

# Articles

## Free Energy Hydride Affinities of Quinones in Dimethyl Sulfoxide Solution

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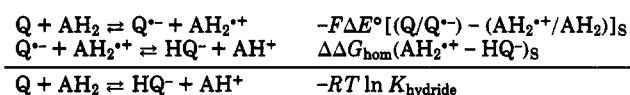
Free energy hydride affinities in solution defined by reaction i [(i)  $Q_s + H^-_s \rightleftharpoons (QH^-)_s$ ,  $\Delta G_{\text{hydride}}(Q)_s$ ] were evaluated for a number of quinones (Q) in dimethyl sulfoxide solution using reaction ii [(ii)  $-\Delta G_{\text{hydride}}(Q)_s = 2.303RTpK_a(QH^-)_s + F\Delta E^\circ_{\text{NHE}}[(Q^-/Q^{2-})_s + (Q/Q^{\cdot-})_s] + C$ ] derived from a thermochemical cycle. The constant C corresponds to  $-FE^\circ_{\text{NHE}}[(H^+/H^\cdot)_s + (H^\cdot/H^-)_s]$  which is equal to 69.9 kcal/mol in dimethyl sulfoxide. The  $pK_a$  of hydroquinone monoanions ( $QH^-$ ) were determined by the overlapping indicator method. Reversible electrode potentials for both the first and second charge transfers to Q were determined by cyclic voltammetry. Values of  $\Delta G_{\text{hydride}}(Q)_{\text{DMSO}}$  ranging from -58 to -101 kcal/mol were observed. Hydride affinities of substituted benzoquinones (BQ) were found to be linearly correlated with the corresponding electron affinities in solution. Equilibrium constants for the reactions of BQ with the NADH model compound 10-methylacridan were estimated, and the thermochemical results are discussed in terms of the one- and the two-step mechanisms of hydride-transfer reactions.

### Introduction

Hydride-transfer reactions between quinones (Q) and dihydroaromatic compounds ( $AH_2$ ) have been studied a great deal in recent years, especially in relation to the conversions of NADH (nicotinamide adenine dinucleotide) analogs to  $NAD^+$ . Much of the discussion of these reactions has been focused upon two different mechanistic possibilities: (i) direct hydride transfer and (ii) coupled electron and hydrogen atom transfers.<sup>3,4</sup> It is now generally accepted that both reaction pathways are feasible and that the pathway followed depends upon the one-electron reduction potential of the carbonyl compound.

It is of interest to note that the thermochemistry of the two mechanisms can be directly related by the cycle shown in Scheme I. The reactions involved in the two-step

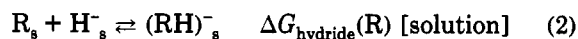
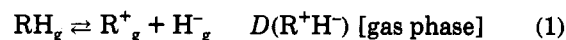
Scheme I  
free energy change



mechanism are shown above the line along with the pertinent free energy changes, and the sum of these, which corresponds to the single-step mechanism, is written below the single line. In this paper, we evaluate hydride affinities of a variety of quinones in dimethyl sulfoxide (DMSO) solution and discuss the thermochemistry of the two mechanisms for reactions of quinones with a typical NADH model compound.

Hydride affinities [ $D(R^+H^-)$ ] of carbenium ions in the

gas phase<sup>5,6</sup> defined as the enthalpy change for reaction 1 have been of great importance as quantitative measures



of the intrinsic stability of the ions. In terms of reaction chemistry, the free energy change for reaction 2 in solution [ $\Delta G_{\text{hydride}}(R)$ ] is more pertinent. The hydride ion acceptor (R) can be either a carbenium ion, a radical, or a closed-shell neutral substrate capable of accepting a hydride ion in a chemical or biochemical process.

Gas-phase data for reaction 1 reported over the past 15 years are primarily for relatively small aliphatic molecules or the corresponding phenyl derivatives.<sup>5,6</sup> In solution, the direct measurement of the free energy changes for reaction (2) is not feasible, due to the very high reactivities of the hydride ion, carbenium ions, and free radicals in solution, depending upon the nature of R.

We have recently derived the free energy hydride affinity values for a variety of carbenium ions<sup>7</sup> and radicals<sup>8</sup> in DMSO and acetonitrile solutions using experimental values of electrode potentials in eqs 3 and 4, respectively, derived from thermochemical cycles. In eqs 3 and 4  $\Delta G_{\text{hom}}(RH)$  are the free energy changes for homolytic RH bond cleavage, and  $\Delta E^\circ$  represents the electrode potential differences of the two redox potentials in parentheses. The subscript s refers to the solvent. As expected, the

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$$-\Delta G_{\text{hydride}}(\text{R}^+)_{\text{s}} = \Delta G_{\text{hom}}(\text{RH})_{\text{s}} - F\Delta E^{\circ}[(\text{H}^+/\text{H}^-) - (\text{R}^+/\text{R}^{\bullet})]_{\text{s}} \quad (3)$$

$$-\Delta G_{\text{hydride}}(\text{R}^{\bullet})_{\text{s}} = \Delta G_{\text{hom}}(\text{RH})_{\text{s}} - F\Delta E^{\circ}[(\text{H}^+/\text{H}^-) - (\text{RH}/\text{RH}^{\bullet})]_{\text{s}} \quad (4)$$

energies required to break the R-H bonds to generate  $\text{R}^+$  and  $\text{H}^-$  are substantially lower in solution as compared to those in the gas phase, a reflection of the exergonic solvation of the ions. Gas-phase data are not available for the heterolytic cleavage of  $\text{RH}^+$ , but the difference between these and the solution values are expected to be considerably less in this case since localization rather than creation of charge is involved in the cleavage reactions. Solution hydride affinities ranging from -76 to -129 kcal/mol for carbenium ions<sup>7</sup> and -22 to -72 kcal/mol for free radicals<sup>8</sup> were observed. A similar strategy was applied to estimate an absolute value of  $-\Delta G_{\text{hydride}}(\text{R}^+)$  for the  $\text{NAD}^+$  model,<sup>9</sup> 10-methylacridinium ion.

Quinones are the most frequently used hydride acceptors in  $\text{NADH}$  model studies<sup>10</sup> as well as in organic synthesis.<sup>11</sup> It is of considerable interest to evaluate  $\Delta G_{\text{hydride}}(\text{Q})_{\text{DMSO}}$  for comparison with those previously determined for carbenium ions and free radicals. Free energy hydride affinities in DMSO solution of quinones of widely varying structures fall in the range -58 to -101 kcal/mol, overlapping those of the free radicals at the low end of the scale and the carbenium ions at more negative values.

## Results and Discussion

**Thermochemical Relationship.** The free energy hydride affinities of quinones in solution, as defined in eq 5 where Q represents a quinone, can be obtained using eq 6 derived from the thermochemical cycle illustrated in Scheme II. The experimental quantities necessary to



$$-\Delta G_{\text{hydride}}(\text{Q})_{\text{s}} = 2.303RT\text{p}K_{\text{a}}(\text{QH}^-)_{\text{s}} + F\Delta E^{\circ}_{\text{NHE}}[(\text{Q}^{\bullet-}/\text{Q}^{2-}) + (\text{Q}/\text{Q}^{\bullet-})]_{\text{s}} + C \quad (6)$$

Scheme II

	free energy changes
$\text{QH}^- \rightleftharpoons \text{Q}^{2-} + \text{H}^+$	$2.303RT\text{p}K_{\text{a}}(\text{QH}^-)_{\text{s}}$
$\text{Q}^{2-} \rightleftharpoons \text{Q}^{\bullet-} + \text{e}^-$	$FE^{\circ}_{\text{NHE}}(\text{Q}^{\bullet-}/\text{Q}^{2-})_{\text{s}}$
$\text{Q}^{\bullet-} \rightleftharpoons \text{Q} + \text{e}^-$	$FE^{\circ}_{\text{NHE}}(\text{Q}/\text{Q}^{\bullet-})_{\text{s}}$
$\text{H}^+ + \text{e}^- \rightleftharpoons \text{H}^{\bullet}$	$-FE^{\circ}_{\text{NHE}}(\text{H}^+/\text{H}^{\bullet})_{\text{s}}$
$\text{H}^{\bullet} + \text{e}^- \rightleftharpoons \text{H}^-$	$-FE^{\circ}_{\text{NHE}}(\text{H}^{\bullet}/\text{H}^-)_{\text{s}}$
$\text{QH}^- \rightleftharpoons \text{Q} + \text{H}^-$	$-\Delta G_{\text{hydride}}(\text{Q})_{\text{s}}$

evaluate  $\Delta G_{\text{hydride}}(\text{Q})_{\text{s}}$  are the  $\text{p}K_{\text{a}}$  of the corresponding hydroquinone monoanion together with the first and second reduction potentials of the quinone.

The equilibrium acidities of the second proton transfer from the corresponding hydroquinone [ $\text{p}K_{\text{a}}(\text{QH}^-)_{\text{DMSO}}$ ] were measured using a modification of the indicator

overlapping method of Bordwell.<sup>12</sup> Details of the procedure are described in the Experimental Section. The first and the second reduction potentials of the quinones,  $E^{\circ}_{\text{NHE}}(\text{Q}/\text{Q}^{\bullet-})_{\text{DMSO}}$  and  $E^{\circ}_{\text{NHE}}(\text{Q}^{\bullet-}/\text{Q}^{2-})_{\text{DMSO}}$ , were determined using cyclic voltammetry (CV) and are reversible values for all but one of the quinones studied.

The electrode potentials for the  $(\text{H}^+/\text{H}^{\bullet})$  and  $(\text{H}^{\bullet}/\text{H}^-)$  redox couples are the only other data necessary for the evaluation of  $\Delta G_{\text{hydride}}(\text{Q})_{\text{DMSO}}$ . These potentials have previously been evaluated in DMSO<sup>7,8,13,14</sup> and are equal to -2.48 and -0.55 V, respectively. The sum of the free energies of these two half-reactions (Scheme II) is equal to 69.9 kcal/mol in DMSO, and this is the constant (C) in eq 6.

**Quinones Studied in this Work.** The data gathered in Table I are for *o*- and *p*-benzoquinones and derivatives designated as *o*-BQ or *p*-BQ. It was feasible to obtain  $\text{p}K_{\text{a}}(\text{QH}^-)$  for all of these compounds. The structures of the unsubstituted quinones are shown below.



**Hydride Affinities of Benzoquinones.** The thermodynamic driving force for the parent *p*-benzoquinone to accept a hydride ion, as reflected by the  $\Delta G_{\text{hydride}}$  value in Table I, is -70 kcal/mol. For *o*-benzoquinone  $\Delta G_{\text{hydride}}(\text{Q})_{\text{DMSO}}$  is 6 kcal/mol less negative and that of tetramethyl-*p*-benzoquinone is still less exergonic by another 6 kcal/mol. Although substituent effects on the neutral quinones cannot be neglected, it is tempting to rationalize the trends in  $\Delta G_{\text{hydride}}(\text{Q})_{\text{DMSO}}$  in terms of the greater effect on the negatively charged hydride adduct,  $\text{QH}^-$ . The substituent effects of the methyl groups<sup>15</sup> destabilize  $\text{QH}^-$  relative to the anion derived from the parent substance. In the case of chloro substituents the electron-withdrawing inductive effect apparently predominates over the electron-donating resonance effect, and the tetrachloro derivatives of both *o*-BQ and *p*-BQ are substantially more negative, by 27 and 13 kcal/mol, respectively. The effect of electron-withdrawing substituents on the hydride affinities of quinones is reflected in the large negative value (-101 kcal/mol) observed for dichlorodicyano-*p*-benzoquinone (DDQ).

A comparison of the  $\text{p}K_{\text{a}}(\text{QH}_2)_{\text{DMSO}}$  values<sup>12b</sup> for *o*-BQH<sub>2</sub> (15.8) and *p*-BQH<sub>2</sub> (19.8) reveals that the former is a substantially stronger acid. One contribution to this acidity difference is surely the stabilization of *o*-BQH<sup>-</sup> by internal hydrogen bonding. Since *o*-BQH<sup>-</sup> is also the product of the combination of *o*-BQ with hydride ion, an argument based on anion stabilization would predict that *o*-BQ is expected to be a better hydride acceptor than *p*-BQ which is contrary to results shown in Table I.

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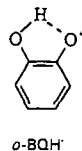
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Table I. Hydride Affinities of Quinones in Dimethyl Sulfoxide Solution

substrate	$pK_a(QH^-)^a$	$E_{NHE}(Q/Q^-)^b$	$E_{NHE}(Q^-/Q^{2-})^c$	$-\Delta G_{hydride}(Q)^d$
<i>p</i> -BQ	26.2	-0.319	-1.208	70
<i>o</i> -BQ	27.2	-0.492	-1.36 <sup>e</sup>	64
tetrachloro- <i>p</i> -BQ	17.6	0.170	-0.645	83
tetrachloro- <i>o</i> -BQ	18.9	0.348	-0.557	91
tetramethyl- <i>p</i> -BQ	27.1	-0.635	-1.498	58
DDQ	13.5	0.622	-0.048	101

<sup>a</sup> Equilibrium acidities of hydroquinone anion in DMSO measured in this work according to the method described in the Experimental Section. <sup>b</sup> Reversible reduction potentials (in V) of quinones in DMSO derived with CV at Pt working electrode (i.d. = 0.2 mm) in 0.1 M Bu<sub>4</sub>NPF<sub>6</sub>/DMSO vs normal hydrogen electrode (NHE); sweep rate, 1 V/s; counter electrode, Pt wire; internal standard, ferrocenium/ferrocene. <sup>c</sup> Reversible reduction potentials (in V) of quinone radical anions derived from the second reductive peaks. Experimental conditions are the same as described in footnote b. <sup>d</sup> Hydride affinities in kcal/mol calculated using eq 6. <sup>e</sup> Irreversible potential.



We have previously shown that hydride affinities of both carbenium ions<sup>7</sup> and free radicals<sup>8</sup> correlate linearly with the corresponding electron affinities in solution. Linear regression analysis of the data in Table I revealed an excellent correlation between  $-\Delta G_{hydride}(Q)_{DMSO}$  and  $FE^{\circ}_{NHE}(Q/Q^-)$  which resulted in eq 7. The data are

$$-\Delta G_{hydride}(Q)_{DMSO} = 1.428FE^{\circ}_{NHE}(Q/Q^-)_{DMSO} + 79.5$$

$$r = 0.9974 \quad (7)$$

plotted in Figure 1 in which the electrode potentials are converted to solution electron affinities by adding 4.44 V, the "absolute" potential of the NHE.<sup>17</sup> The data points for *o*-BQ and *p*-BQ fall very close to the regression line.

The fact that  $\Delta G_{hydride}(Q)_{DMSO}$  is linearly related to the solution electron affinities while the  $pK_a(QH_2)_{DMSO}$ , at least for *o*-BQ and *p*-BQ, have the opposite relation to  $E^{\circ}_{NHE}(Q/Q^-)_{DMSO}$  would appear to provide an argument that the stability imparted by hydrogen bonding to the structure of *o*-BQH<sup>-</sup> has little influence upon the magnitude of  $\Delta G_{hydride}(o\text{-BQ})_{DMSO}$ . More generally this implies that electronic effects of substituents in the neutral quinones must play a major role in the relative values of the hydride affinities.

**A Comparison of Hydride Affinities of Quinones with Those of Carbenium Ions and Free Radicals in Solution.** The hydride affinities of quinones studied in this work cover a range from -58 kcal/mol to -101 kcal/mol. These values fall midway between those found for carbenium ions<sup>7</sup> and for arylmethyl radicals.<sup>8</sup> The radicals have very short lifetimes in solution and are thus out of consideration as hydride acceptors in a practical sense. Carbenium ions, on the other hand, are generally better H<sup>-</sup> acceptors than the quinones, with values of  $\Delta G_{hydride}(R^+)_{s}$  ranging from -76 kcal/mol for the most stable ion (crystal violet cation) to -129 kcal/mol for the least stable carbenium ion studied so far, the *p*-nitrobenzyl cation. It is of interest to note that the DDQ molecule, the best H<sup>-</sup> acceptor in the quinone series, is also a better hydride acceptor by 5-25 kcal/mol than all of the stable carbenium ions reported previously,<sup>7</sup> including the trityl cation, a frequently used reagent for organic syntheses involving hydride transfer reactions.<sup>18</sup>

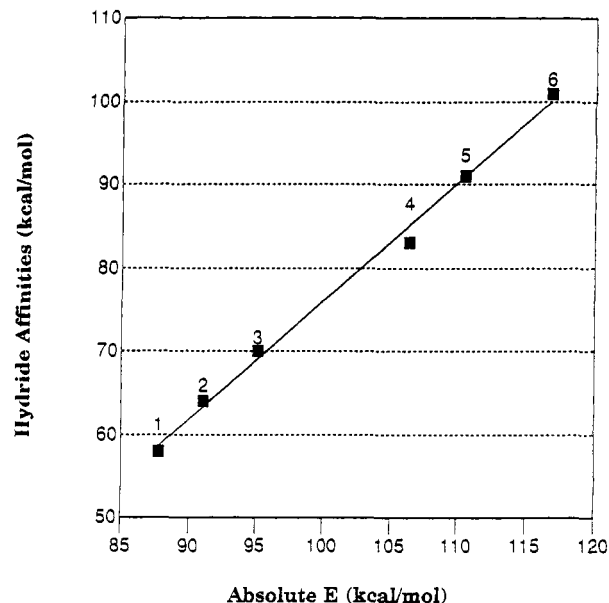
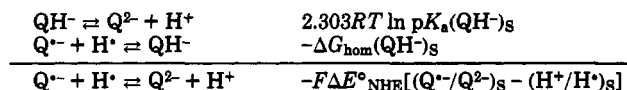


Figure 1. Plot of free energy hydride affinities of quinones vs electron affinities in dimethyl sulfoxide solution: duroquinone (1), *o*-benzoquinone (2), *p*-benzoquinone (3), tetrachloro-*o*-benzoquinone (4), tetrachloro-*p*-benzoquinone (5), and dichlorodicyano-*p*-benzoquinone (6).

**A Comparison of Heterolytic and Homolytic O-H Bond Dissociation Energies of QH<sup>-</sup>.** The appropriate thermochemical cycle to obtain the difference in heterolytic and homolytic bond dissociation free energies in solution is shown in Scheme III.<sup>13</sup> When the solvent is DMSO, Scheme III results in eq 8, in which the last term is equal

Scheme III  
free energy changes



$$\Delta G_{hom}(QH^-)_{DMSO} = 2.303RT \ln pK_a(QH^-)_{DMSO} + FE^{\circ}_{NHE}(Q^{\cdot-}/Q^{2-})_{DMSO} + 57.2 \quad (8)$$

to  $-23.06E^{\circ}_{NHE}(H^{\cdot}/H^+)_{DMSO}$ , for the estimation of  $\Delta G_{hom}(QH^-)_{DMSO}$ . The heterolytic O-H bond dissociation energies of all QH<sup>-</sup> along with the  $pK_a(QH^-)_{DMSO}$  and  $E^{\circ}_{NHE}(Q^{\cdot-}/Q^{2-})_{DMSO}$  necessary for use in eq 8 are summarized in Table III. The differences in heterolytic and homolytic O-H bond dissociation free energies for most of the quinones are very similar in magnitude. The exceptions to this generality are data for QH<sup>-</sup> derived from quinones substituted with either strongly electron-with-

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Table II. Homolytic O-H Bond Dissociation Free Energies of Hydroquinone Anions in DMSO Solution<sup>a</sup>

quinone	p <i>K</i> <sub>a</sub> (QH <sup>-</sup> )	<i>E</i> <sup>o</sup> <sub>NHE</sub> (Q <sup>•-</sup> /Q <sup>2-</sup> )	Δ <i>G</i> <sub>hom</sub> (QH <sup>-</sup> ) <sup>b</sup>	-Δ <i>G</i> <sub>hydride</sub> (QH <sup>-</sup> ) <sup>b</sup>
<i>p</i> -BQ	26.2	-1.208	65.6	70
<i>o</i> -BQ	27.2	-1.36	62.9	64
tetrachloro- <i>p</i> -BQ	17.6	-0.645	66.3	83
tetrachloro- <i>o</i> -BQ	18.9	-0.557	70.1	91
tetramethyl- <i>p</i> -BQ	27.1	-1.498	59.6	58
DDQ	13.5	-0.048	74.5	101

<sup>a</sup> Electrode potential and p*K*<sub>a</sub>(QH<sup>-</sup>) data from Table I. <sup>b</sup> In kcal/mol.

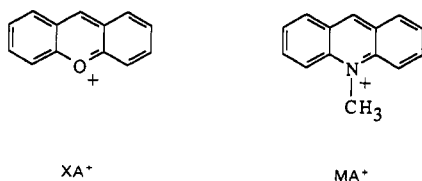
Table III. Thermochemical Data for the Hydride Transfer Reactions of Quinones with 10-Methylacridin in Dimethyl Sulfoxide Solution<sup>a</sup>

quinone	-2.303 <i>RT</i> log <i>K</i> (BQ + MAH) <sup>b</sup>	ΔΔ <i>G</i> <sub>hom</sub> (MAH <sup>•+</sup> - QH <sup>-</sup> ) <sup>c</sup>	- <i>FΔE</i> <sup>o</sup> <sup>d</sup>
<i>p</i> -BQ	7	-25	31
<i>o</i> -BQ	13	-22	35
tetrachloro- <i>p</i> -BQ	-6	-25	20
tetrachloro- <i>o</i> -BQ	-14	-29	16
tetramethyl- <i>p</i> -BQ	19	-19	38
DDQ	-24	-34	10

<sup>a</sup> Free energy changes in kcal/mol. <sup>b</sup> Δ*G*<sub>hydride</sub>(MA<sup>+</sup>) equal to -77 kcal/mol. <sup>c</sup> Δ*G*<sub>hom</sub>(MAH<sup>•+</sup>) equal to 40.7 kcal/mol (see discussion in the text). <sup>d</sup> The difference in electrode potentials of the (BQ/BQ<sup>•-</sup>) and (MAH<sup>•+</sup>/MAH) couples.

drawing or electron-donating substituents. The greatest discrepancy is observed in the data for DDQ.

**Equilibrium Constants for Hydride Transfer Reactions between Quinones and 10-Methylacridin.** Returning again to the thermochemical relationship in Scheme I and the free energy changes involved in the one- and two-step mechanisms of hydride transfer, we note that *K*<sub>hydride</sub> can readily be calculated from the differences in Δ*G*<sub>hydride</sub>(R)<sub>DMSO</sub> of the quinones and of that for 10-methylacridinium ion (MA<sup>+</sup>). In DMSO solution we



observe the potentials of the (MA<sup>+</sup>/MA<sup>•</sup>) and the (MAH<sup>•+</sup>/MAH) couples to be equal to -0.304 and 1.034 V, respectively. Taking XA<sup>+</sup> as a structurally similar model for MA<sup>+</sup>, Δ*G*<sub>hydride</sub>(MA<sup>+</sup>)<sub>DMSO</sub> can be estimated from the difference in potentials for the (XA<sup>+</sup>/XA<sup>•</sup>) and (MA<sup>+</sup>/MA<sup>•</sup>) couples (0.548 V) and Δ*G*<sub>hydride</sub>(XA<sup>+</sup>)<sub>DMSO</sub> (-90 kcal/mol<sup>7</sup>) to be -77 kcal/mol. The homolytic O-H bond dissociation energies for QH<sup>-</sup>, Δ*G*<sub>hom</sub>(QH<sup>-</sup>)<sub>DMSO</sub>, are tabulated in Table II. The only missing quantity to complete the data for Scheme I is Δ*G*<sub>hom</sub>(MAH<sup>•+</sup>)<sub>DMSO</sub>.

The thermochemical cycle shown in Scheme IV was used to derive eq 9. The electrode potential difference below the single line was observed to be equal to 1.34 V. The

Scheme IV

	free energy changes
MAH <sup>•+</sup> ⇌ MA <sup>+</sup> + H <sup>•</sup>	Δ <i>G</i> <sub>hom</sub> (MAH <sup>•+</sup> ) <sub>DMSO</sub>
MA <sup>+</sup> + H <sup>•</sup> ⇌ MAH	-Δ <i>G</i> <sub>hom</sub> (MAH) <sub>DMSO</sub>
MAH <sup>•+</sup> + MA <sup>•</sup> ⇌ MA <sup>+</sup> + MAH	- <i>FΔE</i> [(MAH <sup>•+</sup> /MAH) - (MA <sup>+</sup> /MA <sup>•</sup> )]

$$\Delta G_{\text{hom}}(\text{MAH}^{\bullet+})_{\text{DMSO}} = \Delta G_{\text{hom}}(\text{MAH})_{\text{DMSO}} - F\Delta E[(\text{MAH}^{\bullet+}/\text{MAH}) - (\text{MA}^+/\text{MA}^{\bullet})]_{\text{DMSO}} \quad (9)$$

homolytic C-H bond dissociation free energy, Δ*G*<sub>hom</sub>(MAH)<sub>DMSO</sub>, was estimated to be equal to 71.6 kcal/mol by assuming the same value as that for the corresponding C-H bond of a related model compound, xanthene (XAH).<sup>19</sup> Having the necessary data in hand allowed Δ*G*<sub>hom</sub>(MAH<sup>•+</sup>)<sub>DMSO</sub> to be estimated at 40.7 kcal/mol using eq 9. Equilibrium constants for hydride transfer between BQ and MAH as well as ΔΔ*G*<sub>hom</sub>(MAH<sup>•+</sup> - BQH<sup>-</sup>) and *FΔE*<sup>o</sup>[(BQ/BQ<sup>•-</sup>) - (MAH<sup>•+</sup>/MAH)]<sub>DMSO</sub> are summarized in Table III.



According to Scheme I, the sums of the last two columns in Table III are expected to be equal to the values in the first data column. That this is found to be very nearly the case provides support for the assignment of values to both Δ*G*<sub>hom</sub>(MAH)<sub>DMSO</sub> and Δ*G*<sub>hydride</sub>(MA<sup>+</sup>)<sub>DMSO</sub>. A further check on these assignments is to use eq 3 to calculate Δ*G*<sub>hydride</sub>(MA<sup>+</sup>)<sub>DMSO</sub>. This results in a value of -77.3 assuming Δ*G*<sub>hom</sub>(MAH)<sub>DMSO</sub> to be equal to 71.6 kcal/mol. We have previously estimated Δ*G*<sub>hydride</sub>(MA<sup>+</sup>)<sub>MeCN</sub> at -83 ± 2 kcal/mol. The factor which contributes most to this difference is the potential of the (MA<sup>+</sup>/MA<sup>•</sup>) couple in DMSO (this study, -0.30 V) and in acetonitrile (-0.17 V<sup>20</sup>).

The initial electron transfer equilibria in the two-step mechanism are energetically highly unfavorable for the reaction of all of the quinones with MAH. On the other hand, the hydrogen atom transfer steps between QH<sup>•-</sup> and MAH<sup>•+</sup> are exergonic with free energy changes ranging from -19 to -34 kcal/mol. Although the one-step mechanism<sup>10</sup> is favored for the reactions of all of the quinones in Table IV with 10-methylacridin in DMSO, only those quinones substituted with electron-withdrawing groups are expected to undergo rapid reactions.

## Experimental Section

NMR spectra were obtained on a Varian XL 300 or a JEOL 270 spectrometer. Mass spectra were recorded using a Finnigan MAT ITD GC/MS system. Gas chromatography was performed on a Varian 3700 instrument equipped with a flame ionization detector. The purification of DMSO and the prep-

(19) The homolytic C-H bond dissociation free energy in DMSO at the 9-position of xanthene was calculated to be equal to 71.6 kcal/mol using p*K*<sub>a</sub> and electrode potential data found in ref 7. Relative bond dissociation energies of MAH and other NADH model compounds have been determined in acetonitrile solution.<sup>18</sup>

(20) Klippenstein, J.; Aray, P.; Wayner, D. D. M. *J. Org. Chem.* 1991, 56, 6736.

aration of the dimethyl base were carried out according to a literature procedure.<sup>12a</sup>

**Materials.** Tetrabutylammonium hexafluorophosphate (Aldrich) was recrystallized from  $\text{CH}_2\text{Cl}_2/\text{Et}_2\text{O}$  and vacuum-dried at 110 °C for 10 h before use. *p*-Benzoquinone, *p*-hydroquinone, catechol, tetrachloro-*p*-benzoquinone, tetrachloro-*o*-benzoquinone, DDQ, and duroquinone obtained commercially were recrystallized before use. The *o*-benzoquinone was the product of exhaustive electrolytic oxidation of catechol in the presence of a base and was used *in situ*. All other compounds were prepared using published or modifications of published procedures. The purity of all compounds was verified by melting point determination and spectral analysis.

**Measurements of the Equilibrium Constants of the Second Proton Transfer from Hydroquinones. The  $\text{p}K_a(\text{QH}^-)$ .** The Bordwell indicator overlapping method,<sup>12</sup> which has been applied to the measurements of the first ionization constants for a very extensive collection of weak organic acids, was modified in order to obtain the  $\text{p}K_a$ 's for the second proton transfer from hydroquinones. This modification was based on the assumption that since the second proton transfer is at least four to five powers of 10 times more difficult than the first proton transfer then when an indicator with a  $\text{p}K_a$  close to that of the conjugate acid of the hydroquinone dianion is selected, the generation of the hydroquinone monoanion is quantitative. The experimental procedure of the  $\text{p}K_a$  measurement was basically the same as described in the literature<sup>12</sup> for the determination of the first ionization constant. The data manipulation was as follows.

Equation 10 describes the acid/base equilibrium for the second proton transfer, where  $\text{QH}^-$  is the hydroquinone monoanion,  $\text{In}^-$



is the conjugate base of the indicator HIn which is selected in such a way that its acidity should be within 1.5 p*K* units of the expected value for the  $\text{QH}^-$ . The initial concentration of  $\text{QH}^-$  was assumed to be the same as the amount of hydroquinone  $\text{QH}_2$  titrated into the UV cell which was determined gravimetrically. The initial concentration of  $\text{In}^-$  was obtained from the final optical

absorbance of  $\text{In}^-$  based on Beer's Law which also gave an extinction coefficient ( $\epsilon$ ) of the indicator anion. Following the acid titration, the  $[\text{In}^-]$  at equilibrium was arrived at from the newly recorded optical absorbance and the  $\epsilon$ . The equilibrium concentrations of the other three species in solution were calculated accordingly. These equilibrium concentrations allow the equilibrium constant  $K_{11}$  to be evaluated for use in eq 11 to

$$\text{p}K_a(\text{QH}^-) = \text{p}K_a(\text{HIn}) - \log K_{10} \quad (11)$$

derive the  $\text{p}K_a(\text{QH}^-)$ . The  $\text{p}K_a$  data in Table I are the average values of two or three determinations or from two indicators and are believed to be accurate to  $\pm 0.2$ .

**Electrochemical Measurements.** A standard three-electrode cell equipped with a Teflon top was used for electrode potential measurements. The working electrode was a 0.2-mm diameter Pt disk and was reconditioned by polishing with a 0.05- $\mu\text{m}$  alumina suspension (Stuers, Denmark) between runs. The potentials were measured at a sweep rate of 1 V/s, vs a 0.1 M  $\text{AgClO}_4/\text{Ag}$  reference electrode with the ferrocenium/ferrocene redox couple as the internal standard. The counter electrode was a Pt wire.

**Instrumentation and Data Handling Procedures.** Cyclic voltammetry was performed using a JAS Instrument Systems, Model JDP-165A, potentiostat driven by a Hewlett-Packard 3314A function generator. The signals were filtered with a Stanford Research Systems, Inc., Model SR640 dual channel low pass filter before recording on a Nicolet Model 310 digital oscilloscope. An IBM AT compatible personal computer was used to control the oscilloscope and the function generator via an IEEE interface. The current-potential curves were collected and averaged at trigger intervals selected to reduce periodic noise.<sup>21</sup> The averaged signals were then treated with a digital frequency domain low pass filter before numerical differentiation.

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(21) Lason, E.; Parker, V. D. *Anal. Chem.* 1990, 62, 412.