelles.

Kinetic Measurements

The influence of SDS and urea concentrations on the reaction rate was studied in the presence of a fixed HCl concentration (1.27 × 10^{-2} mol dm^{-3}). Fig. 5 shows typical results of the effect of SDS and urea concentrations on the acid hydrolysis of PEN. Fig. 6 shows the comparative effect found with urea and the alkylureas used in the present work.

The degree of catalysis in each case reaches a maximum with SDS concentration and then decreases as the urea concentration increases. The decrease in the catalytic behaviour is more marked upon increasing the hydrophobic nature of the urea. This characteristic pattern of behaviour may be explained in terms of the pseudo-phase ion-exchange model. It is based on two related key assumptions:[33] (i) the micellar fractional charge, α $(=1-\beta)$, is constant, and (ii) the inert concentrations of the surfactant (Na$^+$ in our case) exchange on a 1 : 1 basis with the reactive ions (H$^+$ in the present case), K_{Na}^{H} being the corresponding equilibrium constant (in the expression for K_{Na}^{H} the concentrations of the ions in the aqueous and micellar pseudo-phases are given as molarities in terms of total solution volumes).

The kinetic data were analysed quantitatively by taking into account (1) the partitioning of PEN between the micellar and aqueous phase, that is, $PEN_w + D_n \rightleftarrows PEN_m$, where D_n is the micellized surfactant, and K_s^N is the corresponding equi-

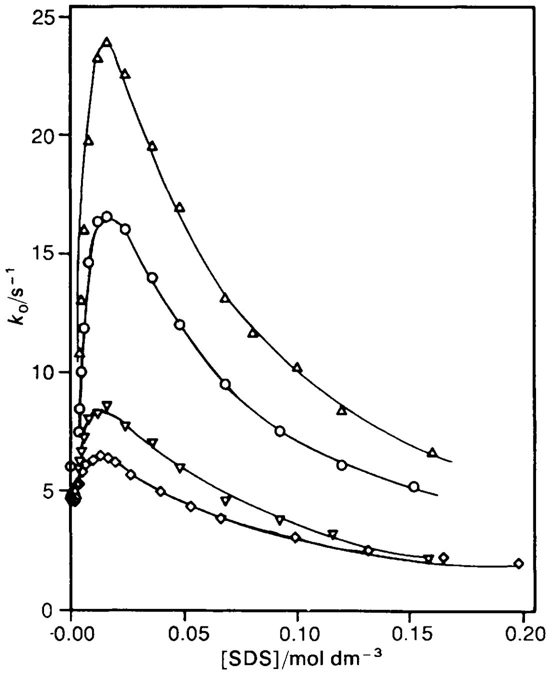


Fig. 6 Influence of [SDS] on the acid hydrolysis of PEN at 0.50 mol dm^{-3} of (△) urea; (○) methylurea; (▽) ethylurea and (◇) tetramethyl urea

librium constant determined as 79 mol^{-1} dm^{-3} in the absence of ureas;[20] (2) the selective ion exchange of the reactive ions with inert counterions, for which K_{Na}^{H} varies over a typical range of 0.6–1,[32] (3) the independent reactivities in the micellar and aqueous phases, and (4) the reduction of $[H^+]$ as a consequence of the protonation of ureas.

Given these considerations, one may use eqn. (5), see below, to attain the apparent acidity constant of ureas in the presence of SDS micelles. In this equation, the quantities in square brackets are molarities referred to the total solution volumes and the subscripts w and m denote aqueous and micellar pseudophases, respectively.

$$K_{UH}(ap) = [U]_w \frac{[H^+]_w + [H^+]_m}{[UH^+]_w + [UH^+]_m} \qquad (5)$$

On the other hand, if as in the experimental conditions of the present work, $[U] \gg [H^+]$, then $[U]_w = [U]_t$, the analytical concentration of urea, which we are assuming to reside only in the aqueous phase; and by analogy, the analytical concentration of H^+ can be given by $[H^+]_t = [H^+]_w + [H^+]_m + [UH^+]_w + [UH^+]_m$.

Quantitative explanation of the experimental data is not difficult if $K_{UH}(ap)$ does not vary appreciably with [SDS]. If we assume this and keep in mind that the overall reaction rate is the sum of the reaction rates in the aqueous and in the micellar pseudo-phases, the observed rate constant can be expressed as:

$$k_0 = \frac{K_{UH}^{(ap)}}{K_{UH}^{(ap)} + [U]_t} \, \frac{k_2^w[H^+]_t + \{(k_2^m/\bar{V})K_s^N - k_2^w\} m_H [D_n]}{1 + K_s^N[D_n]} \qquad (6)$$

Here, k_2^w and k_2^m refer to the bimolecular reaction constants in the aqueous and micellar pseudo-phases, respectively, and $\bar{V}$ is the volume of the micellar pseudo-phase per mole of micellized surfactant, which has been estimated as 0.14 dm^3 mol^{-1} by Bunton *et al.*[34] by identifying the reaction zone of the micellar pseudo-phase with the Stern layer; and $m_H = [H^+]_m/[D_n]$ can be obtained.[35]

Adapting the data in this way is possible if $K_{UH}(ap)$ does not vary appreciably with [SDS]. Fig. 7 presents the variation of k_0 with [urea] at two different concentrations of SDS and with DMU at 0.01 mol dm^{-3} of SDS. Following the same analytical procedure as that described for the reaction in water, and starting from the data shown in Fig. 7, we obtain the values of $K_{UH}(ap)$ reported in Table 4. As one can see, $K_{UH}(ap)$ hardly changes with different [SDS]. This consistency in the values is a consequence of the experimental conditions and the nature of the cations binding to the micellar surface; in this sense, $K_{UH}(ap)$ decreases in the presence of SDS micelles (see Table 2 for K_{UH} values); that is, ureas are more basic in the SDS micellar solutions than in pure water. Following Sepúlveda and co-workers,[36] if the principle of hard and soft acids and bases can be applied in this case, we must expect a higher association of UH^+ to the SDS micellar surface than of the H^+ or Na^+ ions.

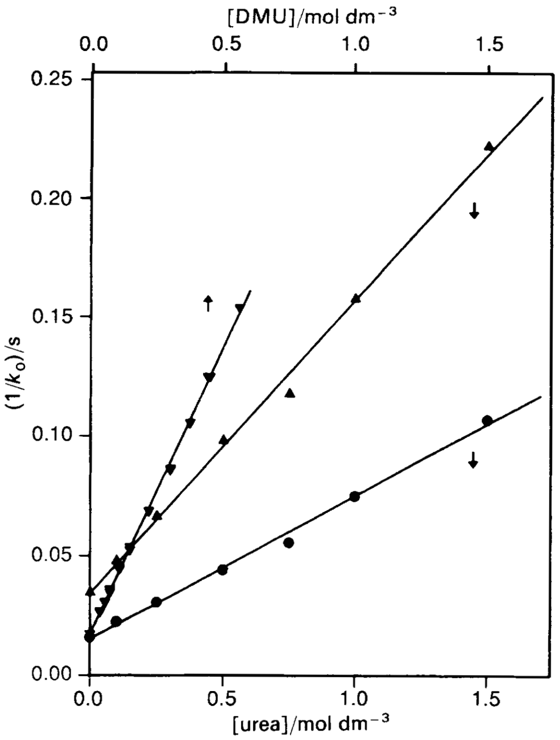


Fig. 7 Reciprocal plot for the influence of [urea] at (▲) [SDS] = 0.12 and (●) [SDS] = 0.012 mol dm^{-3} and of [DMU] at (▼) [SDS] = 0.010 mol dm^{-3} upon the pseudo-first-order rate constant for the acid hydrolysis of PEN

Table 4 Experimental conditions and apparent acid ionization constants of ureas $[K_{UH}(ap)]$ obtained in the presence of SDS

urea	[SDS] /mol dm^{-3}	$K_{UH}(ap)$ /mol dm^{-3}	$(k_2^m/\bar{V})$ /s^{-1}
urea (variable)	0.012	0.27	—
urea (variable)	0.12	0.28	—
urea	—	0.27[a]	317
MEU	—	0.16[a]	319
ETU	—	0.064[a]	329
DMU (variable)	0.0046	0.18	—
DMU (variable)	0.0066	0.098	—
DMU (variable)	0.010	0.082	—
DMU (variable)	0.10	0.091	—
DMU	—	0.095[a]	350
TMU (variable)	0.0053	0.090	—
TMU (variable)	0.13	0.077	—
TMU	—	0.091[a]	321

[a] Values of $K_{UH}(ap)$ obtained from the plots of $1/k_m$ *vs.* [urea].

The experimental kinetic data were fitted to eqn. (6) by the simulation procedure described elsewhere,[35a,b,37] which we have followed in other similar studies applying the ion-exchange model. Constant values of β (0.75–0.40 depending on concentration and urea (type), c.m.c. (kinetic c.m.c. values vary between 1.5×10^{-3} and 4×10^{-3} mol dm^{-3} in the experimental conditions) and k_w $(=k_2^w[H^+]_t K_{UH}/\{K_{UH} + [U]_t\})$ (the experimental first-order rate constant obtained in water in the presence of urea) were used to estimate the values of K_s^N, K_{Na}^H and k_m $(=\{k_2^m/\bar{V}\}K_{UH}(ap)/\{K_{UH}(ap) + [U]_t\})$. We found that values of K_{Na}^H in the range of 0.7–0.8 all gave very good fits to the experimental points, yielding quite similar values of the optimized parameters. We therefore chose $K_{Na}^H = 0.75$ in every case. This is, moreover, the value we found in other studies of acid-catalysed reactions in SDS micelles. The values of the fixed parameters used in the fits, as well as the optimized values of K_s^N and k_m are listed in Table 5. Also given is the root-mean-square deviation of the experimental points from the model. Fig. 5 and 6 (solid lines) show the theoretical fits to the experimental points.

The results obtained for K_s^N in the presence of urea, MEU and ETU barely differ from the value determined in water. This could indicate that, at least for the experimental conditions of this study, the effects of these ureas on the structure of water and of the SDS micelles are not appreciable. Dimethylurea and tetramethylurea give rise to a large increase in the solubility of PEN in water, and thus reduce the value of K_s^N more than two-fold; thus, the water-

Table 5 Values of rate and equilibrium constants used and optimized on fitting experimental data to eqn. (6) by a simulation procedure for the acid hydrolysis of PEN in the presence of SDS micelles and ureas[a]

[urea] /mol dm^{-3}	k_w/s^{-1}	k_m/s^{-1}	K_s^N/dm^3 mol^{-1}	$10^2 \times$ RMSD[b]
		urea		
0[c]	8.2	320	79	3.0
0.10	7.0	229	81	3.3
0.25	5.9	161	86	3.5
0.50	4.9	110	94	3.9
1.0	3.4	67	82	2.4
		methylurea		
0.10	7.7	200	83	2.6
0.25	6.9	132	78.5	2.3
0.5	6.0	82	73	2.4
		ethylurea		
0.10	6.9	134	73	3.8
0.25	6.1	70	74	4.5
0.5	4.5	39	75	4.5
		N,*N*-dimethylurea		
0.11	7.2	175	48	3.5
0.23	5.6	97	52	4.0
		tetramethylurea		
0.10	7.4	159	38	4.8
0.25	5.7	83	35	4.5
0.50	4.6	45	30	4.7

[a] For values of K_{Na}^{H}, β and c.m.c., see text. [b] RMSD, root mean square deviation. [c] Ref. 20.

structure-breaking or hydrophobic-groups-solvating effects of these two ureas are more pronounced than found with urea. Also, the solubilization of DMU and TMU at the micellar surface must be significant, because the degree of micellar charge, α, increases in the presence of both ureas.

Additionally, one can see that the presence of ureas diminishes the observed rate constant obtained both in water, k_w, and at the micellar interface, k_m. This effect is a consequence of the basic nature of the ureas, which are readily protonated and thus reduce the effective concentration of the reactant, H^+. Consequently, in agreement with the expression of k_m (see text), a plot of $1/k_m$ against [urea] (where the k_m values are optimized at various urea concentrations) gives a straight line. The quotient derived from the intercept and the slope of these straight lines is K_{UH}(ap); these constants appear in Table 4, along with the values obtained at fixed [SDS] and with varying [urea]. Meanwhile, from the value of the intercept and $\bar{V} = 0.14$ dm^3 mol^{-1}, one can determine the bimolecular reaction constants at the interface as $k_2^m \approx 45$ dm^3 mol^{-1} s^{-1}, that is, the same value as obtained in the absence of ureas. This result indicates that the ureas do not modify the polarity of the micellar interface, at least not under the experimental conditions of this study. The presence of dimethyl- and tetramethyl-ureas gives rise to a large increase in the solubility of PEN in water and a strong decrease in the value of the rate constant obtained in the micellar phase. These effects could indicate that TMU is dissolved in the micellar surface, thus separating the head groups of the surfactant in the micelle, which should increase the degree of ionization of the micelle and the association of protonated DMU or TMU to the micellar interface.

The fact that alkylureas in the presence of SDS show a greater inhibitory effect than urea can be attributed to the weaker hydrophobic nature of urea compared with the alkylureas. Subsequently, on increasing the hydrophobicity of the protonated urea, this latter species associates more at the micellar surface, thus reducing in greater proportion the effective $[H^+]$. The hydrophobic nature of the ureas is reflected in the value determined for K_{UH}(ap). As shown in Table 4, K_{UH}(ap) decreases on increasing the hydrophobicity of the urea substituent or on increasing the degree of methylation of the urea molecule. Such an increase in the basic nature of a substrate in the presence of SDS micelles has been encountered and quantitatively explained in the case of amines.[38]

Finally, from the activation parameters determined above, it may be concluded that the urea does not affect the solvation state of the reactants or the transition state, since $\Delta H^{\ddagger}$ for the reaction in the micelle scarcely varies in the presence of urea. However, the fact that $\Delta S^{\ddagger}$ is slightly greater in the presence of urea could indicate a change in the solvation state of the reactants in the micellar interface. Evidently, as a consequence of the more restricted movement of the reactants in the interface, the entropy of activation in this medium decreases as compared with the value obtained for the reaction in water; this entropic effect, combined with the lower polarity of the interface, is the main cause for the lower reactivity in the micellar pseudo-phase compared with the reactivity in water.

Conclusions

The results of this study demonstrate that the presence of ureas reduces the pseudo-first-order rate constants obtained in the acid hydrolysis of PEN both in the aqueous and in the micellar pseudo-phases. This inhibition is quantitatively interpreted through the action of ureas as weak bases. This basic character, moreover, increases in the presence of SDS micelles, as we have found with amines. DMU and TMU increase the solubility in water of PEN, a very hydrophobic species. This result can be explained by the two mechanisms of urea action. Both ureas also increase the degree of ionization of the SDS micelles. This means that these ureas are dissolved at the micellar interface, thus decreasing the electrostatic repulsions of the sulfate head groups. Nevertheless, our results indicate that the solubilization process of ureas at the micellar surface does not change the polarity of the interface, since k_2^m has the same value as that found in the absence of urea. These results also predict that the alkyl-substituted ureas are more efficient denaturants than the parent compound.

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