The Oxovanadyl(IV) catalysed Oxidation of Adrenaline by Molecular Oxygen

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The oxidation of adrenaline (H_2LH^+) by molecular oxygen in the presence of catalytic amounts of VO^{2+} ions has been followed using a Clark-type oxygen electrode. The empirical rate law $-d[O_2]/dt = k_{obs.}[O_2] + p$ was obtained in which p is a small constant (only observable at low pH) and $k_{obs.}$ is given by the relationship (i), where $[VO]_T$ and $[L]_T$ are total initial concentrations of

$$k_{obs.} = \frac{A[VO]_T[L]_T}{B[H^+]^2 + C[L]_T + D[L]_T^2} + \frac{E}{[H^+]} + F \qquad \text{(i)}$$

VO^{2+} and adrenaline respectively. It is demonstrated that this behaviour is consistent with the involvement of both VO^{2+} and V^{2+} species and furthermore that it also calls for the presence of a tris(adrenaline) species, $V(LH)_3^+$. A kinetically determined equilibrium constant for reaction (ii) is

$$VO(LH)_2 + H_2LH^+ \rightleftharpoons V(LH)_3^+ + H_2O \qquad \text{(ii)}$$

reported. The necessary proton– and metal–ligand equilibrium constants were obtained by pH- and ^1H n.m.r.-titration techniques. All solutions were made up to a constant ionic strength (I = 0.100 mol dm^{-3}) with KNO_3 at 25.00 ± 0.02 °C.

The oxidation of adrenaline to adrenochrome has been the subject of many investigations in the past and the (overall) two-electron steps shown in the Scheme are well established.[1-3] (Note that the nomenclature listed in the Scheme is used throughout this paper; phenolic protons are written to the left of the symbol L, the amino proton to the right.)

The choice of vanadium(IV), *i.e.* VO^{2+}, as the catalyst was made on two grounds. First preliminary experiments showed it to be an effective catalyst giving rates measurable by means of an oxygen-sensitive electrode, and secondly in the hope that this catalyst would encourage two-electron oxidation steps. The latter point would seem to be confirmed in that the addition of species such as acrylonitrile gave no signs of the formation of polymeric material during the course of the reaction; solutions allowed to stand after completion of a run did, however, react positively.

The use of an oxygen-electrode method and initial oxygen saturation ($[O_2]$ initially 1.3×10^{-3} mol dm^{-3}) meant that at least a ten-fold range of adrenaline concentrations could be studied whilst maintaining the effective constancy of the latter during an oxidation run. The maintenance of constant $[H^+]$ by means of a pH-stat obviated the use of buffers (potential complexing agents!) and enabled a wide range of $[H^+]$ to be investigated. The thermostatted bath was oil-filled, and this prevented interference by photocatalysis.

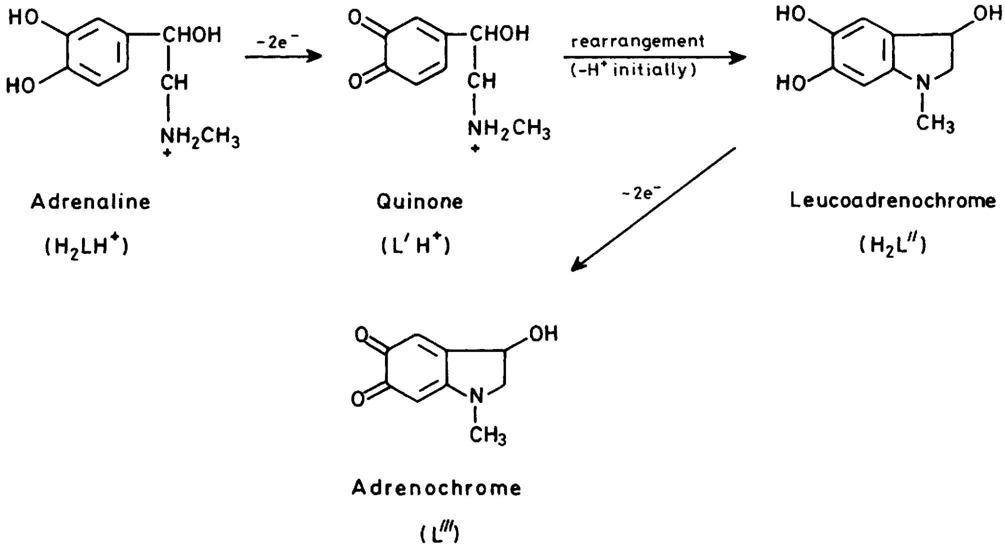
Results

Stoicheiometry.—The stoicheiometry of the reaction could not, unfortunately, be established beyond all doubt, but the following evidence strongly supports that given in the Scheme. (*i*) By use of minimal excess of oxygen it was found that two moles of oxygen were consumed for each mole of adrenaline oxidised. (*ii*) The pink colour that developed during the reaction was characterised by maxima at *ca.* 225, 305, and 485 nm in the u.v.–visible spectrum, indicating that it was indeed due to adrenochrome.[4,5] (*iii*) The production of adrenochrome was followed in several runs by using the maximum at 485 nm and a molar absorption coefficient of 3.77×10^3 dm^3 mol^{-1} cm^{-1} obtained by dissolving pure adrenochrome in 0.100 mol dm^{-3} KNO_3 solution (a value of 3.66×10^3 dm^3 mol^{-1} cm^{-1} is derivable from data reported by Mattok[5]). The concentration of adrenochrome was found to increase almost linearly with time (*i.e.* zero order) until *ca.* 80% of the oxygen had been consumed, at which point it levelled off sharply. By the time that no more oxygen could be detected the adrenochrome concentration had started to fall again, but this could not be followed as the solutions had started to become cloudy (probably due to melanin formation[4]). In no experiment in which the adrenochrome concentration was measured did it exceed 93% of the theoretical, but this is not surprising in view of subsequent reactions that take place.

In acid solutions adrenochrome decomposes, even in the absence of oxygen or catalysts to form precursors of melanin, a brown polymeric material.[4] Furthermore, vanadium(IV) compounds themselves react with hydrogen peroxide,[6] one of the products of the reaction. Fortunately this latter reaction seemed not to take place during the greater part of the reaction with oxygen studied here as is demonstrated by the absence of free radicals ($OH^•$) in the solutions during the course of the reaction. Indeed, even after the reaction with oxygen was virtually complete it was some time before the existence of free radicals (which are formed in the reaction[6] between VO^{2+} and H_2O_2) could be demonstrated. This is almost certainly because the very strong complexes with adrenaline stabilise the lower oxidation state of vanadium, and indeed may not interact with H_2O_2 in this pH range. Finally it is important to note that the concentration of H_2O_2 is never very high at any stage of the reaction.

Dependence of the Rate on $[O_2]$.—Over the pH range studied (4.2—5.4) the oxygen dependence was strictly first order in $[O_2]$, except for results obtained at the lowest pH where plots of $-d[O_2]/dt$ *vs.* t showed a small but definite intercept. The basic experimental rate law is thus as in equation (1), where p is small

$$-d[O_2]/dt = k_{obs.}[O_2] + p \qquad (1)$$



Scheme.

Table 1. Typical values of first-order rate constants, $k_{obs.}$, for the rate of consumption of oxygen: I = 0.100 mol dm^{-3} (KNO_3), T = 25.00 ± 0.02 °C, total vanadyl concentration, $[VO]_T$ = 2.941 × 10^{-4} mol dm^{-3}

pH	$[L]_T$/mol dm^{-3}	$10^2\ k_{obs.}$/min^{-1}	pH	$[L]_T$/mol dm^{-3}	$10^2\ k_{obs.}$/min^{-1}
4.28	0.0248	3.29	4.97	0.0249	4.19
	0.0199	2.79		0.0199	4.76
	0.0149	2.67		0.0150	5.05
	0.0099	2.21		0.0099	5.58
	0.0050	1.51			
			5.09	0.0239	4.38
4.54	0.049	4.13		0.0191	4.98
	0.0199	3.81		0.0174	5.64
	0.0149	3.45		0.0149	4.99
	0.0099	3.08		0.0125	5.90
	0.0050	2.60		0.0099	6.13
4.79	0.0249	3.74	5.36	0.0249	4.97
	0.0199	4.05		0.0199	5.50
	0.0149	4.68		0.0174	6.22
	0.0124	4.63		0.0150	6.33
	0.0099	5.35		0.0099	6.51

and only apparent at the highest $[H^+]$ studied. Therefore quantitative values of p, the oxygen-independent term, are not used in the ensuing discussion. Typical values of $k_{obs.}$ are listed in Table 1.

Dependence of $k_{obs.}$ on $[VO]_T$ and $[L]_T$.—A series of runs was carried out in which the total initial vanadyl concentration, $[VO]_T$, was varied between 1.0 × 10^{-4} and 6.0 × 10^{-4} mol dm^{-3} at various, constant, adrenaline concentrations, $[L]_T$. These results are summarized in Figure 1 in which it can be seen that there is an intercept on the $k_{obs.}$ axis when $k_{obs.}$ is plotted against $[VO]_T$, and that this intercept is independent of $[L]_T$. This established the form of $k_{obs.}$ to be as in equation (2), *i.e.*

$$k_{obs.} = \{a[VO]_T \cdot f([L]_T) + b\} \qquad (2)$$

to contain a term, b, independent of both $[VO]_T$ and $[L]_T$.

The dependence of the first term of equation (2) on $[L]_T$ was expected to be complex because of the interaction of vanadyl and adrenaline in solution and thus the form of $f([L]_T)$ was elucidated as follows. Figure 2 shows the variation of the slopes of the $[VO]_T$ *vs.* $k_{obs.}$ curves with both $[L]_T$ and $[H^+]$. The general pattern of these curves taken together with the fact that $f([L]_T)$ most likely reflects the speciation of the solutions led to an attempt to fit the curves to equation (3). For this purpose, c

$$\text{Slope } (k_{obs.}\ vs.\ [VO]_T) = c[L]_T/(d + e[L]_T + f[L]_T^2) \qquad (3)$$

was arbitrarily assigned the value 10^3 and d, e, and f determined algebraically from the data. Values of d, e, and f are listed in Table 2 and the goodness-of-fit illustrated by using them to draw the solid curves through the experimental points in Figure 2.

Dependence of Kinetic Constants on $[H^+]$.—The overall variation of the rate of reaction with hydrogen ion concentration is extremely complex, but fortunately it proved relatively straightforward to obtain simple relationships between the constants obtained at each (constant) pH and $[H^+]$.

First, Figure 3 shows that e and f [equation (3)] are independent of $[H^+]$ whereas d varies linearly with $[H^+]^2$. Secondly, the intercepts, b, of the $k_{obs.}$ *vs.* $[VO]_T$ curves are

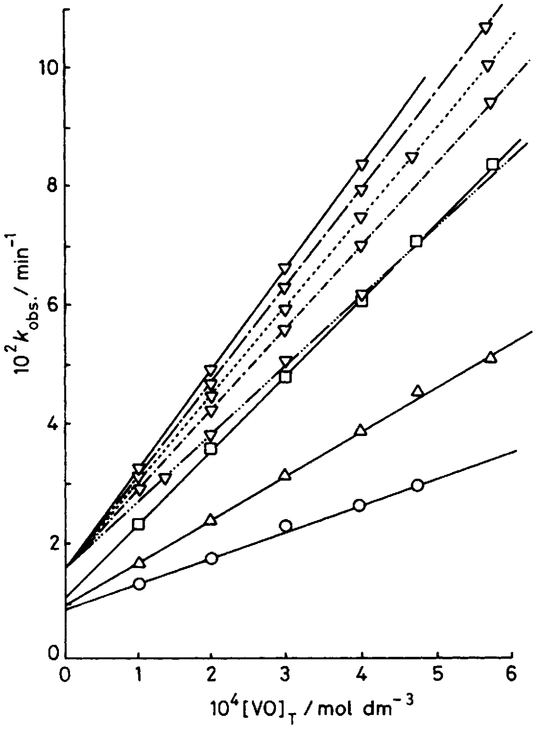


Figure 1. Dependence of $k_{obs.}$ on catalyst concentration at various pH values and adrenaline concentrations: pH 4.28 (○), 4.54 (△), 4.79 (□), 5.36 (▽); $[L]_T$/mol dm^{-3} = 0.0100 (——), 0.0125 (—·—·—), 0.0175 (·····), 0.0200 (–·–), 0.0250 (—··—)

plotted against $1/[H^+]$ in Figure 4 and clearly fit the relationship (4). Thus the overall rate expression can be written

$$b = E/[H^+] + F \qquad (4)$$

as in equation (5). The oxygen-independent term p, although

$$-d[O_2]/dt = \left(\frac{A[VO]_T[L]_T}{B[H^+]^2 + C[L]_T + D[L]_T^2} + \frac{E}{[H^+]} + F\right)[O_2] + p \qquad (5)$$

clearly observable at low pH, could not be extracted from the data; the values of the constants are as follows: $A = 10^3$ (arbitrarily assigned), $B = 1.01 \times 10^8$, $C = 3.53$, $D = 180$, $E = 2.21 \times 10^{-8}$, $F = 8.82 \times 10^{-3}$.

Determination of Proton and Metal Stability Constants.—The proton– and metal–ligand equilibrium constants were obtained in the usual way, that is by titration of various ligand and metal + ligand mixtures with potassium hydroxide under oxygen-free nitrogen maintaining a constant ionic strength (0.100 mol dm^{-3} KNO_3) and a temperature of 25.00 ± 0.02 °C. The results are collated as log β values in Table 3 where β is defined by the general expressions (6) and (7). Because the

$$m(VO^{2+}) + lL^{n-} + hH^+ \rightleftharpoons (VO)_mL_lH_h^{(2m+h-ln)+} \qquad (6)$$

$$\beta_{mlh} = \frac{[(VO)_mL_lH_h^{(2m+h-ln)+}]}{[VO^{2+}]^m[L^{n-}]^l[H^+]^h} \qquad (7)$$

proton association constants for the amino proton and one of the phenyl protons overlap, the microconstants relating to these protons were required: these had previously been determined by ^1H n.m.r. measurements.[7]

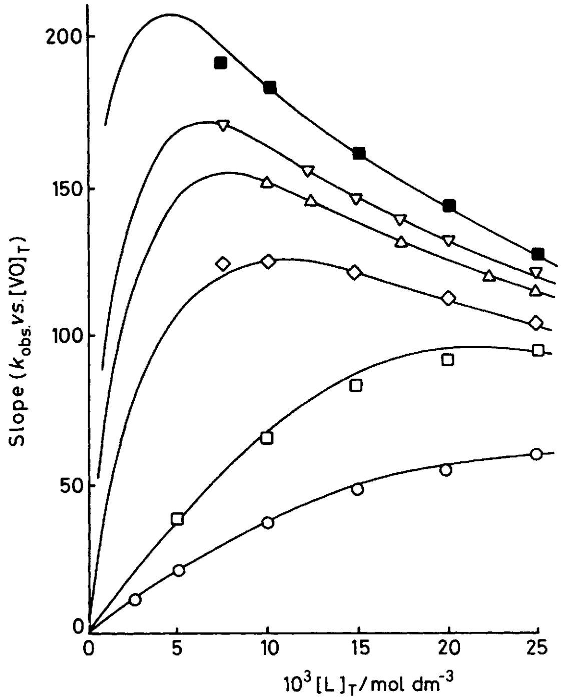


Figure 2. Variation of slope ($k_{obs.}$ *vs.* $[VO]_T$) with total adrenaline concentration. Experimental points and lines drawn using the data in Table 2. pH 4.28 (○), 4.54 (□), 4.79 (◇), 4.97 (△), 5.09 (▽), 5.36 (■)

Interpretation of the Rate Law.—It now remains to show that the overall rate expression (5) is indeed explicable in terms of a plausible reaction mechanism. In order greatly to simplify the presentation of this the following abbreviations have been adopted: adrenaline = H_2LH^+, quinone = $L'H^+$, leucoadrenochrome = H_2L'', adrenochrome = L'''; total concentration of vanadium = $[VO]_T$, total concentration of adrenaline = $[L]_T$.

The reactions postulated are (8)—(17). Note that the mono-

(*a*) Reduction of VO^{2+}

$$VO(LH)^+ \xrightarrow{k_1} V^{2+} + L'H^+ \qquad (8)$$

$$VO(L) \xrightarrow{k_2} V^{2+} + L' \qquad (9)$$

$$V^{2+} + H_2LH^+ \rightleftharpoons V^{II}(LH)^+ + 2H^+ \qquad (10)$$

(*b*) Formation of leucoadrenochrome

$$L' \xrightarrow{k_3} H_2L'' \qquad (11)$$

(*c*) Interaction of VO^{2+} and V^{2+} species

$$V^{II}(LH)^+ + VO(LH)^+ \xrightarrow{k_4} VOV^{4+} + 2LH^- \qquad (12)$$

(*d*) Reactions with O_2

$$VO(LH)_2 + O_2 \xrightarrow{k_5} VO(LH)^+ + L'H^+ + O_2{}^{2-} \qquad (13)$$

$$V^{II}(LH)^+ + O_2 \xrightarrow{k_6} V^{2+} + L'H^+ + O_2{}^{2-} \qquad (14)$$

$$V^{II}L'' + O_2 \xrightarrow{k_7} V^{2+} + L''' + O_2{}^{2-} \qquad (15)$$

$$VOV^{4+} + O_2 + H_2O \xrightarrow{k_8} 2VO^{2+} + O_2{}^{2-} + 2H^+ \qquad (16)$$

$$O_2{}^{2-} + 2H^+ \xrightarrow{\text{fast}} H_2O_2 \qquad (17)$$

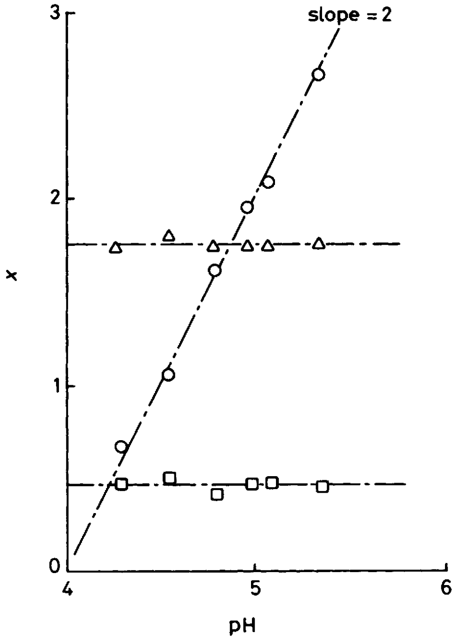


Figure 3. Dependence of x on pH of reaction mixtures, where $x = -\log d$ (○), $-\log e + 1$ (□), or $-\log f + 4$ (△)

Table 2. Values of d, e, and f required to fit the experimental data to the curves: slope ($k_{obs.}$ *vs.* $[VO]_T$) = $c[L]_T/(d + e[L]_T + f[L]_T^2)$ (c arbitrarily assigned the value 10^3)

pH	10^3d	e	$10^{-2}f$
4.28	219	3.44	1.91
4.54	85.7	3.26	1.61
4.79	24.0	3.79	1.92
4.97	11.2	3.66	1.82
5.09	8.15	3.96	1.86
5.36	2.22	3.56	1.66

co-ordinated VO^{2+} species do not react with O_2 whereas the di-co-ordinated species, $VO(LH)_2$, does. This would be explicable if the $VO(LH)_2$ species was electronically closely related to a V^{II} species co-ordinated to two semiquinone radicals, *i.e.* a large degree of charge transfer takes place, but does not lead to a complete redox reaction. The chain reaction is in fact maintained in this scheme by V^{II} (*i.e.* V^{2+}) species.

The consumption of oxygen thus follows: equation (18). The

$$-d[O_2]/dt = k_5[VO(LH)_2][O_2] + k_6[V^{II}(LH)^+][O_2] + k_7[V^{II}L''][O_2] + k_8[VOV^{4+}][O_2] \quad (18)$$

steady-state approximations (19)—(25) then apply. Substitution

$$d[VOV^{4+}]/dt = 0 \Longrightarrow k_4[V^{II}(LH)^+][VO(LH)^+] = k_8[VOV^{4+}][O_2] \quad (19)$$

$$d[V^{II}]/dt = 0 \Longrightarrow k_1[VO(LH)^+] + k_2[VO(L)] = k_4[V^{II}(LH)^+][VO(LH)^+] \quad (20)$$

$$\Longrightarrow [V^{II}(LH)^+] = k_1/k_4 + k_2/K^{H}_{VO(LH)}k_4[H^+] \quad (21)$$

$$d[L'']/dt = 0 \Longrightarrow k_3[L'] = k_7[V^{II}L''][O_2] \quad (22)$$

$$d[L']/dt = 0 \Longrightarrow k_1[VO(LH)^+] + k_2[VO(L)] + k_5[VO(LH)_2][O_2] + k_6[V^{II}(LH)^+][O_2] = k_3[L'] \quad (23)$$

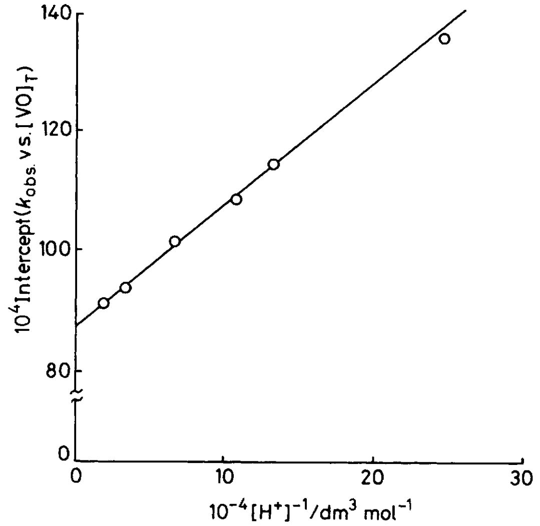


Figure 4. Variation of intercept ($k_{obs.}$ *vs.* $[VO]_T$) with $[H^+]^{-1}$

Table 3. Potentiometrically determined equilibrium constants for the vanadyl–adrenaline system at 25.00 ± 0.02 °C, I = 0.100 mol dm^{-3} (KNO_3)

	Concentration maxima		
Species[a]	occur at pH	with percentage VO^{2+}	log β
HL^-			13.15 ± 0.15
H_2L			23.02 ± 0.01
H_3L^+			31.65 ± 0.01
$VO(OH)^+$	5.4	1	−5.66[b]
$[VO(OH)]_2^{2+}$	5.4	1	−6.67[b]
$VO(LH)^+$	4.4	70	26.81 ± 0.01
$VO(L)$	5.3	16	21.42 ± 0.03
$VO(LH)_2$	8.2	90	51.63 ± 0.02
$VO(LH)(L)^-$	10.2	40	42.01 ± 0.02
$VO(L)_2^{2-}$	11.0	85	32.01 ± 0.02
$VO(LH)^+ \rightleftharpoons VO(L) + H^+$			−5.39
$VO(LH)_2 \rightleftharpoons VO(LH)(L)^- + H^+$			−9.62
$VO(LH)(L)^- \rightleftharpoons VO(L)_2^{2-} + H^+$			−10.00

[a] Note that the protonated ligand, adrenaline, is written H_3L^+ and not H_2LH^+ as elsewhere, since the values do not refer to microconstants. [b] Ref. 9.

into equation (18) yields equation (24), which must now be expressed in terms of $[VO]_T$ and $[L]_T$.

$$-d[O_2]/dt = 2k_5[VO(LH)_2][O_2] + 2k_6\{k_1/k_4 + k_2/k_4[H^+]K^{H}_{VO(LH)}\}[O_2] + 2k_4[V^{II}(LH)^+][VO(LH)^+] \quad (24)$$

In order to do this we require the equilibrium constants listed in Table 4, noting that values for the metal ion–ligand interactions were obtained by using the relevant microconstants (allowing for complex formation *via* the phenolic protons only). Although there was no evidence for a tris(adrenaline) complex from the e.m.f. data, the addition of a third ligand had to be allowed for in order to satisfy the observed kinetics. Since, as will be shown below, this had to involve no further loss of protons after the formation of the bis(adrenaline) complex, $VO(LH)_2$, reaction (25) was postulated. This would explain the

Table 4. Equilibria in the vanadyl–adrenaline system; symbols and values used in the kinetic studies

Equilibrium	Symbol	log value
$LH^- + 2H^+ \rightleftharpoons H_2LH^+$	β_2^H	21.91
$LH^- + VO^{2+} \rightleftharpoons VO(LH)^+$	$\beta_{VO(LH)}$	17.04
$2LH^- + VO^{2+} \rightleftharpoons VO(LH)_2$	$\beta_{VO(LH)_2}$	32.08
$VO(L) + H^+ \rightleftharpoons VO(LH)^+$	$K^H_{VO(LH)}$	5.39
$VO(LH)_2 + H_2LH^+ \rightleftharpoons V(LH)_3{}^+ + H_2O$	$\beta_{V(LH)_3}$	1.67*

* Kinetically determined, (see text).

$$VO(LH)_2 + H_2LH^+ \overset{\beta_{V(LH)_3}}{\rightleftharpoons} V(LH)_3{}^+ + H_2O \quad (25)$$

difficulty in obtaining evidence for it from pH-titration data, and furthermore parallels the existence of the $[V(L)_3]^{2-}$ ion in the catechol–vanadyl system.[8–10] It would also explain the relative slowness to reach equilibrium noted for the titration of these systems.[9,10]

Since the adrenaline remains protonated over the pH range studied (4.2—5.4) we can write equations (26) and (27). The total vanadyl concentration is given by equation (28) which

$$[L]_T = [H_2LH^+] \quad (26)$$

$$[LH^-] = [L]_T/\beta_2^H\,[H^+]^2 \quad (27)$$

$$[VO]_T = [VO^{2+}] + [VO(OH)^+] + [VO(LH)^+] + [VO(L)] + [VO(LH)_2] + [V(LH)_3{}^+] \quad (28)$$

$$= [VO(LH)^+] + [VO(L)] + [VO(LH)_2] + [V(LH)_3{}^+] \quad (29)$$

simplifies to equation (29) since the concentrations of free VO^{2+} and hydroxo species are negligible in the solutions used for the kinetic runs.[9,11] This yields equation (30). The first term of the theoretically derived rate equation (24) thus becomes equation (31) using equations (27) and (30).

$$[VO]_T = [VO^{2+}]\{\beta_{VO(LH)}[L]_T/\beta_2^H[H^+]^2 + \beta_{VO(L)}[L]_T/\beta_2^H[H^+]^3 + \beta_{VO(LH)_2}[L]_T{}^2/(\beta_2^H)^2[H^+]^4 + \beta_{V(LH)_3}\beta_{VO(LH)_2}[L]_T{}^3/(\beta_2^H)^2[H^+]^4\} \quad (30)$$

$$\frac{2k_5[VO]_T[L]_T[O_2]}{\{\beta_2^H\beta_{VO(L)}[H^+]/\beta_{VO(LH)_2}\} + \{\beta_2^H\beta_{VO(LH)}[H^+]^2/\beta_{VO(LH)_2}\} + [L]_T + \beta_{V(LH)_3}[L]_T{}^2} \quad (31)$$

If we now assume that the concentration of VO(L) is small compared with that of $VO(LH)^+$ and that the kinetic·data are only accurate enough to detect the term in the denominator in equation (31) that depends on $[H^+]^2$, then to a fair approximation the $[O_2]$-dependent term becomes that shown by (32). This exactly parallels the first term in the experimental rate equation (5) and thus the experimental values of *A*, *B*, *C*, and *D* can be used to calculate the equilibrium constants appearing in the term (32); log $[\beta_2^H\beta_{VO(LH)}/\beta_{VO(LH)_2}]$ = 7.49, when calculated in this way, can be compared with the value 6.95 obtained from the thermodynamic data. This is extremely close agreement under the circumstances and hence gives considerable justification for the use of *A*, *C*, and *D* to obtain a value of $\beta_{V(LH)_3}$; calculation from *A*, *C*, and *D* yields log $\beta_{V(LH)_3}$ = 1.67.

$$\frac{2k_5[VO]_T[L]_T[O_2]}{\{\beta_2^H\beta_{VO(LH)}[H^+]^2/\beta_{VO(LH)_2}\} + [L]_T + \beta_{V(LH)_3}[L]_T{}^2} \quad (32)$$

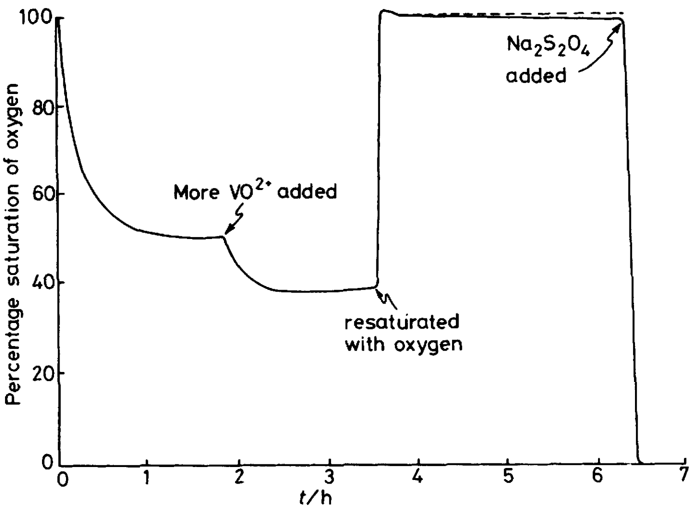


Figure 5. High-pH studies: effect of precipitation of $VO_2{}^+$-containing polymer

The only rate constant that can be obtained from the data is k_5 for which the mean value is 142 dm^3 mol^{-1} min^{-1} or 2.37 dm^3 mol^{-1} s^{-1}. Comparing the experimental rate equation (5) with (24) shows that $E = 2k_2k_6/k_4K^H_{VO(LH)}$ and $F = 2k_6k_1/k_4$ and that $p = 2k_4[V^{II}(LH)^+][VO(LH)^+]$. Although this value of *p* can in no way be completely substantiated, it seems logical in that it would increase with $[VO(LH)^+]$ and this indeed reaches a maximum at pH $\simeq$ 4.4 (Table 3).

If, on the other hand, $k_1 \simeq k_2$ [which seems reasonable, see equations (8) and (9)], then $F/E \simeq K^H_{VO(LH)}$. It is pleasing then to see that log (*F*/*E*) obtained from the values quoted above is 5.50 which compares well with the experimentally determined value of log $K^H_{VO(LH)}$ in Table 4 (*i.e.* 5.39). Conversely, if log $K^H_{VO(LH)}$ is taken to have the experimentally determined value of 5.39, then $k_1 = 1.6\,k_2$. This again would seem to be reasonable as the nitrogen proton could well facilitate electron transfer by interaction with the vanadyl oxygen atom (the side-chain of adrenaline is certainly long enough) thus making reaction (8) more favourable than reaction (9). A study of the anaerobic oxidation of adrenaline by vanadyl ions is at present being undertaken and this should enable values of k_1 and k_2 to be obtained independently.

Finally, a weakening of the V=O bond is almost certainly also involved in the charge transfer (or at least partial charge transfer) invoked as the reason for the interaction of O_2 with the $VO(LH)_2$ species rather than with the mono-chelated species [equation (13)]. This in turn must be linked with the formation of $V(LH)_3{}^+$ in which the vanadyl oxygen has been displaced completely.

High-pH *Studies.*—Reactions carried out at higher pH, namely between 6.5 and 8, were particularly interesting, in that the uptake of oxygen ceased at well above zero concentration despite the large excess of adrenaline. This was accompanied by the formation of a fine dark blue precipitate. On addition of more vanadyl ion the consumption of O_2 continued again, but again stopped short of total removal. The solution could then be resaturated with oxygen, whereupon the solution remained saturated for a considerable time until removed with metabisulphite. This behaviour is illustrated in Figure 5.

We interpret these results on the basis that the precipitate is a melanin-like polymer that entraps the vanadyl ions within it. Thus the solution is *totally free of transition-metal ions* and under these conditions there is no reaction of the adrenaline with oxygen. Precisely the same effect is observable in the case of the iron(III)-catalysed oxidation of L-3,4-dihydroxyphenylalanine.[12] It is possible, therefore, that the instability of solutions of catechols towards oxidation by air is due to the presence of traces of transition-metal ions and that their complete exclusion would lead to stable solutions. Note that complexing agents such as ethylenediaminetetra-acetate are not successful in achieving this because of the lability of the complexes.

Experimental

Adrenaline was Fluka 'Puriss' grade and was used without further purification. Stock solutions of $VOCl_2$ were prepared from V_2O_5 solutions in sulphuric acid and reduced by sulphur dioxide. To avoid the problem of sulphate interaction with vanadyl ion, barium chloride was used to precipitate quantitatively the sulphate as $BaSO_4$. The resulting solution of vanadyl chloride was checked for vanadyl by permanganate titration and standardized potentiometrically for $[H^+]$.

For both kinetic and thermodynamic work experiments were made in 0.100 mol dm^{-3} KNO_3 and a thermostatted oil-bath was used to maintain the temperature at 25.00 ± 0.02 °C.

A Radiometer E5046 oxygen-sensitive electrode was used to measure the rate of oxygen consumption. The pH was measured with a Russell CMAT micro combination electrode and was maintained at a constant value in kinetic runs by a Radiometer TT1 pH-stat fitted with an SBR2 recorder. Pure oxygen was bubbled through the cell containing adrenaline in 0.100 mol dm^{-3} KNO_3 until the $[O_2]$ had reached its standard value. The cell was then closed to the atmosphere and after a stable $[O_2]$ reading had been achieved the catalyst solution was added through a micro-syringe. The resulting $[O_2]$ *vs.* time curve was recorded on a Watanabe pen recorder. In the kinetic study the concentrations of adrenaline were varied in the range 0.005—0.025 mol dm^{-3} and the concentrations of vanadyl from 1×10^{-4} to 6×10^{-4} mol dm^{-3}. Kinetic measurements were performed at six different pH values in the range 4.28—5.36.

In the pH-titration work the concentration of adrenaline was between 2×10^{-3} and 4×10^{-3} mol dm^{-3} and the metal ion:ligand ratio was 1:1, 1:2, 1:4, or 1:6. The dissociation constant of the second phenolic hydroxy group of the ligand was determined as described earlier.[13] All titrations were made under an atmosphere of pure N_2 [O_2 content reduced by use of vanadium(II) sulphate solution]. The potentiometric measurements were made on a Radiometer PHM64 instrument.

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References

1 J. Harley-Mason, *J. Chem. Soc.*, 1950, 1276.
2 J. D. Bu'Lock and J. Harley-Mason, *J. Chem. Soc.*, 1951, 712.
3 R. A. Heacock, *Chem. Rev.*, 1959, **59**, 181.
4 J. D. Bu'Lock, *J. Chem. Soc.*, 1961, 52.
5 G. L. Mattok, *J. Chem. Soc. A*, 1965, 4278.
6 H. B. Brooks and F. Sicilio, *Inorg. Chem.*, 1971, **10**, 2530.
7 R. F. Jameson, G. Hunter, and T. Kiss, *J. Chem. Soc., Perkin Trans. 2*, 1980, 1105.
8 R. P. Henry, P. C. H. Mitchell, and J. E. Prue, *J. Chem. Soc. A*, 1971, 3392.
9 R. P. Henry, P. C. H. Mitchell, and J. E. Prue, *J. Chem. Soc., Dalton Trans.*, 1973, 1156.
10 D. K. Bhatacharya and S. N. Baneiji, *Z. Anorg. Allg. Chem.*, 1962, **118**, 315.
11 A. Komura, M. Hayashi, and H. Imanaga, *Bull. Chem. Soc. Jpn.*, 1977, **50**, 2927.
12 R. F. Jameson, *Proc. 11th Int. Conf. Co-ord. Chem.*, Haifa, 1968, 23.
13 J. E. Gorton and R. F. Jameson, *J. Chem. Soc. A*, 1968, 2616.

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