

# The Reduction Mechanism of the >C=O Group. Part 4. The Electrochemical Reduction of Isonicotinamide in an Aqueous Medium

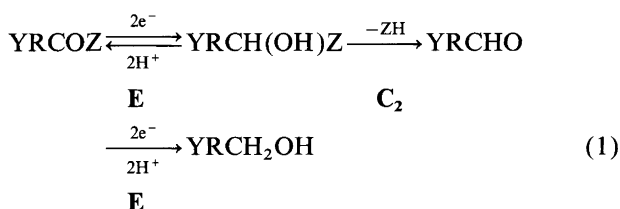
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The electrochemical reduction of isonicotinamide NRCONH<sub>2</sub> (INA), in which NR represents the pyridine ring, has been examined in an aqueous medium between  $H_0 = -3$  and pH 13.75. For pH < 6, as shown by cyclic voltammetry and polarography, a global  $2e^-$ ,  $2H^+$  reversible transfer is followed by two first-order chemical reactions and a  $2e^-$ ,  $2H^+$  reduction (EC<sub>1</sub>C<sub>2</sub>E process), as in the case of methyl isonicotinate NRCOOMe. The first chemical reaction C<sub>1</sub> is an internal proton transfer (rate constant  $\kappa_1$ ). The second chemical reaction C<sub>2</sub> (rate constant  $\kappa_2$ ), which is much slower than C<sub>1</sub>, involves the loss of NH<sub>3</sub> to give 4-pyridinecarbaldehyde, which is more easily reduced than INA. Reaction C<sub>2</sub> is base-catalyzed but no acid catalysis was found even in very acidic media ( $H_0 = -3$ ). At pH > 6, another first-order reaction C<sub>1</sub>' (rate constant  $\kappa_1'$ ) competes with C<sub>1</sub> to give 1,4-dihydropyridine-4-carboxamide which is not further reducible. The rate constant  $\kappa_1$  is independent of pH between pH 1.8 and 5. Above pH 6,  $\kappa_1'$  is independent of pH, whereas log  $\kappa_1$  decreases linearly with a slope of -2; this result can be interpreted as a consequence of adsorption, the actual rate constant  $\kappa_{s1}$  of the reaction, which occurs on the electrode surface, being inversely proportional to the concentration of the H<sup>+</sup> ions.

It was shown previously by Lund<sup>1</sup> that below pH 6–7, a global  $2e^-$  electrochemical reduction of the C=O bond of isonicotinamide (INA) leads to the *gem*-hydroxylamine GHA, the analogue of the hydrate of 4-pyridinecarbaldehyde, which can lose an NH<sub>3</sub> molecule fairly rapidly, thus giving the aldehyde which is immediately reduced [EC<sub>2</sub>E reaction, eqn. (1) in which YR represents the pyridine ring, protonated or not, and Z = NH<sub>2</sub>].

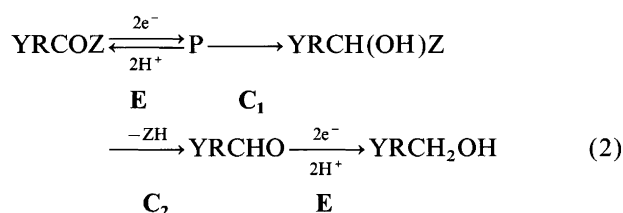


When the solution becomes more basic, a  $2e^-$  reduction of YRCONH<sub>2</sub> yields a dihydropyridine DH.<sup>1</sup> This global reaction scheme was used in subsequent studies<sup>2,3</sup> of INA.

A detailed analysis of the reduction of the analogue

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of INA, methyl isonicotinate<sup>4</sup> [eqn. (1) with Z = OMe] led to the conclusion that the reaction actually proceeds via a more complex sequence, namely EC<sub>1</sub>C<sub>2</sub>E; C<sub>1</sub> is also first order and much faster than C<sub>2</sub> [eqn. (2)].



As shown in the present paper, the same type of intermediate P is also formed during the reduction of INA.

## Experimental

INA was a commercial product.

The experiments were carried out in aqueous solutions at a temperature of 20 °C. For pH > 1.80, Britton–Robinson buffers were used with 1 mol dm<sup>-3</sup> NaCl as the supporting electrolyte except for solutions where the influence of the buffer was studied; in this case, a

KCl-HCl or KCl-KHCO<sub>3</sub> buffer was also used.<sup>5</sup> All the pH values were measured after the experiments. For pH < 1.80, H<sub>2</sub>SO<sub>4</sub> was employed and the H<sub>0</sub> values were taken from literature data.<sup>6</sup> The potentials are referred to the saturated calomel electrode (SCE). For pH < 1, the error due to the junction potential and the effect of the viscosity were corrected as described previously.<sup>7</sup>

The equipment used in earlier work<sup>7</sup> was employed to obtain the voltammograms and the polarograms. During the polarography, the drop time  $\tau$  was 0.5 s and the rate of capillary flow  $m$  was 1.08 mg s<sup>-1</sup>.

**Results**

We studied INA by polarography and cyclic voltammetry (CV) between H<sub>0</sub> = -3 and pH = 13.75.

*Polarography.* A single wave was observed over the whole pH range. The variations of the half-wave potentials E<sub>1/2</sub> are shown in Fig. 1. The limiting current of the single wave varied with pH as shown in Fig. 2. These variations agree with those given by Lund.<sup>1</sup> We have also indicated in the figure the variations of the limiting current of 4-pyridinecarbaldehyde.<sup>8</sup>

*Voltammetry at a carbon rotating disk electrode.* In a strongly basic medium (NaOH 0.5 mol dm<sup>-3</sup>), INA gave one reduction wave at about the same potential as observed by polarography.

*Cyclic voltammetry at a mercury electrode.* As in the case of methyl isonicotinate, the cyclic voltammograms showed that the global electrochemical reaction was reversible (a reoxidation peak was observed) when the sweep rate was large enough. Examples of variations of the cathodic and anodic peak potentials E<sub>pc</sub> and E<sub>pa</sub> are

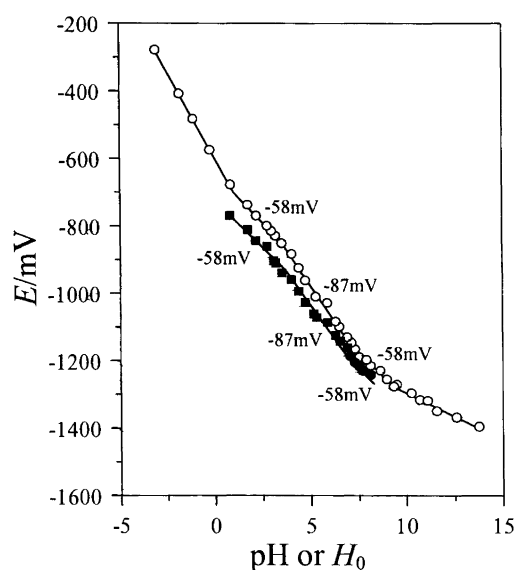


Fig. 1. Variations of the potentials with pH or H<sub>0</sub>: (○) E<sub>1/2</sub>; (■) E°; (●) E°; C = 10<sup>-4</sup> mol dm<sup>-3</sup>.

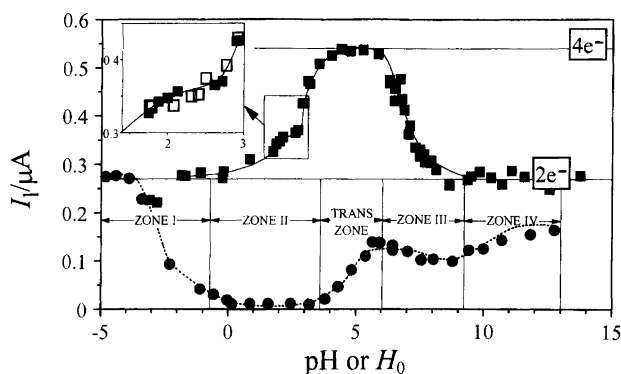


Fig. 2. Variations of the limiting current with pH or H<sub>0</sub> for INA (■) in BR buffers; (□) in KCl-HCl or KCl-KHCO<sub>3</sub> buffers and for 4-pyridinecarbaldehyde (●); C = 10<sup>-4</sup> mol dm<sup>-3</sup>.

shown in Figs. 3 and 4. Below pH 6, the shape of the peaks indicated a diffusion-controlled process<sup>9a</sup> when log  $\nu$  was smaller than 1.5-2 (Fig. 3); for larger values, the electrochemical reaction was of a surface nature.<sup>10</sup> Above pH 6, the peaks had a 'diffusional' shape whatever the sweep rate (Fig. 4).

The changes of the slopes are analyzed in the Discussion, however it must be noted that the equilibrium potential for the surface<sup>10</sup> (E°) or heterogenous<sup>9b</sup> (E°)

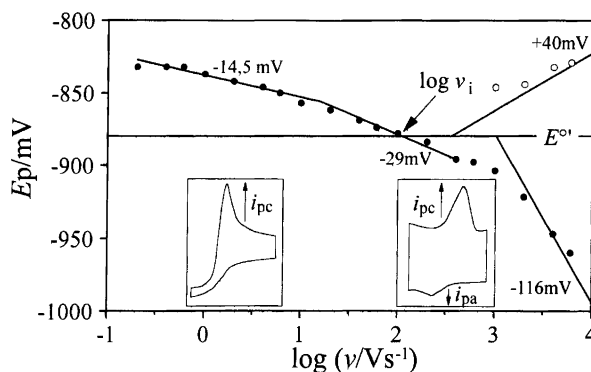


Fig. 3. Variations of E<sub>pc</sub> (●) and E<sub>pa</sub> (○) with log  $\nu$ ; C = 10<sup>-4</sup> mol dm<sup>-3</sup>; pH = 2.7 and examples of voltammograms for  $\nu = 1$  V s<sup>-1</sup> and 2000 V s<sup>-1</sup>.

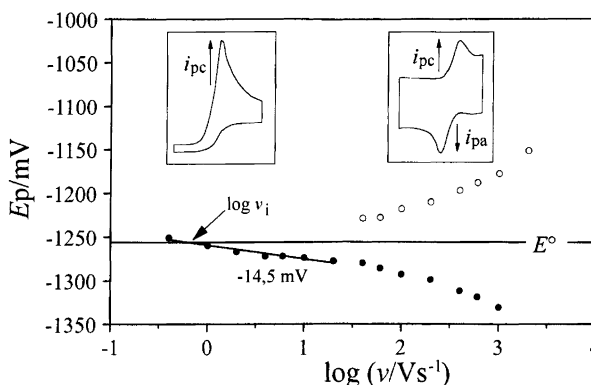


Fig. 4. Variations of E<sub>pc</sub> (●) and E<sub>pa</sub> (○) with log  $\nu$ ; C = 5 × 10<sup>-4</sup> mol dm<sup>-3</sup>; pH = 7.8 and examples of voltammograms for  $\nu = 2$  V s<sup>-1</sup> and 1000 V s<sup>-1</sup>.

reactions can easily be obtained in the region where both the cathodic and anodic peaks are seen; the values we found are reported in Fig. 1.

*Cyclic voltammetry at a carbon electrode.* In a strongly basic medium (NaOH 0.5 mol dm<sup>-3</sup>), the voltammogram showed one reduction peak A at about -1500 mV and one anodic peak B at about +300 mV (Fig. 5).

## Discussion

*The first 2e<sup>-</sup> electrochemical reaction.* The first 2e<sup>-</sup> electrochemical reaction can be described by using a three-dimensional scheme<sup>11</sup> such as that of Fig. 6 in which the molecules protonated on the pyridine nitrogen lie in the bottom plane. If the protonations are fast (at equilibrium), the kinetics of the system can be treated mathematically;<sup>11</sup> the solutions obtained are valid for a heterogeneous reaction or for a surface reaction when a Langmuir isotherm is obeyed. The scheme is kinetically equivalent to two simple successive reactions with apparent rate constants  $ke_1^*$  and  $ke_2^*$  and apparent standard potentials  $E_{r1}$  and  $E_{r2}$  (Fig. 6); when the two stages overlap ( $E_{r1} < E_{r2}$ ), it is thus equivalent to a simple 2e<sup>-</sup> reaction, with an apparent equilibrium potential  $E^\circ = 0.5(E_{r1} + E_{r2})$ .

*The nature of the intermediate P.* If INA were reduced directly to the *gem*-hydroxylamine GHA [eqn. (1)], we would not obtain an anodic peak in CV; GHA, like the hydrate of 4-pyridinecarbaldehyde,<sup>8</sup> is certainly not electrochemically reoxidizable in the potential range available at mercury.

As in the case of isonicotinic acid and its derivatives,<sup>4,12</sup> the 2e<sup>-</sup> reduction of INA involves the intermediates P [eqn. (2)] which are shown in Fig. 6. Formally, as shown in Ref. 4, these intermediates can be written as molecules in which the carbon of the substituent bears a minus charge (sp<sup>3</sup> hybridization) and which can be protonated

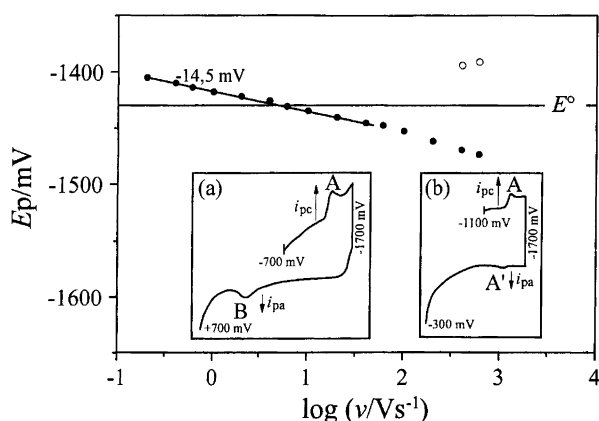


Fig. 5. Variations of  $E_{p,c}$  (●) and  $E_{p,a}$  (○) with  $\log v$  on an Hg electrode;  $C = 10^{-4}$  mol dm<sup>-3</sup>; pH = 13.7 and examples of voltammograms for (a)  $v = 40$  V s<sup>-1</sup> on a carbon electrode and (b)  $v = 400$  V s<sup>-1</sup> on a mercury electrode.

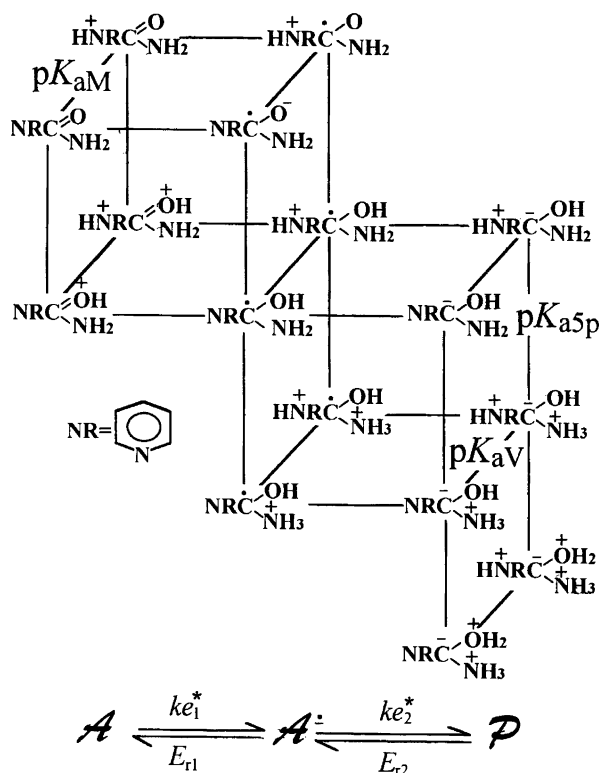


Fig. 6. The three-dimensional scheme.

on different sites (nitrogen of the pyridine ring, nitrogen and/or oxygen of the substituent). In Fig. 6, the protonations on the substituent are represented preferentially on the nitrogen but forms involving protonations on the oxygen can also be written. These diverse forms can actually possess quinoidal structures<sup>13</sup> (with an sp<sup>2</sup> carbon) which may be prevalent.

*The equilibrium diagram.* The variation with increasing pH of the equilibrium potential  $E^\circ$  or  $E^{\circ'}$  for the first 2e<sup>-</sup> reduction (Fig. 1) consists of three linear segments with slopes of -58 mV, -87 mV and -58 mV, indicating the successive involvement of two protons, three protons and two protons. Since we know the initial compounds, we can determine the domains of existence of the products of the reduction (Fig. 7). The reduced forms indicated in the figure represent only a few of the many possible mesomeric structures. The value of 3.6 found for  $pK_{aM}$  (Figs. 6 and 7) is in good agreement with that (3.61) given in the literature.<sup>14</sup>

*The transformation of P.* Generally speaking, the occurrence of a chemical reaction can affect both the half-wave potential and the current. (a) When a chemical reaction follows a reversible electrochemical reduction (EC mechanism), the half-wave potential  $E_{1/2,e}$  or the peak potential  $E_{p,c}$  of the cathodic CV peak is shifted towards positive potentials. If the rate constant  $\kappa$  of the chemical reaction is large enough (purely kinetic case), the shift (for mean current in polarography) is given by

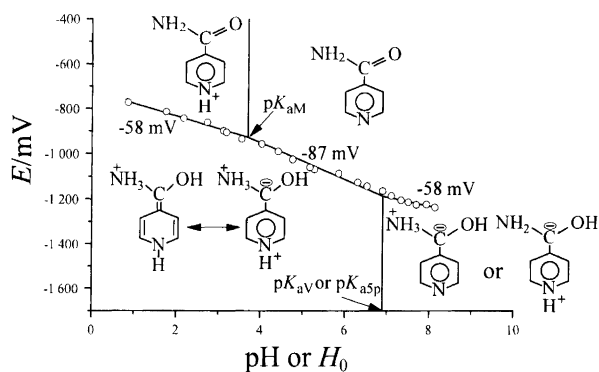


Fig. 7. The equilibrium diagram. Only the quinoidal structure of the product below pH 7 is represented, due to space constraints. For a definition of  $pK_a$ s, see Fig. 6.

eqns. (3) and (4) for a heterogeneous E reaction followed by a volume C reaction<sup>15,16</sup> and eqns. (5) and (6) when both E and C are surface reactions.<sup>17,18</sup>

$$E_{1/2,c} = E^\circ + \frac{2.3RT}{nF} \log 0.886 + \frac{2.3RT}{2nF} \log \kappa\tau \quad (3)$$

$$E_{p,c} = E^\circ - \frac{RT}{nF} 0.78 + \frac{2.3RT}{2nF} \log \frac{RT\kappa}{nFv} \quad (4)$$

$$E_{1/2,c} = E^\circ + \frac{2.3RT}{nF} 0.261 + \frac{2.3RT}{nF} \log \kappa\tau \quad (5)$$

$$E_{p,c} = E^\circ + \frac{2.3RT}{nF} \log \frac{RT\kappa}{nFv} \quad (6)$$

In CV, the anodic peak then disappears. If  $E^\circ$  or  $E^\circ'$  is accessible, the rate constant can be calculated from these equations.

The variations of  $E_{1/2,c}$  or  $E_{p,c}$  are the same for an ECE reaction<sup>19</sup> or, in the present case, for an ECCE reaction, because they are brought about by the chemical reaction immediately following the first reversible electrochemical reaction.

(b) In the case of an ECE or ECCE mechanism, an increase of the total current  $\bar{i}_k$  over that  $\bar{i}_d$  due to the first process is observed. The ratio of the mean currents  $(\bar{i}_k + \bar{i}_d)/\bar{i}_d$  as a function of  $\kappa\tau$  can be found in Ref. 19 for a volume reaction.

In the present case, the slopes of the asymptotes at slow sweep rates do indicate that the first  $2e^-$  reaction is followed by a first order chemical reaction. The slopes are  $-14.5$  or  $-29$  mV/log  $v$  according to whether the electrochemical reaction is of a heterogeneous or surface nature (Figs. 3 and 4). The rate constant  $\kappa_1$  of the chemical reaction can be determined in certain pH zones when the appropriate asymptote can be associated with  $E^\circ$  or  $E^\circ'$  [cf. eqns. (4) and (6)].

Below pH 5, the constant  $\kappa_1$  determined by using eqns. (2)–(6) is always larger (by about two to five orders of magnitude) than the constant  $\kappa_d$  (Fig. 8) measured from the ratio  $(\bar{i}_k + \bar{i}_d)/\bar{i}_d$  according to Ref. 19. If we introduce the values of  $\kappa_1$  in Nicholson and Shain's treatment,<sup>19</sup>

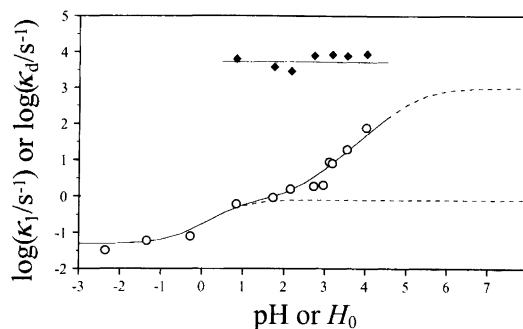


Fig. 8. Variations of the logarithm of the rate constants with pH: ( $\blacklozenge$ )  $\log \kappa_1$  calculated by cyclic voltammetry (the electrochemical and chemical reactions are of a surface nature); ( $\circ$ ) and (—)  $\log \kappa_d$ : experimental points and simulated curve.

we obtain the result that the current should always correspond to  $4e^-$ , so that  $\kappa_1$  is the rate constant of the first chemical reaction. Therefore,  $\kappa_d$  is the rate constant for the deamination of INA; it is indeed of the same order of magnitude as the dehydration rate constant determined previously for the dihydrate  $YRCH(OH)_2$ .<sup>8</sup> As in the case of methyl isonicotinate<sup>4</sup> and isonicotinic acid,<sup>12</sup> it can be assumed that P is transformed into GHA following an internal proton transfer from the protonated site (nitrogen or oxygen of the substituent) to the carbon which bears the minus charge. The relative slowness of this reaction may be connected with the change in hybridization (from  $sp^3$  to  $sp^2$ ) at the carbon of the substituent, analogous to that observed in the protonation of enolate or nitronate ions.

*The deamination reaction for INA.* It should be noted that for INA, an alternative to loss of  $NH_3$  would be dehydration, leading to the imine  $YRCH=NH$  instead of the aldehyde. However, this path can be excluded for INA since it has been shown by Lund that formation of the amine, which would result from the reduction of the imine, does not occur for INA itself, although a certain amount of amine is formed in the case of isonicotinamide in strongly acidic media.<sup>1</sup>

Before discussing the deamination reaction, we recall the behavior of 4-pyridinecarbaldehyde. The limiting current of the  $2e^-$  reduction wave is determined by the equilibrium between the free and hydrated forms and by the dehydration rate.<sup>8</sup> Four zones can be defined (Fig. 2): In zone II, the aldehyde exists mostly in its hydrated form, and the current is controlled by the dehydration rate. In zone III, the dehydration rate is small, so that the part of the current due to the fraction of the free aldehyde existing at equilibrium in solution is large. In zones I and IV, the current increases because the dehydration is acid- and base-catalyzed respectively.

In the case of INA, between pH 0 and 4, the current increases in two successive steps (Fig. 2). Since the current found for the aldehyde in this region is small, the increase must be due to a catalysis of the deamination

reaction. The variations of the current with pH remain identical when the nature of the buffer is changed (Fig. 2) so that the step cannot be due to catalysis by a component of the buffer. Formally, the variations of the current can be interpreted on the basis of the deamination reaction in equilibrium, as represented in Fig. 9. We can thus write eqn. (7),  $C_T$  being the sum of the concentrations of  $\text{GH}_2^{2+}$ ,  $\text{GH}^+$  and  $\text{G}$ .

$$-\frac{dC_T}{dt} = \kappa_{d1} C_{\text{GH}_2^{2+}} + \kappa_{d2} C_{\text{GH}^+} + \kappa_{d3} C_{\text{G}} \quad (7)$$

From the definition of the  $pK_a$ s, we easily obtain eqn. (8).

$$-\frac{dC_T}{dt} = \frac{\kappa_{d1} + \kappa_{d2} K_{a1} [\text{H}^+]^{-1} + \kappa_{d3} K_{a1} K_{a2} [\text{H}^+]^{-2}}{1 + K_{a1} [\text{H}^+]^{-1} + K_{a1} K_{a2} [\text{H}^+]^{-2}} \times C_T \equiv \kappa_d C_T \quad (8)$$

A simulated curve is shown in Fig. 8 together with the experimental points. The best fit is obtained for  $\log \kappa_d$  calculated with  $\kappa_{d1} = 0.08 \text{ s}^{-1}$ ,  $\kappa_{d2} = 1 \text{ s}^{-1}$ ,  $\kappa_{d3} = 900 \text{ s}^{-1}$ ,  $pK_{a1} = 1$  and  $pK_{a2} = 5.2$ . The value of  $\kappa_{d1}$  is fixed between 0.07 and 0.08 by the plateau below  $H_0 = -1$  (Fig. 8). That of  $\kappa_{d2}$ , which is determined by the plateau between pH 1.5 and 2 (Fig. 8), is less precise. Values in the range 0.7–1.3 give satisfactory results; then  $pK_{a1}$  must lie between 0.5 and 1. As for  $\kappa_{d3}$  and  $pK_{a2}$ , several sets of values are possible, for example, 700 and 5, 7000 and 6, 70000 and 7, etc. In any case,  $pK_{a2}$  must be larger than 5 and  $\kappa_{d3}$  is probably not very large, if it is assumed that it has the same order of magnitude as the dehydration rate constants (cf. Refs. 8, 12, 21 and next paragraph).

The formulation of the species  $\text{GH}_2^{2+}$ ,  $\text{GH}^+$  and  $\text{G}$  is not straightforward. At first sight, we could assume that the appearance of this step is linked to a change in the protonation of the pyridine nitrogen. However, this should not be the case, because a similar step is also observed for the *N*-methyl derivative;<sup>20</sup> moreover, this shows that the species must be protonated on the pyridine nitrogen. Our results can be compared to those obtained<sup>12,21</sup> for the dehydration of the hydrate of 1-methylpyridinium-4-carbaldehyde (Fig. 10). The forms involved are shown in Fig. 10(a); for the constants, we have  $\kappa_{d1}' > 2 \times 10^2 \text{ s}^{-1}$ ,  $\kappa_{d2}' = 0.4 \text{ s}^{-1}$ ,  $\kappa_{d3}' = 20 \text{ s}^{-1}$ ,  $pK_{a1}' < -3$  and  $pK_{a2}' = 12$ . Since  $\kappa_{d2}' < \kappa_{d1}'$ , the reaction is acid-catalyzed below  $H_0 = 0$  whereas  $\kappa_{d3}' > \kappa_{d2}'$  leads to base catalysis above pH 9.

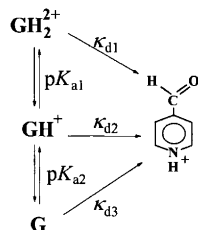


Fig. 9. The deamination reaction.

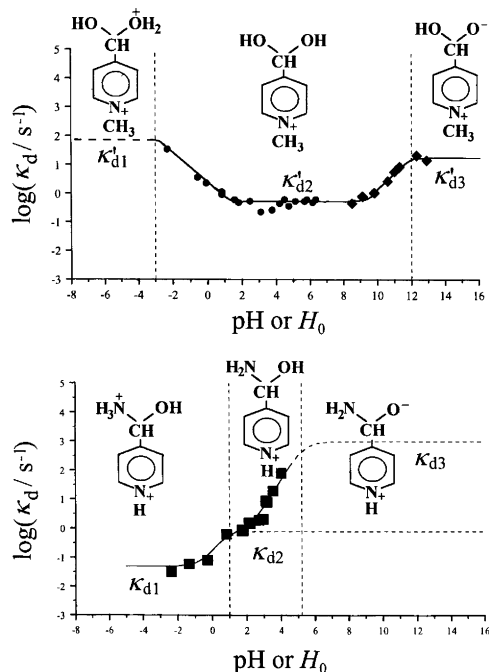


Fig. 10. (a) Variations of  $\log \kappa_d$  with pH for 1-methylpyridinium-4-carbaldehyde ion: (●) taken from Ref. 12, (◆) calculated from Ref. 21. (b) Variations of  $\log \kappa_d$  with pH for INA.

By analogy, we should have the forms indicated in Fig. 10(b) for the *gem*-hydroxylamine. Since  $\kappa_{d1} < \kappa_{d2} < \kappa_{d3}$ , no increase in the current is observed in acidic media. However, it is difficult to explain why  $\kappa_{d1}$  is smaller than  $\kappa_{d2}$ , i.e. why the form  $\text{GH}_2^{2+}$  is more stable than the form  $\text{GH}^+$ ; the value of  $pK_{a2}$  seems reasonable but that for  $pK_{a1}$  ( $\approx 1$ ) is perhaps more difficult to compare with that of  $pK_{a1}' (< -3)$  for the dihydrate.

Another explanation might be that  $pK_{a1}$  is very negative (below about  $H_0 = -7$ ), so that the acid catalysis cannot be observed under our experimental conditions, but this seems unlikely; besides, the step cannot be accounted for.

*The reduction of the pyridine ring above pH 6.* It has been shown by Lund<sup>1</sup> that, for INA and its analogues, a  $2e^-$  reduction of the pyridine ring occurs in alkaline media, yielding the corresponding dihydropyridine, which can be reoxidized to INA or its analogues either by oxygen or electrochemically.

*A priori*, the dihydropyridine DH can be obtained either by direct reduction of INA (path a, Fig. 11) or via a rearrangement of species P (path b, Fig. 11). In order to distinguish between the two possibilities, we carried out a voltammetric study of INA at a mercury electrode at  $\text{pH} = 13.70$  ( $\text{NaOH } 0.5 \text{ mol dm}^{-3}$ ). The slope ( $-14 \text{ mV}$ ) of the peak potential of A (Fig. 5) indicates that the  $2e^-$  reduction of INA is followed by a first-order chemical reaction which should correspond to

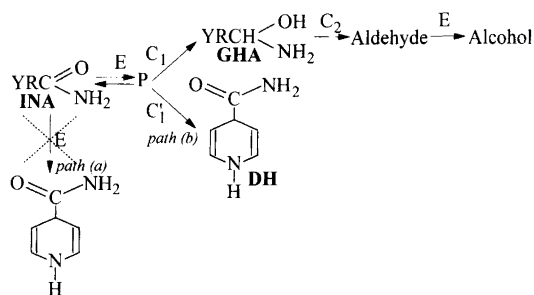


Fig. 11. The reaction path in an alkaline medium.

the rearrangement of P. When the sweep rate is increased, a reoxidation peak A' can indeed be seen at about the same potential as that of peak A [Fig. 5(b)]. No anodic peak corresponding to the reoxidation of the dihydropyridine DH is observed because it is hidden by the anodic discharge current; however, the voltammogram at a carbon electrode indeed shows an oxidation peak at about +300 mV [Fig. 5(a)]. Moreover, in view of the continuity existing in the evolution of the reduction potentials, and since P is formed at high sweep rates, it can be inferred that P is also the intermediate at slower sweep rates in this medium. This shows that the formation of DH occurs through a rearrangement of P; this reaction  $C_1'$  (Fig. 11) thus competes with the rearrangement of P to the *gem*-hydroxylamine GHA (reaction  $C_1$ ).

The competition between reactions  $C_1$  and  $C_1'$  modifies both the height of the wave (it is responsible for the decrease of the current above pH 6) and the variations of the potentials. We now analyze these two effects in turn.

(a) Height of the wave. None of the species remain adsorbed during the time used for the measurement at slow sweep rates (long measurement times) in CV (Fig. 4); this is also, *a fortiori*, true in polarography, since the measurement time ( $\tau$ ) is even larger. Both reactions  $C_1$  and  $C_1'$  being fast, we can thus use the reaction layer concept<sup>22</sup> to analyze the situation.

The flux of INA reaching the electrode is given in eqn. (9), where  $D_{\text{INA}}$  is the diffusion coefficient,  $\delta$  the thickness of the diffusion layer and  $C_T$  the analytical concentration.

$$f_{\text{INA}} = \frac{D_{\text{INA}}}{\delta} C_T \quad (9)$$

Since P disappears completely in the reaction layer, an equivalent flux  $f_P$  can be written for it, eqn. (10), in

$$f_P = \frac{D_P}{\mu_1} C_P^S + \frac{D_P}{\mu_1'} C_P^S = f_{P1} + f_{P1'} \quad (10)$$

which  $D_P$  is the diffusion coefficient of P, and  $C_P^S$  the concentration of P in the reaction layers of thickness  $\mu_1$  (reaction  $C_1$ ) and  $\mu_1'$  (reaction  $C_1'$ ).

The current is given by eqn. (11), in which  $A$  is

$$\bar{i} = 4FAf_{P1} + 2FAf_{P1'} \quad (11)$$

the area of the electrode, since along the upper and lower paths four and two electrons are involved, respectively.

The reaction layers are defined by eqn. (12).<sup>22</sup>

$$\mu_1 = \left( \frac{D_{\text{INA}}}{\kappa_1} \right)^{1/2} \quad \text{and} \quad \mu_1' = \left( \frac{D_{\text{INA}}}{\kappa_1'} \right)^{1/2} \quad (12)$$

By expressing that  $f_{\text{INA}} = f_P$  (it is assumed that  $D_{\text{INA}} = D_P$ ), we thus obtain eqn. (13) for the dimensionless current, in which the current  $\bar{i}_d$  is that for the first  $2e^-$

$$\bar{I} = \frac{\bar{i}}{\bar{i}_d} = \frac{2(\kappa_1/\kappa_1')^{1/2} + 1}{(\kappa_1/\kappa_1')^{1/2} + 1} \quad (13)$$

process, eqn. (14). Equation (13) enables us to determine the ratio  $\kappa_1/\kappa_1'$ .

$$\bar{i}_d = 2FADC_T \delta^{-1} \quad (14)$$

(b) Variations of  $E_{pc}$  and  $E_{1/2}$ . The problem is defined by eqns. (15) and (16),<sup>23</sup> with the boundary conditions;

$$\frac{\partial C_{\text{INA}}}{\partial t} = D_{\text{INA}} \frac{\partial^2 C_{\text{INA}}}{\partial x^2} \quad (15)$$

$$\begin{aligned} \frac{\partial C_P}{\partial t} &= D_P \frac{\partial^2 C_P}{\partial x^2} - \kappa_1 C_P - \kappa_1' C_P \\ &= D_P \frac{\partial^2 C_P}{\partial x^2} - (\kappa_1 + \kappa_1') C_P \end{aligned} \quad (16)$$

at  $x = \infty$ ,  $C_{\text{INA}} = C_T$ ,  $C_P = 0$ ; at  $t = 0$ ,  $C_{\text{INA}} = C_T$ ,  $C_P = 0$  and for  $t > 0$  and  $x = 0$ , the Nernst equation (17), in which the potential  $E$  is constant for polarography and  $E = E_i + \nu t$  for CV, and the equality of the fluxes at the electrode surface, eqn. (18).

$$\frac{C_{\text{INA}}}{C_P} = \exp \left[ \frac{nF(E - E^0)}{RT} \right] \quad (17)$$

$$D_{\text{INA}} \left( \frac{\partial C_{\text{INA}}}{\partial x} \right)_{x=0} = -D_P \left( \frac{\partial C_P}{\partial x} \right)_{x=0} \quad (18)$$

We have thus the same equations as those for an EC mechanism with an apparent constant  $\kappa = \kappa_1 + \kappa_1'$  for both polarography<sup>23</sup> and CV,<sup>24,25</sup> which allows us to calculate their sum  $\kappa_1 + \kappa_1'$  using eqns. (2) and (4).

The individual values of  $\kappa_1$  and  $\kappa_1'$  can be deduced from their ratio and their sum.

The variations of  $\log \kappa_1$  and  $\log \kappa_1'$  with pH are shown in Fig 12. For the reaction  $P \rightarrow \text{DH}$ ,  $\log \kappa_1'$  is independent of pH, whereas a slope of  $-2$  is obtained for the variations of  $\log \kappa_1$  ( $P \rightarrow \text{GHA}$ ). This result, which implies that  $\kappa$  has the form  $\kappa_0[\text{H}^+]^{-2}$ , seems *a priori* difficult to explain; however, it can be accounted for on the basis of a surface vs. volume aspect of the chemical reaction;<sup>26</sup> when a heterogeneous electrochemical reaction is followed by a first-order surface chemical reaction whose rate constant  $\kappa_{s1}$  is of the form  $\kappa_{s0}[\text{H}^+]^{-1}$ , the rate constant  $\kappa_1$ , which is measured experimentally, is an apparent 'volume' constant proportional to  $\kappa_{s1}$ <sup>26</sup> and it is indeed proportional to  $[\text{H}^+]^{-2}$ .

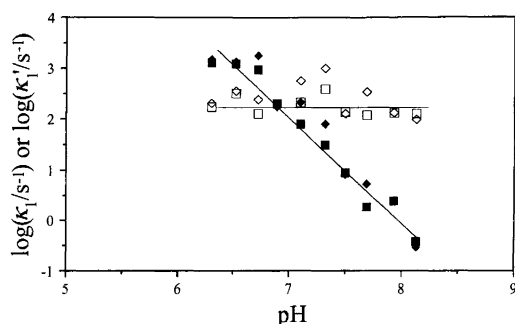


Fig. 12. Variations of  $\log \kappa_1$  and  $\log \kappa_1'$  with pH: (■, ◆)  $\log \kappa_1$ ; (□, ◇)  $\log \kappa_1'$ . For □ and ■, the sum  $\kappa_1 + \kappa_1'$  was obtained by polarography, for ◇ and ◆ by CV.

The 'heterogenous' nature of the electrochemical reaction (even though this 'heterogenous' character is only apparent<sup>26,27</sup>) is attested by the 'diffusional' character of the voltammetric peaks (Figs. 4 and 5).

This analysis was confirmed by the study of isonicotinic anilide;<sup>20</sup> in this case, the electrochemical reaction is of a surface nature and a slope of  $-1$  is then logically found.<sup>20</sup>

## Conclusions

The studies of Lund, based on polarographic data and on macroscale electrolyses<sup>1</sup> have firmly established the general mechanism of eqn. (1). Our work shows that a fast first-order reaction, yielding an intermediate P, occurs after the first  $2e^-$  reversible reduction of the C=O group. Up to pH about 6, P rearranges (reaction  $C_1$ ) to the *gem*-hydroxylamine, which loses  $NH_3$  to give the aldehyde; the same type of process is found during the reduction of isonicotinic acid and its analogues.<sup>4,12</sup> Above about pH 6, moreover, P reacts differently (reaction  $C_1'$ ) to give the 1,4-dihydropyridine-4-carboxamide isolated by Lund;<sup>1</sup> this reaction competes with the rearrangement to the *gem*-hydroxylamine. Our results show that the dihydropyridine is not formed below pH about 6 because under these circumstances, reaction  $C_1$  becomes faster than reaction  $C_1'$ .

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