

Kinetics & Mechanism of Hydrolysis of N-Phenylbenzohydroxamic Acid in Sulphuric Acid & Perchloric Acid

KALLOL K GHOSH & SHIV G TANDON*

Department of Chemistry, Ravishankar University, Raipur 492 010

Received 9 April 1984; revised and accepted 21 August 1984

The kinetics of the acid-catalysed hydrolysis of N-phenylbenzohydroxamic acid [$C_6H_5 \cdot N(OH) \cdot C(=O) \cdot C_6H_5$] have been investigated in sulphuric and perchloric acids over a wide range of acidity. The rate-acidity profile is characterised by three distinct features: (i) an initial increase in rate, with increase in [acid], passing through a maximum; (ii) decrease in rate with a minimum; and (iii) increase in rate at high [acid]. The kinetic data have been analysed by means of solvent isotope effect, salt effect, hydration parameter (r) treatment and linear free energy relation. These different rate-acidity dependences have been treated uniformly as a function of H_A and the water activity to yield reaction parameters which are explicable in terms of a change from a bimolecular mechanism (A-2) to a unimolecular mechanism (A-1). The activation parameters, particularly $\Delta S^\ddagger$ are fully consistent with this mechanism.

A few reports are available¹⁻⁴ on the kinetics and mechanism of hydrolysis of N-substituted hydroxamic acids at low (5 mol dm^{-3}) acidity and the evidence to arrive at a plausible mechanism is inadequate. We therefore thought it of interest to investigate the rate of hydrolysis of N-phenylbenzohydroxamic acid (PBHA) [$C_6H_5 \cdot N(OH) \cdot C(=O) \cdot C_6H_5$] over a wide range of concentrated aqueous sulphuric (0.72 to 14 mol dm^{-3}) and perchloric acids (1.16 to 10 mol dm^{-3}) in dioxane-water medium (10% , v/v) in the hope of obtaining detailed information about the various possible mechanisms and the conditions under which they occur.

Materials and Methods

N-Phenylbenzohydroxamic acid was prepared according to literature procedure and characterized by m.p., elemental analyses, UV, IR, PMR and mass spectral data. Sulphuric acid (BDH, AR) and perchloric acid (E. Merck, GR) were used and their solutions were standardized by titration against standard alkali. Dioxane (BDH, AR) was used after purification. Deuterium oxide, D_2O (isotopic purity 99.8%) and D_2SO_4 in D_2O (isotopic purity $>95\%$) were obtained from BARC, Bombay. Ferric-chloride solution used in colorimetric procedure was prepared by the standard method.

The hydrolysis rates were determined from the rate of disappearance of hydroxamic acid in aliquots withdrawn at regular time intervals colorimetrically using a SPEKOL spectrophotometer (Carl-Zeiss Jena) at 530 nm . The colour reaction of PBHA with Fe^{3+} ions giving a purple complex at $pH 1.0 \pm 0.2$ was used for its colorimetric determination. Initial [PBHA] in all kinetic runs was around $6.5 \times 10^{-3} \text{ mol dm}^{-3}$.

Results and Discussion

The observed first order rate constants in sulphuric and perchloric acids are given in Table 1. The rates in these acids follow the order $H_2SO_4 > HClO_4$. It has earlier been shown that the rates of hydrolysis of carboxylate^{5a} in these acids (carboxylates are hydrolysed by an A-2 mechanism) follow the order $H_2SO_4 > HClO_4$, whereas the rates of hydrolysis of esters in these media (esters are hydrolysed by an A-1 mechanism) follow the order $HClO_4 > H_2SO_4$. Thus Bunton *et al.*^{5a} have suggested that the order of acids

Table 1—Observed Pseudo-first Order Rate Constants for Hydrolysis of N-Phenylbenzohydroxamic Acid (PBHA)

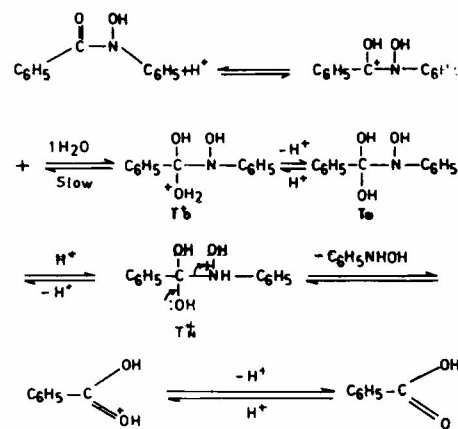
[Temp = 55°C ; solvent: Aq. dioxane (10% , v/v)]					
[Acid] (mol dm^{-3})	$k\psi \cdot 10^3$ (min^{-1})	[Acid] (mol dm^{-3})	$k\psi \cdot 10^3$ (min^{-1})		
	H_2SO_4	$HClO_4$	H_2SO_4	$HClO_4$	
0.72	1.12	—	5.00	8.40	5.27
1.16	—	2.03	5.50	8.50	5.10
1.45	2.40	2.54	5.80	—	4.92
1.45 ^a	3.85	—	6.00	8.70	—
1.75	3.01	—	6.22	—	4.46
2.02	3.43	—	6.50	7.67	4.00
2.20	—	3.47	7.00	6.50	3.65
2.50	4.47	—	7.50	5.60	2.81
2.90	5.30	4.23	7.50 ^a	5.34	—
2.90 ^a	8.06	—	8.00	5.25	3.13
3.25	6.03	—	8.50	4.70	3.55
3.50	6.52	4.70	9.00	4.36	3.93
3.75	6.90	—	10.00	8.10	6.68
4.06	—	5.09	10.50	11.70	—
4.11	7.62	5.13	11.00	18.61	—
4.11 ^a	10.93	—	14.00	464.2	—
4.50	8.14	5.20	—	—	—

(a) D_2SO_4 in D_2O .

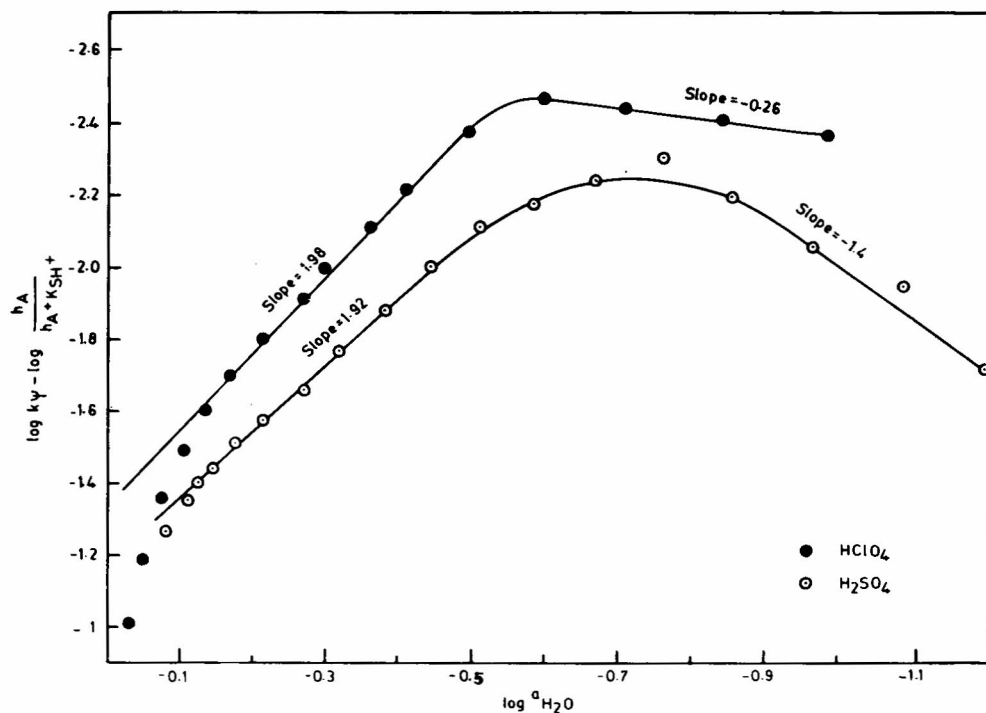
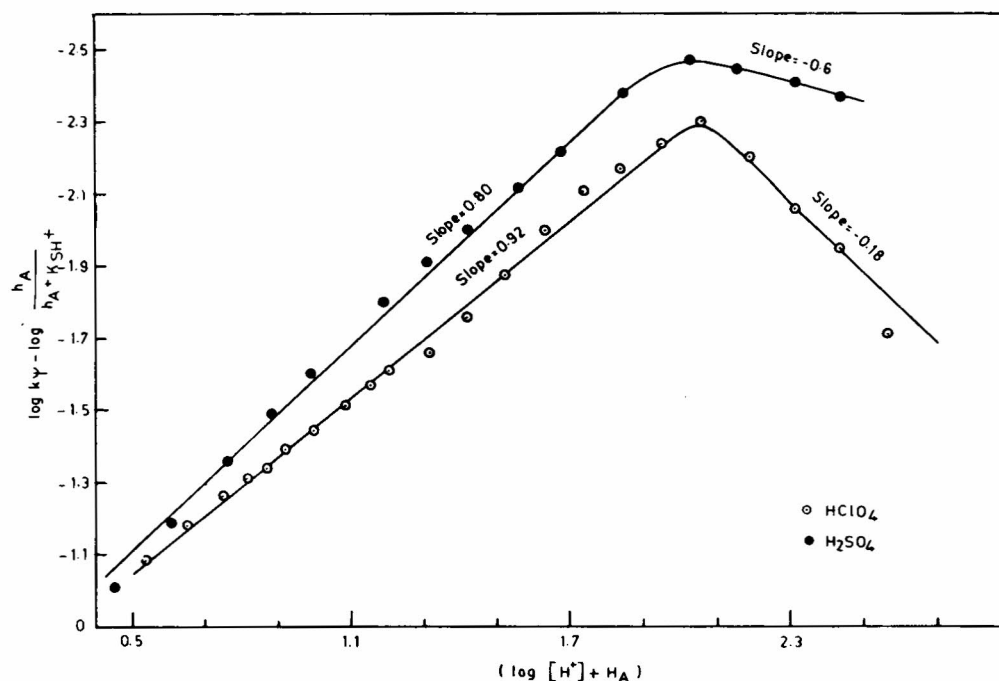
can serve as diagnostic of mechanism. A similar order in the rates of hydrolysis has been observed in the acid-catalysed hydrolysis of benzamide^{5b} and carboxylic anhydrides^{5c}, all of which undergo hydrolysis by a bimolecular attack of water in the transition state. Examination of these data reveals that in the case of sulphuric acid, the rate increases with [acid] up to about 5.5 mol dm⁻³, then decreases up to [acid] around 9.0 mol dm⁻³. In the case of perchloric acid, the same trend is observed with a maximum at 5.0 mol dm⁻³ and minimum at 7.5 mol dm⁻³ of [acid]. These maxima are a consequence of the opposing effects of increasing protonation and decreasing water activity. While studying the acid-catalysed hydrolysis of hexanohydroxamic acid over a wide range of mineral acids Mane and Jagdale⁶ have concluded that protonation is complete at the maximum. This does not appear to be true since the position of the maximum does not necessarily correspond to the point of complete protonation. For example, the PBHA is half-protonated in sulphuric acid of H_A -1.22, i.e. 2.9 mol dm⁻³ (9.6%), and does not get fully protonated in sulphuric acid, since the point of 99% protonation corresponds to $H_A > -3.62$, whereas the maximum in sulphuric acid occurs at about 5-6 mol dm⁻³ acid. In perchloric acid half-protonation occurs at ≈ 2.2 mol dm⁻³ acid and 99% protonation at > 10 mol dm⁻³; the maximum in the rate-acidity profile occurs at 5.0 mol dm⁻³. Similar rate maxima in strongly acidic media have been observed in other acid-catalysed hydrolysis reactions, including those of amides, esters, anilides, nitriles, thioacids and other carboxylic acid derivatives. In such cases, the rate maxima have been attributed to two factors: (i) moderately basic substrates approach complete protonation at higher acidities, and any additional increase in acid concentration (or activity) does not proportionately increase further the degree of protonation of the substrates; (ii) the activity coefficients including that of water itself change. The activity of water decreases with increasing acid activity in concentrated solutions of mineral acids; therefore, the rate of reaction tends to decrease if more water molecules—in whatever capacity either as a solvent or as a nucleophile—are involved in the transition state relative to the protonated substrate.

Recently Edward and colleagues studied sulphuric acid catalysed hydrolysis of ethyl benzoate⁷ and 4-nitroacetanilide⁸. By analogy it is now suggested that hydrolysis of PBHA occurs by two distinct mechanism; (i) in acid concentration up to 9.0 mol dm⁻³ H₂SO₄ and 7.5 mol dm⁻³ HClO₄ by an A-2 mechanism, in which one or more water molecules of water add on to the protonated PBHA in the rate-determining step; and (ii) in acid concentrations above

9.0 mol dm⁻³ H₂SO₄ and 7.5 mol dm⁻³ HClO₄ by an A-1 mechanism in which conjugate acid of PBHA undergoes unimolecular fission without participation of water molecules. According to Lobo *et al.*⁹, who studied the protonation of N-methylhydroxamic acids above 60% H₂SO₄ the activity of water decreases very sharply and this might render the N-protonated cation more unstable than the O-protonated one. Also the O-protonated hydroxamic acid involved in mechanism is the only detectable protonated species present at any acidity, as recently confirmed by Yates¹⁰ in the hydrolysis of amides. In Scheme 1 the rate-determining step is attack of a water molecule on the O-protonated PBHA to give tetrahedral intermediate T₀⁺, deprotonation of which gives the neutral species T₀⁰; further proton transfer results in the intermediate T_N⁺.



In acid media, strong enough to support the formation of protonated ester, T_N⁺ can lose C₆H₅NHOH directly. The Yates-McClelland hydration treatment and Bunnett-Olsen LFER treatment of the kinetic data support these mechanistic interpretations. In 1.16 to 7.5 mol dm⁻³ H₂SO₄ and HClO₄ the points in the *r* plot of Yates-McClelland hydration treatment (Fig. 1) lie close to a straight line with slopes: $r = 1.92$ (in H₂SO₄) and $r = 1.98$ (in HClO₄). These values are close to the value ~ 2 found by Yates and McClelland for an A_{AC} hydrolysis of a variety of acetate esters. These plots do show some curvature in HClO₄ and scattering in H₂SO₄ especially when the points at lower acidities are excluded in calculation of *r*-value. Regions where $r \sim 2$ corresponds to hydrolysis by an A_{AC} type of mechanism in which the slow step of the reaction involves the concerted attack of two molecules of nucleophile water upon the protonated PBHA species, leading to the transition complex and subsequent hydrolysis resulting in the fission of N-acyl bond giving benzoic acid and phenylhydroxylamine as products. Regions where *r* becomes negative correspond to change from the A-2 to the A-1

Fig. 1—Yates-McClelland hydration parameter (r) plots for hydrolysis of PBHAFig. 2—Bunnett-Olsen LFER ϕ plots for hydrolysis of PBHA

mechanism, in which PBHA conjugate acid undergoes unimolecular fission. Water is essentially not involved in the rate-determining step, hence the water activity is effectively zero. In HClO_4 $r = -0.26$ in the most concentrated regions. The two linear regions and the sharp break are consistent with a change in mechanism

from A_{AC2} to A_{AC1} . In H_2SO_4 above 9.0 mol dm^{-3} acid the slope r becomes -1.4 , in the range expected for an A_{A1} mechanism. Such anomalies are difficult to explain. Truly $A-1$ process would be completely independent of $\log a_{\text{H}_2\text{O}}$ since formally water is neither consumed nor released in the rate-determining step.

Bunnett-Olsen plots shown in Fig. 2 are linear (upto $9.0 \text{ mol dm}^{-3} \text{ H}_2\text{SO}_4$ and $7 \text{ mol dm}^{-3} \text{ HClO}_4$) with slopes $\phi = 0.80$ and 0.92 in H_2SO_4 and HClO_4 respectively. Such slope values are large but not unprecedented for an A-2 hydrolysis. Again, the indication of change in mechanism when [acid] exceeds 7.5 mol dm^{-3} (HClO_4) and 9.0 mol dm^{-3} (H_2SO_4) is evident from the fact that ϕ changes sign (-0.18 for HClO_4 and -0.60 H_2SO_4). The value of ϕ for A-2 hydrolysis of amides is generally about $+0.6$ while for A-1 hydrolysis it is lower and may be negative. Bunton *et al.*^{5a} have shown that the effect of added perchlorates (4 mol dm^{-3}) is negligible or slightly negative in A-2 acid-catalysed ester hydrolysis. Added perchlorate (NaClO_4) (0.94 mol dm^{-3} to 2.8 mol dm^{-3}) has either negligible or minor retarding effect. For example in $2.2 \text{ mol dm}^{-3} \text{ HClO}_4$ $k_{\psi} \times 10^3 \text{ min}^{-1}$ values are 3.47 and 3.45 in the absence and presence of 2.8 mol dm^{-3} of added NaClO_4 whereas in $2.9 \text{ mol dm}^{-3} \text{ HClO}_4$ the values are 4.23 and 4.24 respectively in the absence and presence of $2.1 \text{ mol dm}^{-3} \text{ NaClO}_4$.

The solvent isotope effects ($k_{\text{D}_2\text{O}}/k_{\text{H}_2\text{O}}$) are found to be 1.60 and 0.95 in 1.45 and $7.5 \text{ mol dm}^{-3} \text{ H}_2\text{SO}_4$ respectively. Deuterium oxide is a weaker base than water and therefore nucleophilic attack by D_2O in A-2 mechanism will be less effective than H_2O . However, the substrate will be able to compete with the solvent for a deuteron in D_2O more effectively than for a proton in H_2O and thus the concentration of the protonated species will be greater in D_2O than in H_2O . Consequently, the rate should be faster in the former. As in the case of amides, the effect will be to shift the maximum in the rate-acid profile to lower acidities. The above solvent isotope effects are consistent with a reaction involving a rapid pre-equilibrium followed by bimolecular attack by deuterium oxide or water. When protonation is substantially complete, $k_{\text{D}_2\text{O}}/k_{\text{H}_2\text{O}}$ is around 1 ; when the pre-equilibrium is significant (at lower acidities) $k_{\text{D}_2\text{O}}/k_{\text{H}_2\text{O}} > 1$. Thus, the magnitude of the solvent isotope effect in a given acid strength reflects the relative importance of the protonation.

Another line of evidence indicating a change from an A-2 to an A-1 mechanism comes from measurements of the entropy of activation ($\Delta S^\ddagger$). Acid-catalysed hydrolyses of esters and amides proceeding by an A-1 mechanism have $\Delta S^\ddagger$ of about 0 to $41.84 \text{ J mol}^{-1} \text{ K}^{-1}$, while those proceeding by A-2 mechanism have $\Delta S^\ddagger$ of -62.8 to $-125.5 \text{ J mol}^{-1} \text{ K}^{-1}$ (ref. 11). The values of activation parameters are listed in Table 2. In the region of concentrated sulphuric acid where the

Table 2—Temperature Dependence of Rate of Acid-Catalysed Hydrolysis of PBHA

[Acid] (mol dm^{-3})	Temp ($^\circ\text{C}$)	$10^3 \times k_{\psi}$ (min^{-1})	$\log A$ s^{-1}	$\Delta S^\ddagger$ ($\text{J mol}^{-1} \text{ K}^{-1}$)	E_a (kJ mol^{-1})
2.9	45	2.10			
	55	5.30	8.67	-79.6	80.0
	65	12.58			
5.0	45	3.25			
	55	8.40	9.18	-69.8	81.9
	65	20.33			
7.0	55	6.50			
			10.15	-51.3	88.7
	65	17.02			
10.0	45	2.70			
	55	8.10	11.72	-21.3	97.9
	65	24.20			
11.0	45	5.60			
	55	18.61	13.35	+9.7	105.9
	65	59.80			
14.0	35	32.57			
	45	124.6	15.68	+54.7	111.8
	55	464.2			

mechanism is presumed to have changed from A-2 to A-1, E_a (energy of activation) has increased significantly and $\Delta S^\ddagger$ has changed from negative value, (typical of A-2 reactions) to a small positive value (more typical of A-1 process).

Acknowledgement

One of the authors (KKG) is grateful to the CSIR, New Delhi for the award of a junior research fellowship. The authors are also indebted to Prof. C H Rochester, University of Dundee.

References

- Subbarao C & Tandon S G, *Indian J Chem*, **14A** (1976) 765.
- Berndt D C & Ward I E, *J org Chem*, **41** (1976) 3297.
- Mollin J & Kucerova T, *Colln Czech chem Commun*, **41** (1976) 2885.
- Tandon S G & Ghosh K K, *Bull chem Soc Belg*, **91** (1982) 5.
- (a) Bunton C A, Crabtree J H & Robinson L, *J Am chem Soc*, **90** (1968) 1258.
(b) Bunton C A, Farber S J, Miubank A J G O, Connor C J & Turney T A, *J chem Soc Perkin II*, (1979) 1869.
(c) Bunton C A & Findler J H, *J org Chem*, **30** (1965) 1365.
- Mane B S & Jagdale M H, *Indian J Chem*, **15A** (1977) 1086.
- Edward J T & Wong S C, *J Am chem Soc*, **99** (1977) 7224.
- Edward J T, Derald G D & Wong S C, *J Am chem Soc*, **102** (1978) 1023.
- Lobo A M, Prabhakar S & Fonseca M T C, *Tetrahedron Lett*, **36** (1977) 3167.
- Cox R A & Yates K, *Can J Chem*, **59** (1980) 1560.
- Yates K, *Acc chem Res*, **4** (1971) 136.