The Mechanism of a Rate-Controlled Cross-Disproportionation between Phenoxyl Radicals[1]

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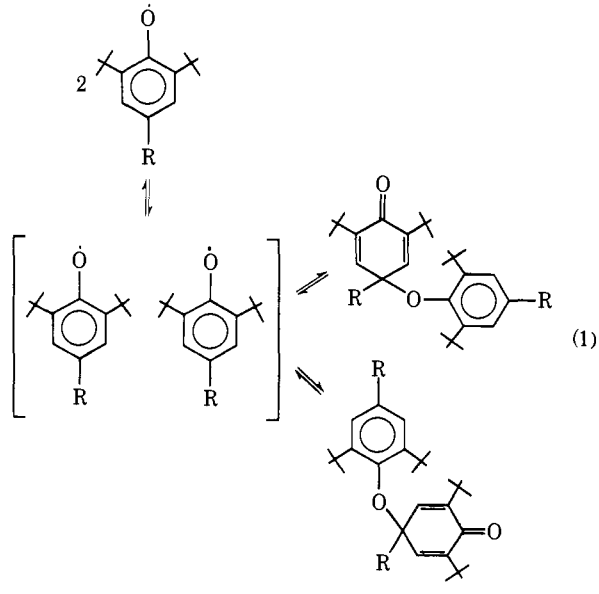
Abstract: The reaction of galvinoxyl with trihydrogalvinoxyl leads to hydrogalvinoxyl in the stoichiometric proportions of 2:1:3, respectively. The experimental rate law, measured by following the decay of the galvinoxyl esr signal, was found to be second order in galvinoxyl, first order in trihydrogalvinoxyl, and inverse order in hydrogalvinoxyl. The activation parameters were $\Delta H^{\ddagger} = 7.7$ kcal/mol and $\Delta S^{\ddagger} = -42.8$ gibbs/mol, indicating that steric congestion is the principal factor in the rate-limiting step. The isotope effect for hydrogen transfer from carbon to oxygen was found to be 3.6, but no isotope effect for hydrogen transfer from oxygen to oxygen was observed. This implies that the methylene hydrogen of trihydrogalvinoxyl rather than the hydroxyl hydrogen is transferred in the slow step. Although the free energy of activation and thus the rates of hydrogen transfer were constant in a large number of diverse solvents, it was found that $\Delta H^{\ddagger}$ and $\Delta S^{\ddagger}$ for the hydrogen transfer in benzene, acetone, acetonitrile, and cyclohexane followed a linear relationship with an isokinetic temperature of 293.8°K. The mechanism is discussed in terms of initial fast reversible oxygen–oxygen hydrogen transfer between galvinoxyl and trihydrogalvinoxyl, followed by rate-controlled cross-disproportionation between galvinoxyl (G) and the dihydrogalvinoxyl (GH_2) radical.

Phenoxyl radicals, sterically blocked by *tert*-butyl groups in the 2,6 positions, may suffer self-coupling or self-disproportionation depending on the structure of the para substituent. Phenoxyls bearing small- or medium-sized groups at the para position with no α hydrogens, *e.g.*, acetyl, pivaloyl, carbomethoxy, and cyano substituents, exist as dimers at subambient temperatures, but on heating dissociate into measurable concentrations of the monomeric radicals. A recent study on the mechanism of self-coupling using the nmr line-width broadening technique has revealed a number of interesting features about this fundamental step in free-radical reactions.[3] It was shown that a short-lived (10^{-7} sec) caged radical pair intervened between the monomeric and dimeric phenoxyls (eq 1).



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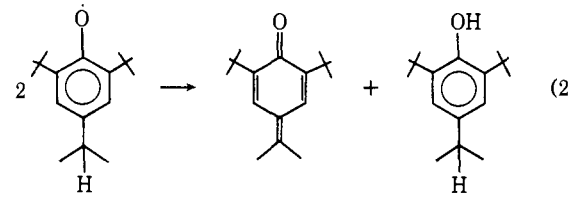
On the other hand, the self-disproportionation process predominates when primary or secondary alkyl groups serve as para substituents (eq 2). Several mechanistic



(2)

studies on this hydrogen-transfer step have been documented.[4,5]

When the para substituent is a tertiary alkyl group, *e.g.*, trityl, *tert*-butyl, or *tert*-amyl, the phenoxyl radicals are perfectly stable and exist in solution as monomers rather than dimers even at subambient temperatures.[6] Their stability is due to the fact that self-coupling is sterically prohibited; however, when exposed to hydrogen donors such as hydrocarbons, amines, mercaptans, or phenols, these phenoxyls are rapidly converted into their respective phenols and cross-coupling products. The most detailed mechanistic study of such a hydrogen-transfer process relates to the reaction of tri-*tert*-butylphenoxyl with unhindered phenols (eq 3).[7–10] In all the cases examined, the hydrogen-transfer step (k_1) is rate determining, followed by fast cross-coupling (k_2). In fact, using the nmr and esr techniques, it was possible to observe short-lived (10^{-9} sec), hydrogen-bonded complexes between tri-*tert*-butylphenoxyl and tri-*tert*-butylphenol.[11,12]

(1) Financial assistance from the National Science Foundation, the Petroleum Research Fund, and the A. P. Sloan Foundation is gratefully acknowledged.

(2) M.S. Thesis, Aug 1969.

(3) D. J. Williams and R. Kreilick, *J. Amer. Chem. Soc.*, **90**, 2775 (1968).

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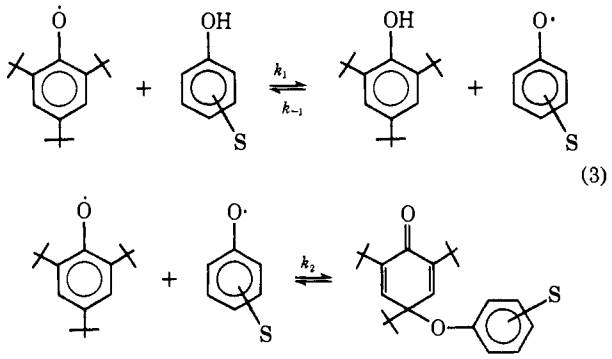
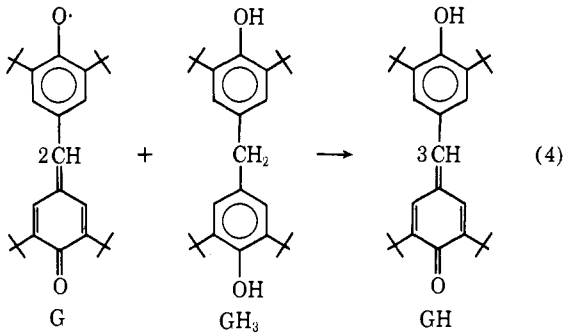


(3)

Several years back it was observed that the stable phenoxyl radical galvinoxyl (G)[13] reacted with its bisphenolic precursor trihydrogalvinoxyl (GH_3) to give the ketophenol hydrogalvinoxyl (GH), as shown below (eq 4).[14] The mechanism of this reaction should re-



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semble that of eq 3 in that the first step should consist of reversible hydrogen transfer to produce GH and the GH_2 radical; however, the subsequent step must entail cross-disproportionation of the (G, GH_2) radical pair rather than cross-coupling in order to account for the fact that hydrogalvinoxyl (GH) is the exclusive product (eq 5). In fact, in view of the structural

$$\begin{aligned} G + GH_3 &\underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} GH + GH_2 \\ G + GH_2 &\xrightarrow{k_2} 2GH \end{aligned} \qquad (5)$$

uniqueness of the galvinoxyl system, it is conceivable that the cross-disproportionation (k_2) is rate determining, preceded by fast, reversible hydrogen transfer (k_1, k_{-1}). In this paper we report on the mechanism of this novel reaction of a phenoxyl radical.

Experimental Section

Materials. **Galvinoxyl** was prepared by the procedure of Coppinger.[15] Iodometric titration of a sample recrystallized six times from isooctane was shown to be 98 ± 1% pure.

Hydrogalvinoxyl was prepared by treating a solution of 10.6 g (0.025 mol) of galvinoxyl in 150 ml of benzene with an excess of sodium iodide (20 g) dissolved in acetone. The deep brown color instantly changed to an intense purple red, which on addition of acetic acid gave a brownish yellow solution. Then 25 ml of a 1 *N* sodium thiosulfate solution was added to decolorize the liberated iodine. The bright canary yellow benzene layer was washed several times with water, dried over anhydrous magnesium sulfate, and, after removal of the solvent, 8.5 g of crude hydrogalvinoxyl was obtained. The material was purified by chromatography on a silicic acid column (25 g of silicic acid/g of substance) eluting with chloroform. The eluted product (mp 156–156.5°) on recrystallization from aqueous ethanol gave a total of 7.0 g (66%) of canary yellow crystals, mp 156.5–157° (lit.[16] mp 158–159°).

Trihydrogalvinoxyl was obtained from the Ethyl Corporation and purified by chromatography on an alumina column (50 g of alumina/g of substance) eluting with carbon tetrachloride. The combined nonyellow fractions on recrystallization from pentane gave the pure product, mp 154.5–155°.

3,5,3′,5′-Tetra-*tert*-butyl-4,4′-dihydroxybenzophenone[17] was prepared by chromic acid oxidation of hydrogalvinoxyl in acetone. A 61.5% yield was obtained, mp 222–223° (aqueous ethanol).

3,5,3′,5′-Tetra-*tert*-butyl-4,4′-dihydroxydiphenyldeuteriomethane[17] was prepared from the above benzophenone by reduction with lithium aluminum deuteride in 80% yield, mp 154.5–155° (aqueous ethanol).

O-Deuterated phenols were prepared by refluxing 0.1 *M* solutions in O-deuterated methanol for 6 hr and allowed to crystallize. The crystals were collected and dried under high vacuum; all operations were carried out in the drybox.

Solvents were purified according to standard procedures.[18]

Products and Stoichiometry. A solution of 422 mg (1.00 mmol) of galvinoxyl and 212.5 (0.5 mmol) of trihydrogalvinoxyl in 30 ml of freshly distilled benzene was deaerated with pure nitrogen gas, sealed, and allowed to stand at 25° until the color of the solution had changed from deep brown to bright canary yellow (3 days). The solvent was evaporated and the residue was recrystallized from aqueous methanol, affording pure hydrogalvinoxyl, mp 156.5–157°.

The stoichiometry of the galvinoxyl–trihydrogalvinoxyl reaction was determined spectrophotometrically. Stock solutions of galvinoxyl (2×10^{-3} *M*) and trihydrogalvinoxyl (1.25×10^{-3} *M*) were prepared in pure benzene. Appropriate aliquots of each stock solution were pipetted into a 10-ml volumetric flask and diluted to the mark, and their absorbances measured at 773 nm. The results are summarized in Table I.

Table I. Stoichiometry of the Galvinoxyl–Trihydrogalvinoxyl Reaction in Benzene

Solution[a]			
ml of G (2×10^{-3} *M*)	ml of GH_3 (1.25×10^{-3} *M*)	Molar ratio of GH_3 to G	Absorbance[b] (773 nm)
5	0	0.000	0.614
5	1	0.125	0.454
5	2	0.250	0.300
5	4	0.500	0.071

[a] Diluted to 10-ml total volume. [b] Measured after 6 hr of reaction time.

Kinetics. The reaction tube used in the kinetic measurements is displayed in Figure 1. Arm A was charged with a solution of galvinoxyl and hydrogalvinoxyl, while arm B was charged with a solution of trihydrogalvinoxyl. All solutions were prepared double the desired final concentration in order to make up for the dilution factor. For the degassed runs, the tube was connected to a high-vacuum system (Pope Co.) and the unmixed solutions were degassed simultaneously by the freeze–pump–thaw method (five cycles) and then sealed under a nitrogen atmosphere at the constriction. For the undegassed runs, arm B was sealed with a rubber septum. At time t_0 the solution in arm B was transferred carefully into arm A, thoroughly mixed, afterward shaken into the analyzer tube (arm C), and placed into the cavity of the Varian E-3 electron spin resonance spectrometer.

Prior to mixing of the solutions, the esr spectrometer was allowed to warm up until maximum stability was achieved (usually 2 hr). The variable temperature assessory was adjusted to the desired temperature and the temperature fluctuations minimized to ±0.5°K. This was only possible by regulating the water inlet temperature

(13) The official nomenclature is 2,6-di-*tert*-butyl-(3,5-di-*tert*-butyl-4-oxo-2,5-cyclohexadiene-1-ylidene)-*p*-tolyloxy for galvinoxyl (G), 2,5,-3′,5′-tetra-*tert*-butyl-4,4′-dihydroxydiphenylmethane for trihydrogalvinoxyl (GH), and 3,5,3′,5′-tetra-*tert*-butyl-4-(3,5-di-*tert*-butyl-4-hydroxybenzal)cyclohexadione for hydrogalvinoxyl (GH).

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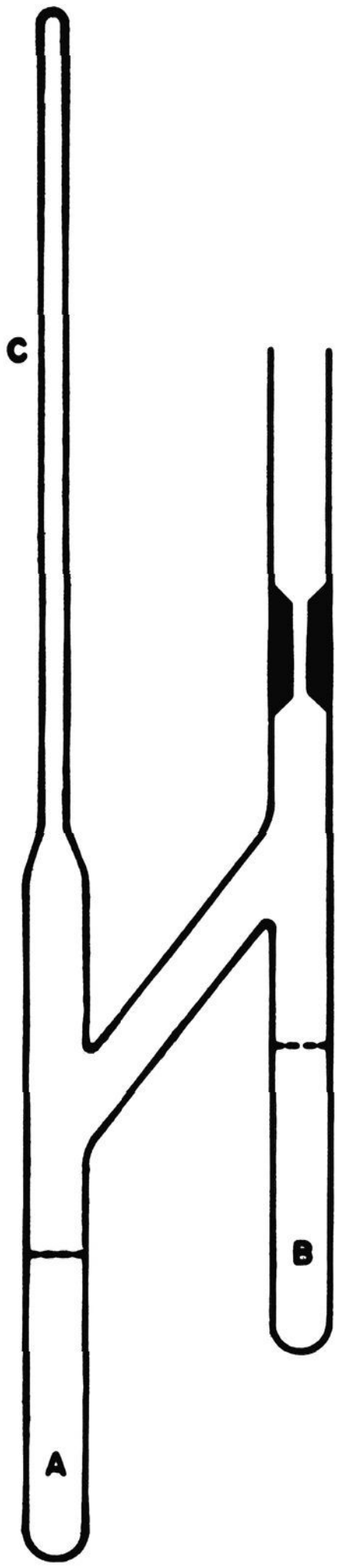


Figure 1. Kinetic reaction tube for the galvinoxyl–trihydrogalvinoxyl reaction.

of the electromagnet within 1–2°K of the desired experimental temperature, employing a closed loop heat exchanger system (Neslab Co.). Once the esr system had stabilized, the appropriate magnetic field setting was selected by scanning a standard sample of galvinoxyl, prepared at the desired concentration of the kinetic run. The signal was overmodulated to avoid hyperfine structure, and the magnetic field dial was locked at maximum signal height. The standard galvinoxyl sample was removed while leaving all the settings unchanged. The recorder needle was placed at t_0 + 25 sec, the above described mixing operation carried out as rapidly as feasible (less than 25 sec), the reaction tube placed into the esr cavity, and the recorder started to trace the signal at exactly t_0 +

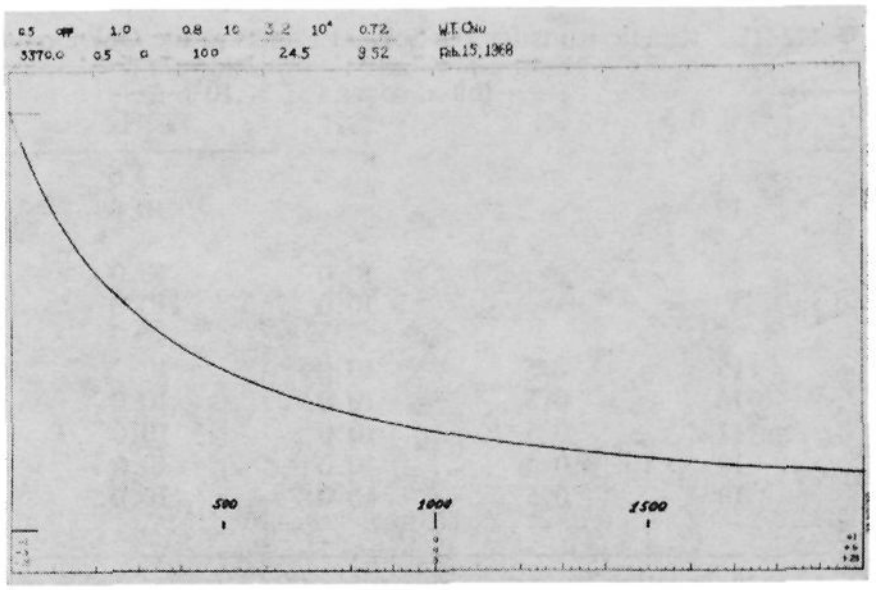


Figure 2. Esr signal height *vs.* time record of the galvinoxyl–trihydrogalvinoxyl reaction.

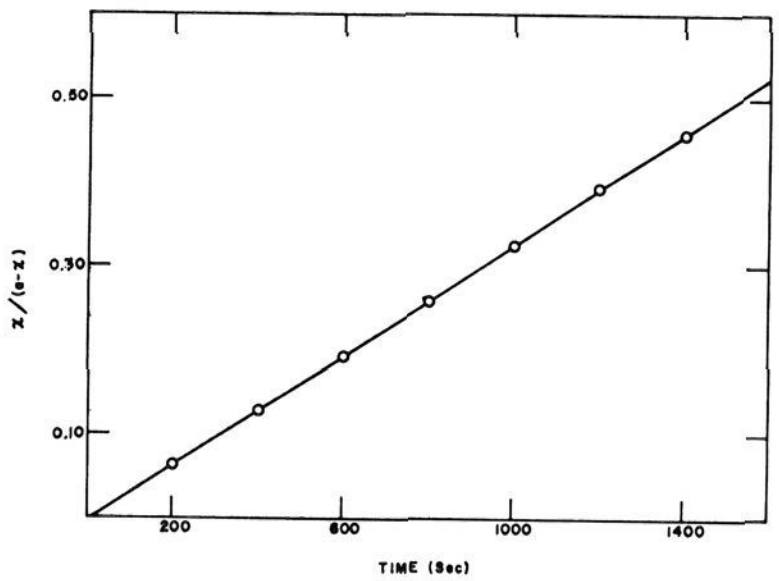


Figure 3. Pseudo-second-order plot of the galvinoxyl consumption in the galvinoxyl–trihydrogalvinoxyl reaction.

25 sec. A typical signal height *vs.* time record is exhibited in Figure 2.

The runs were analyzed in terms of pseudo-second-order kinetics utilizing an IBM-1130 computer. Straight-line plots of $x/(a - x)$ *vs.* t were observed through two half-lives for most cases. A typical graph is demonstrated in Figure 3. The kinetic runs for establishing the rate law are summarized in Table II. The

Table II. Kinetic Runs for the Rate Law of the Galvinoxyl–Trihydrogalvinoxyl Reaction in Cyclohexane at 295°K

Run	Initial concn ($M \times 10^3$) G	GH	GH_3	k_{obsd},[a] M^{-1} sec^{-1}
1	0.50	10.0	10.0	4.95 ± 0.08
2	0.50	20.0	20.0	5.06 ± 0.06
3	0.50	30.0	30.0	4.92 ± 0.04
4	0.50	20.0	10.0	5.24 ± 0.12
5	0.50	30.0	10.0	4.88 ± 0.06
6	0.50	10.0	20.0	4.24 ± 0.22
7	1.00	20.0	20.0	2.96 ± 0.32
8	0.25	10.0	10.0	6.72 ± 0.36
9	0.05	10.0	10.0	8.20 ± 0.30
10	0.50	10.0	10.0	5.12[b]
11	0.50	10.0	10.0	5.03[c]

[a] Average of at least three independent measurements, but only one representative run has been tabulated; $k_{obsd} = 2k_1k_2/k_{-1}$. [b] In the presence of oxygen gas. [c] In the presence of nitrogen gas.

solvent data are given in Table III and the data for the calculation of the activation parameters according the Eyring equation are given in Table IV. The kinetic isotope effects are collected in Table V. In order to minimize hydrogen–deuterium exchange

Table III. Kinetic Runs for the Solvent Effects in the Galvinoxyl–Trihydrogalvinoxyl Reaction at 295°K

Run	Initial concn ($M \times 10^3$) G	GH	GH_3	Solvent	Dielectric constant (25°)	k_{obsd},[b] M^{-1} sec^{-1}
1	0.5	10.0	10.0	Cyclohexane	2.02	4.94 ± 0.08
12	0.5	10.0	10.0	Carbon tetrachloride	2.22	5.59 ± 0.45
13	0.5	10.0	10.0	Benzene	2.27	3.98 ± 0.01
14	0.5	10.0	10.0	Tetrahydrofuran	8.20	4.82 ± 0.07
15	0.5	10.0	10.0	Pyridine	12.30	6.90 ± 0.23
16	0.5	10.0	10.0	Acetone	20.70	5.94 ± 0.06
17	0.5	10.0	10.0	Ethanol[a]	24.30	16.60 ± 1.78
18	0.5	10.0	10.0	Acetonitrile	36.20	7.08 ± 0.08
19	0.5	10.0	10.0	Dimethylformamide	37.00	8.96 ± 0.16

[a] Large deviation from pseudo-second-order kinetics; value represents initial slope of 10% reaction. [b] $k_{obsd} = 2k_1k_2/k_{-1}$.

Table IV. Kinetic Runs for the Solvent Effects on the Activation Parameters of the Galvinoxyl[a]–Trihydrogalvinoxyl[b] Reaction in the Presence of Hydrogalvinoxyl[c]

Run	Solvent	Temp,[d] °K	k_{obsd},[e] M^{-1} sec^{-1}	$\Delta H^{\pm}$, kcal M^{-1}	$\Delta S^{\pm}$, gibbs M^{-1}	$\Delta G^{\pm}$, at 303°K, kcal M^{-1}
20	Benzene	293	3.072			
21	Benzene	303	4.730	7.21	−45.37	20.96
22	Benzene	313	7.454			
23	Cyclohexane	293	4.780			
24	Cyclohexane	303	8.160	7.68	−42.84	20.66
25	Cyclohexane	313	11.840			
26	Acetonitrile	293	4.318			
27	Acetonitrile	303	7.098	9.99	−35.33	20.70
28	Acetonitrile	313	13.708			
29	Acetone	293	3.522			
30	Acetone	303	6.608	10.37	−34.32	20.77
31	Acetone	313	11.734			

[a] 0.5×10^{-3} M initial concentration. [b] 10.0×10^{-3} M initial concentration. [c] 10.0×10^{-3} M initial concentration. [d] Controlled within ±0.5°K. [e] Average of at least three independent measurements, but only one representative run has been tabulated; $k_{obsd} = 2k_1k_2/k_{-1}$.

Table V. Kinetic Runs for the Isotope Effects of the Galvinoxyl–Trihydrogalvinoxyl in Cyclohexane at 295°K

Run	Reaction	k_{obsd},[a] M^{-1} sec^{-1}
32	G–GH–GH_3 in H_2O	4.92 ± 0.50
33	G–GD[b]–GHD_2[b] in D_2O	4.95 ± 0.56
34	G–GH–GDH_2[c] in H_2O	1.37 ± 0.03
35	G–GD[b]–GD_3[d] in D_2O	1.40 ± 0.06

[a] Average of at least three independent measurements, but only one representative run has been tabulated; $k_{obsd} = 2k_1k_2/k_{-1}$. [b] Deuterated at phenolic hydroxyl groups. [c] Deuterated at methylene carbon only. [d] Deuterated at phenolic hydroxyl groups and methylene carbon.

of the O-deuterated phenols due to adventitious water, the stock solutions were saturated with deuterium oxide and allowed to equilibrate several hours. Also, a few drops of deuterium oxide were placed into each arm of the reaction tube.

Discussion of the Results

Products and Stoichiometry. When galvinoxyl (G) is treated with trihydrogalvinoxyl (GH_3), the exclusive reaction product is hydrogalvinoxyl (GH), identified by mixture melting point and spectral data with the authentic material. This product is formed quantitatively as evidenced by the fact that the product solution at 385 nm showed an ϵ of 32,800, while for pure hydrogalvinoxyl ϵ is 33,800 at this wavelength. The results of Table I indicate that the stoichiometry of the G–GH_3 reaction is 2:1:3, respectively, for G, GH_3, and GH (eq 6).

$$2G + GH_3 \longrightarrow 3GH \quad (6)$$

Experimental Rate Law. The kinetics of this reaction, as already described in the Experimental Section, was studied by following the decay of the galvinoxyl esr signal as a function of time. Initial efforts revealed that the reaction was so fast that by the time the reaction tube was placed into the esr cavity (less than 25 sec), most of the signal had already decayed. However, in the presence of excess hydrogalvinoxyl the rate could be slowed down sufficiently to carry out the kinetics *via* conventional techniques without recourse to rapid detection techniques or flow methods. A reaction sample, allowed to stand 14 days, representing t_∞, showed no esr signal. By varying initial concentrations of galvinoxyl, hydrogalvinoxyl, and trihydrogalvinoxyl, the experimental rate law was established as $-d[G]/dt = k_{obsd}[G]^2[GH_3][GH]^{-1}$. The pertinent kinetic data (runs 1–11) are summarized in Table II. Of particular interest is the fact that k_{obsd} does increase appreciably (runs 7, 8, and 9) with decreasing galvinoxyl concentration, *e.g.*, a 20-fold decrease in [G] causes a 2.8-fold increase in k_{obsd}.

In order to establish whether molecular oxygen interfered with the G–GH_3 reaction,[14] the kinetics were examined in the absence and presence of oxygen. As runs 10 and 11 in Table II show, k_{obsd} is constant within

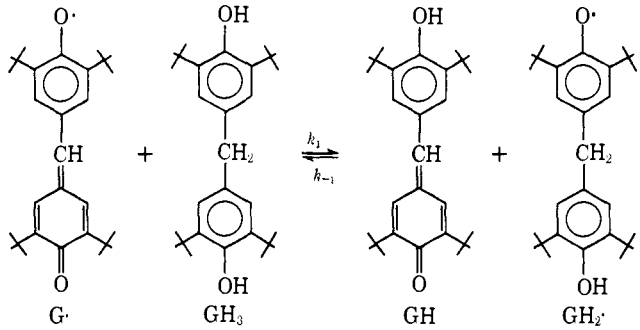
the experimental error. Thus the G–GH_3 reaction is insensitive to molecular oxygen.

Solvent Effect. To learn about the polar character of the G–GH_3 reaction, the kinetics were measured in a large variety of solvents. The data in Table III (runs 1 and 12–19) show that the k_{obsd} and thus the free energies of activation are quite insensitive to solvent polarity and solvent structure. A variation of about 18-fold in the dielectric constant caused only a twofold change in k_{obsd}. However, considering that one might be studying the kinetics of this reaction near the isokinetic temperature,[4,6] k_{obsd} was measured at 293, 303, and 313°K in benzene, cyclohexane, acetonitrile, and acetone. $\Delta H^{\pm}$ and $\Delta S^{\pm}$ were then computed from this data using the Eyring equation. The activation data (runs 20–31) are summarized in Table IV.

Thus the G–GH_3 reaction is solvent sensitive but $\Delta G^{\pm}$ and thus k_{obsd} remain effectively constant as a result of opposing trends in the solvent effects on $\Delta H^{\pm}$ and $\Delta S^{\pm}$. In fact, these activation data yield an excellent straight-line plot of $\Delta H^{\pm}$ *vs.* $\Delta S^{\pm}$ (Figure 4), with a slope of 293.8°K.

Isotope Effects. In order to unravel the finer details of the mechanism of this hydrogen-transfer process, the hydrogen–deuterium isotope effects for oxygen and carbon were determined. The data are collected in Table V (runs 32–35). As mentioned in the Experimental Section, to minimize hydrogen–deuterium exchange at the phenolic hydroxyl groups, the kinetics of the respective O-deuterated substances were conducted in the presence of deuterium oxide. As control experiments, the undeuterated substances were conducted in the presence of regular water. Runs 32 *vs.* 33 and 34 *vs.* 35 indicate that the hydrogen–deuterium isotope effect for oxygen is negligible. Runs 32 *vs.* 34 and 33 *vs.* 35 clearly show an appreciable hydrogen–deuterium isotope effect for carbon, *i.e.*, k_{C-H}/k_{C-D} = 3.55 ± 0.30.

Mechanism. From the above results the following conclusions emerge on which a reasonable mechanism



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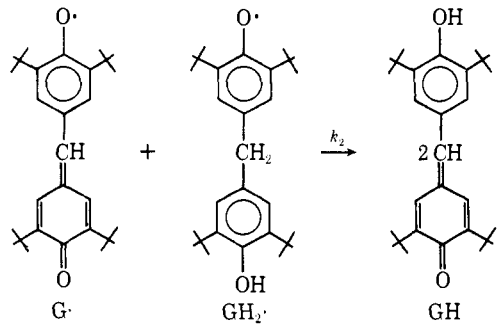


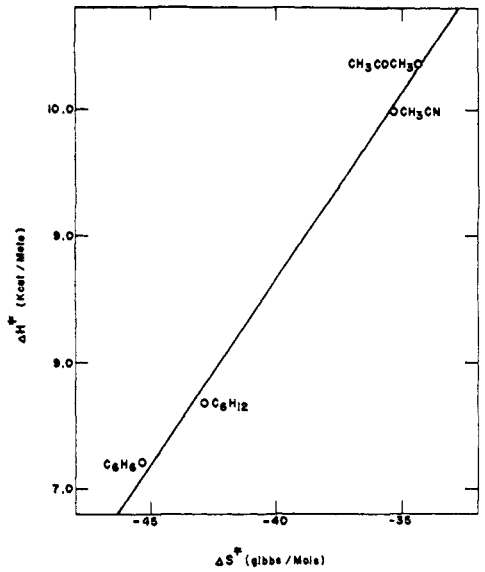


Figure 4. Activation enthalpy *vs.* entropy plot for the galvinoxyl–tryhdyrogalvinoxyl reaction.

for the hydrogen-transfer process of the galvinoxyl–trihydrogalvinoxyl reaction must be based: (a) the hydrogalvinoxyl is the exclusive product; (b) the limiting rate law suggests a reversible step involving GH, followed by an irreversible slow step; (c) the low-activation enthalpy implies that energywise the hydrogen transfer proceeds with relative ease, but the large and negative entropy of activation demands a highly structured transition state; (d) the observation of an isokinetic temperature of 293.8°K implies strong dipolar interaction of the solvent with the activated complex for the hydrogen transfer between the radicals G and GH_2; (e) the presence of a hydrogen–deuterium isotope effect for carbon but none for oxygen illlustrates that the hydrogen transfer originates from the central methylenic group of trihydrogalvinoxyl in the slow step and not from the phenolic hydroxyl group.

A reasonable mechanism which accommodates the above experimental conclusions is given in eq 7. In shorthand notation we have

$$G + GH_3 \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} GH_2 + GH$$

$$G + GH_2 \xrightarrow{k_2} 2GH$$

Summation of both steps gives the experimentally established stoichiometry

$$2G + GH_3 \longrightarrow 3GH$$

Employing the steady-state approximation,[8,10] it has been shown that this two-step mechanism leads to the rate equation

$$-d[G]/dt = 2k_1k_2[GH_3][G]^2/(k_2[G] + k_{-1}[GH])$$

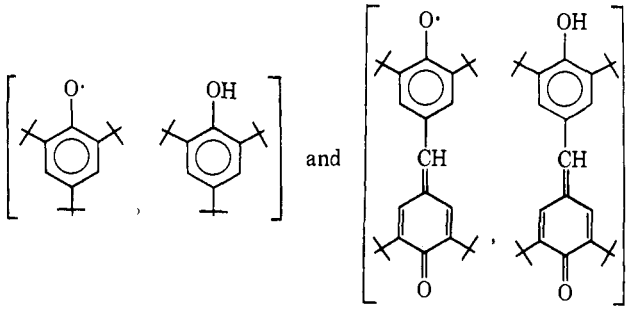
Now, if the assumption is made that $k_{-1}[GH] \gg k_2[G]$, dropping $k_2[G]$ in the denominator of the above rate equation yields the experimental law

$$-d[G]/dt \cong k_{obsd}[GH_3][G]^2/[GH]$$

where k_{obsd} equals $2k_1k_2/k_{-1}$.

It is now apparent that the experimental rate law is a limiting expression, only applicable under conditions

where $k_{-1}[GH] \gg k_2[G]$. To illustrate that at the initial concentrations of $[G] = 5 \times 10^{-4}$ M, $[GH] = 1 \times 10^{-2}$ M, and $[GH_3] = 1 \times 10^{-2}$ M, the above inequality holds, thus enabling us to examine the complex hydrogen-transfer kinetics in terms of simple pseudo-second-order kinetics, we have estimated k_1, k_{-1}, and k_2 from literature data and our own results. Mahoney, Ferris, and DaRooge[9] have shown that the differences in the standard heats of formation, that is, $\Delta H_f°[ArO] - \Delta H_f°[ArOH]$, are 29.14 and 26.4 kcal/mol, respectively, for the pairs

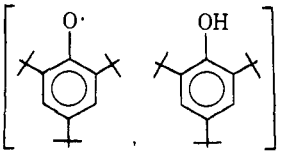


The standard heat of reaction for the first reversible step in our mechanism is given then by

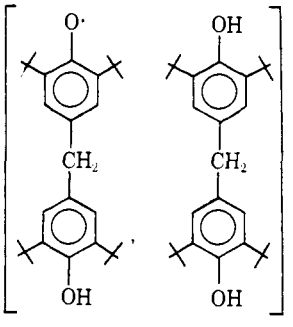
$$\Delta H° = (\Delta H_f°[GH_2] - \Delta H_f°[GH_3]) - (\Delta H_f°[G] - \Delta H_f°[GH])$$

$$\Delta H° = (29.14 - 26.40) \text{ kcal/mol} = 2.74 \text{ kcal/mol}$$

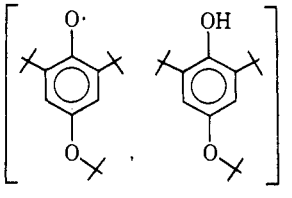
We are of course assuming here that the measured difference in the standard enthalpies of formation for the pair



is a reasonable estimate for the pair

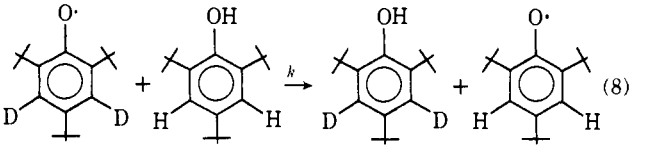


Structurally, the two pairs are very closely related and this assumption is expected to be a valid one. Furthermore, the standard entropy differences for the hydrogen transfers are not known. But in view of the fact[19] that the standard entropy difference is zero for the pair



it is justified to assume that the standard entropy difference is also zero for our hydrogen transfer.

With this simplification and the estimated reaction heat value, we establish that $\Delta G° \cong \Delta H° \cong 2.74$ kcal/mol. From the relationship $K = k_1/k_{-1} = e^{-\Delta G°/RT}$, we calculate that $k_1/k_{-1} = 0.01$ at 298°K. Since $k_{obsd} \cong 2k_1k_2/k_{-1} = 5.00$ M^{-1} sec^{-1} at these experimental conditions, after substituting $k_1/k_{-1} = 0.01$ we determine that $k_2 = 250$ M^{-1} sec^{-1}. A reasonable value for k_1 can be estimated from the data of Arick and Weissman.[12] They have shown by esr kinetics that for the hydrogen transfer in eq 8 the bimolecular rate constant is $k = 250$ M^{-1} sec^{-1}. If



this value can be taken as a reasonable guess for our reaction, we have $k_1 \cong 250$ M^{-1} sec^{-1}.[20] Substituting into $k_1/k_{-1} = 0.01$, we obtain that $k_{-1} \cong 25{,}000$ M^{-1} sec^{-1} and, therefore, reasonable values for the three essential bimolecular rate constants k_1, k_{-1}, and k_2 in our mechanism are on hand.

It is interesting to note that the first step (k_1) and the second step (k_2) proceed approximately at the same rate; however, due to the large reversible step (k_{-1}), the second step (k_2) becomes the rate-determining step because of the very low steady-state concentration of GH_2 that builds up. At the initial concentrations of $[GH] = 10^{-2}$ M and $[G] = 0.05 \times 10^{-2}$ M, we have $k_{-1}[GH] = 250$ sec^{-1} and $k_2[G] = 0.125$ sec^{-1}. Clearly, in the concentration range at which most of our kinetic data were collected it follows that $k_{-1}[GH] \gg k_2[G]$ and thus pseudo-second-order kinetics obtain.

Furthermore, it is also evident that as the galvinoxyl concentration is stepped up relative to hydrogalvinoxyl and trihydrogalvinoxyl, the condition $k_{-1}[GH] \gg k_2 \cdot [G]$ progressively deteriorates. Thus, deviations from the above limiting rate law should become more pronounced at higher concentrations of G. As a consequence, k_{obsd} is expected to decrease with increasing [G] because analysis of the rate data in terms of the limiting rather than the complete rate expression underrates the denominator. Indeed, the rate data in Table II attest to this effect since k_{obsd} decreases by a factor of about 3 when the [G] is increased 20-fold.

The most convincing corroboration of the suggested mechanism is the kinetic isotope effect. The first step in our mechanism (eq 7) engages a fast, reversible hydrogen transfer between the hydroxyl group of GH_3 and the oxyl site of G. The second step involves rate-determining hydrogen transfer from the methylene group of GH_2 to the oxyl site of G. Therefore, we anticipate a primary isotope effect in the hydrogen transfer from carbon to oxygen but none from oxygen to oxygen. These are exactly the experimental conclusions reached in Table V. The intermediate value of the primary isotope effect implies a relatively unsymmetrical or partially bent structure of the activated complex for the hydrogen transfer between G and

(19) C. D. Cook, C. B. Depatic, and E. S. English, *J. Org. Chem.*, **24**, 1356 (1959).

(20) This value appears to be a little large compared to the k_1 values experimentally measured by Mahoney and DaRooge[10] for the reaction shown in eq 3.

GH_2.[21] Examination of molecular models implies that the approach of the oxyl end of G toward the methylenic hydrogen of GH_2 should have an unsymmetrical or partially bent configuration due to steric congestion.

The activation data (Table IV) bear further evidence on the suggestion of steric hindrance in the approach of the G and GH_2 radicals. While the enthalpies of activation ($\Delta H^{\ddagger}$) are quite low for this hydrogen transfer, ranging between 7.21 and 10.37 kcal/mol, the entropies of activation ($\Delta S^{\ddagger}$) are very large, ranging between −45.37 and −34.32 gibbs/mol, but more important they are negative in sign. In other words, a highly structured transition is required for the hydrogen transfer. Again, as already mentioned in the interpretation of the isotope effect, the activation data confirm the supposition that a selective approach for the hydrogen transfer between the G and GH_2 radicals is mandatory.

Quite intriguing is the effect of the solvent on the activation parameters. Although the free energies of activation ($\Delta G^{\ddagger}$) remain constant at 295°K in a variety of solvents (Table III), we notice that the activation enthalpy ($\Delta H^{\ddagger}$) and entropy ($\Delta S^{\ddagger}$) increase in the order of hexane, benzene, acetonitrile, and acetone (Table IV). In the Gibbs-Helmholtz equation the enthalpy and entropy have opposite signs and therefore the increase in $\Delta H^{\ddagger}$ is offset by the increase in $\Delta S^{\mp}$. What is surprising is that these two opposing factors are so well balanced at 295°K that the free energies of activation and thus the k_{obsd} are constant within the experimental error (Table III). However, as we have shown in Figure 4, the slope of the $\Delta H^{\ddagger}$ *vs.* $\Delta S^{\ddagger}$ plot is 293.8°K. The isokinetic temperature of this reaction is thus very near to the experimental temperature (295°K) at which the solvent rate data were measured.

To account for the interesting entropy–enthalpy correlation observed in the hydrogen transfer between G and GH_2 we offer the following mechanistic interpretation concerning the transition state for this reaction.[4] In the reactant state G + GH_2 of the slow step (k_2) the radical sites are solvated by means of dipolar interaction with the solvent. As these solvated radicals approach each other to form the activated complex

$$[\overset{\delta\cdot}{G}H\cdots H\cdots\overset{\delta\cdot}{G}]^{\ddagger}$$

solvating solvent molecules must be pushed aside in order to allow a close contact of the methylenic hydrogen of GH_2 and the oxyl center of G in the hydrogen transfer between G and GH_2. Once the hydrogen-transfer act is completed, solvent molecules reorganize to solvate the product state 2GH. In view of the steric demand in the approach of G and GH_2 as pointed out before, desolvation of the respective radical sites to form the activated complex imposes a sizable energy barrier to this hydrogen transfer, but is accompanied by entropy production in the transition state. Therefore, in a polar solvent the radicals are more strongly solvated and a greater energy is required for desolvation, but at the same time a greater degree of disorder is created in the transition state relative to the reactant state. Consequently, an increase in the solvent polarity is expected to manifest itself in a larger activation enthalpy (greater desolvation barrier) which is counterbalanced by a more positive activation entropy (greater degree of disorder due to desolvation). These are the experimental trends observed for the hydrogen transfer between G and GH_2.[22]

In conclusion, the mechanism suggested in eq 7 accommodates all of the experimental facts concerning the G–GH_3 reaction in a reasonable way. The unique and novel feature of this work is that the second step, *i.e.*, the cross-disproportionation of the radical pair G,GH_2, is the rate-determining step. Usually, for example, in the reaction of 2,4,6-tri-*tert*-butylphenoxyl with substituted phenols,[8] such a radical pair preferentially cross-couples (eq 3) and the reversible hydrogen transfer between the phenolic hydroxyl and the oxyl radical is rate determining. Steric factors can hardly be responsible for the distinct behavior between galvinoxyl and 2,4,6-tri-*tert*-butylphenoxyl since para-substituted 2,6-di-*tert*-butylphenoxyls are known to dimerize,[3,4,23] with standard free energies of dissociation ranging between 3 and 6 kcal/mol and corresponding to dissociation constants of 10^{-3}–10^{-5} for the dimer. Presumably the equilibrium constant for cross-coupling of G and GH_2 into G–GH_2 must lie heavily on the side of the dissociated radical pair.[24] The most probable reason for this contrast is the additional resonance stabilization in galvinoxyl *vs.* para-substituted 2,6-di-*tert*-butylphenoxyls.

The cross-disproportionation of the radical pair [G, GH_2] rather than its cross-coupling appears to be the only efficient channel for the destruction of galvinoxyl in the G–GH_3 reaction. However, still more interesting is the fact that this cross-disproportionation is rate limiting as a result of the great steric demand for this hydrogen transfer. Usually, disproportionation of radical pairs, for example, the conversion of two ethyl radicals into ethane and ethylene, is so fast that the rates are diffusion controlled.[25] To the best of our knowledge, our hydrogen-transfer study provides the first example of a rate-controlled *cross*-disproportionation of a pair of phenoxy radicals. A rate-controlled *self*-disproportionation was already cited in eq 2, in which a pair of 2,6-di-*tert*-butyl-4-isopropylphenoxy radicals disproportionate to give quinone and phenol products with a rate constant of 2.0 M^{-1} sec^{-1}, about 100 times slower than observed here for the G,GH_2 pair. Most likely, steric factors are responsible for this rate difference.

(21) K. B. Wiberg, "Physical Organic Chemistry," Wiley, New York, N. Y., 1962, p 361.

(22) On the basis of dielectric constants it is expected that the activation enthalpy for acetonitrile should be higher than that for acetone, while in fact experimentally the reverse is observed. Probably specific solvation effects operate which are not incorporated in our crude model. Heats of solvation data could be helpful in shedding some understanding on this apparent anomaly.

(23) H. D. Becker, *J. Org. Chem.*, 30, 982 (1965).

(24) Efforts to pick up the esr signal of the GH_2 radical by employing galvinoxyl which was deuterated at the vinylic position unfortunately failed because of the complexity of the esr signal of the deuterated galvinoxyl.

(25) S. W. Benson "Thermochemical Kinetics," Wiley, New York, N. Y., 1968, p 100; S. W. Benson, *Advan. Photochem.*, 1, 10 (1964); J. A. Kerr and A. F. Trotman-Dickenson, *Progr. React. Kinet.*, 1, 11 (1961).