laser, typically 10^{-4} cm^{-1}. The combination of diode lasers with inert rare gas waveguides then represents 4 orders of magnitude improvement in resolution over existing surface sensitive spectroscopic techniques.

It is of interest to also note that in the lowest order guided modes very high field densities are attained in the slab and at the metal/dielectric interface in the case of TM modes. The total injected power in these experiments was typically between 5 and 20 μW. As much as 70% of this field is contained in the TM_0 mode where the thickness of the slab is $\sim$3 μm. For a 50-μm beam waist, this fluence then corresponds to 6 W/cm^2. The effect of this high-field density in our present experiments is that of line broadening. These high field densities, coupled with the narrow radiation bandwidth, represent a formidable spectral brightness that could be used in nonlinear spectroscopic applications.[11]

These demonstrations were performed in a first-generation design of a waveguiding assembly. There are obvious improvements that can be implemented for enhancement of sensitivity and mode selectivity. Prism couplers[19] or a grating etched in the copper substrate could be used for the injection of radiation at better defined angles and for improved efficiency in output coupling. In order to sample a larger number of molecules either in the slab or at the slab/metal interface, line focusing as opposed to the present point focusing geometry could be used. These improvements, and the incorporation of the assembly in a ultrahigh-vacuum chamber, are presently being implemented in order to apply the technique to submonolayer metal surface adsorbates.

Acknowledgment. Acknowledgment is made to the donors of the Petroleum Research Fund, administered by the American Chemical Society, for partial support of this research. The award from the Camille and Henry Dreyfus Foundation for "Newly Appointed Faculty in Chemistry" is also gratefully acknowledged. The diode lasers were obtained through a grant from the NSF Instrumentation program.

Registry No. Cu, 7440-50-8; Ar, 7440-37-1; CH_3F, 593-53-3.

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Reduction Potentials and Exchange Reactions of Thiyl Radicals and Disulfide Anion Radicals

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Redox equilibria between RS$^\bullet$ and $-S\therefore S-^\bullet$ radicals, and between these types of radical and phenoxyl and chlorpromazine ($ClPz^{\bullet 2+}$) radicals, have been investigated in aqueous solutions at pHs over the range 6-10 to obtain a self-consistent set of redox potentials for the reactions $PhO^\bullet + H^+ + e^- = PhOH$ (4), $RS\therefore SR^- + 2H^+ + e^- = 2RSH$ (11), and $RS^\bullet + H^+ + e^- = RSH$ (18), in sulfur systems with alkyl R groups. Absolute standard potentials were calculated on the basis of E° = 0.83 V for the chlorpromazine couple. The results for E°_4 (=1.35 $\pm$ 0.02 V) and E°_{18} (=1.33 $\pm$ 0.02 V) were in agreement with values calculated from thermodynamic data within the known uncertainties. E°_{18} was found to exhibit a falloff when electron-rich groups, such as the two methyls of penicillamine or the CO_2^- of β-mercaptoacetic acid, were present on the carbon adjacent to the S atom. However, the effect was relatively small ($\sim$10-14 mV). E°_{11} was 1.72 $\pm$ 0.02 V for β-mercaptoethanol. The corresponding potentials for the cyclic anions of dithiothreitol, dithioerythritol, and lipoamide were the same within experimental error, but the uncertainties were larger ($\pm$0.04 V). For the reaction $e^- + -S-S- = -S\therefore S-^\bullet$ (22), the magnitude of E°_{22} was calculated to be -1.60 V, showing that only strongly reducing species could donate electrons to disulfide. Rate constants for several of the forward and backward reactions in the equilibria were also determined.

Introduction

Sulfhydryl radicals (RS$^\bullet$ species) are of interest in the redox chemistry of inorganic[1,2] and bioinorganic[3,4] sulfur-containing systems. Also, these radicals and the disulfide anions ($RS\therefore SR^-$) formed on their complexation by thiolate anion are involved in protection mechanisms for biological systems subjected to ionizing radiation or other forms of free radical damage.[5,6] In all of these systems a knowledge of the redox potentials is of great value in assessing various mechanisms of reaction and estimating reaction rates.

Recently we reported thermodynamic calculations[7] that gave values of E° for the couples $RS^\bullet/RS^-$, $RS\therefore SR^-/2RS^-$, and $RSSR/RS\therefore SR^-$. With these as a guide, experiments were designed to equilibrate different radical systems with each other and with couples of known reduction potential, so that more exact experimental E° values could be obtained. This paper reports a study of the equilibria between several sulfur-containing radicals, chlorpromazine, and phenol. In addition, the effects of structure of the aliphatic R groups of various thiyl or sulfhydryl radicals have been examined. The technique employed was pulse radiolysis. The reactions producing the radicals are well documented,[8,9] and the procedures were the same as those described in ref 10 and 11.

As in many studies of this kind all solvated electrons (e^-_{aq}) were converted to $^\bullet OH$ radicals by reaction with N_2O.[8,9] Concentrations of solutes were adjusted so that those radicals reacted with either a sulfhydryl molecule (RSH) or phenol (PhOH) on a 10^{-7}-s time scale. The thiyl or phenoxyl ($PhO^\bullet$) radicals, which are respectively produced in the two cases, then initiated the exchange

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reactions, which occurred with other solutes on a 10–50-μs time scale.

Experimental Section

Materials. β-Mercaptoethanol (β-Me), mercaptoacetic acid (MAA), dithiothreitol ($D(SH)_2$ or DTT), dithioerythritol (DTE), chlorpromazine ($ClPz^+$) and penicillamine (PenSH) were purchased from Sigma Chemical Co. β-Mercaptopropionic acid (β-MPA) was from Fluka, and phenol was from Fisher Scientific Co. Dihydrolipoamide ($L(SH)_2$) was prepared from lipoamide (LS_2) by the method of Reed et al.[12] All the other chemicals were of the purest grade available and used as supplied. The water for preparing solutions was redistilled, first from alkaline potassium permanganate and then from acid dichromate.[8]

Apparatus and Procedures. All experiments were conducted at room temperature (23 ± 2 °C) with solutions containing phosphate buffer which was 10 mM unless otherwise stated. In order to prevent their oxidation by air, solutions with sulfhydryl compounds were degassed during preparation by prebubbling with nitrous oxide, which had been passed through a column of Ridox (Fisher Scientific Co.). The pH was finally adjusted with $HClO_4$ or NaOH, and the solutions were saturated with nitrous oxide before pulsing at 296 ± 2 K.

The pulse radiolysis experiments were performed by utilizing 2-μs pulses of electrons with doses of 4–10 Gy/pulse from a 1.5-MeV Van de Graaff generator. Details of the generator and dosimetry have been described earlier.[11,13]

Results

Experimental data are presented for four types of radical exchange reactions, and thermodynamic calculations of E° for the phenoxyl/phenol couple are given. These results are presented in sections 1–5.

1. Reaction between Phenoxyl Radicals and Disulfhydryls. This is exemplified for dithiothreitol:[14]

$$PhO^\bullet + D(SH)_2 = PhOH + DS_2^{\bullet-} + H^+ \qquad (1)$$

The equilibrium expression is

$$K_1 = \frac{[PhOH][DS_2^{\bullet-}][H^+]}{[D(SH)_2][PhO^\bullet]} \qquad (2)$$

in which the concentrations of PhOH, H^+, and $D(SH)_2$ are essentially constant due to the small number of radicals formed and the buffer action on $[H^+]$. The kinetics and equilibrium concentrations for this reaction were studied at pHs 7, 8, 9, and 10 in 0.1 M phenol solutions with 10 mM phosphate buffer. The initial absorbance directly after the pulse was due to the radicals produced from the attack of OH on phenol and due to the proportion of disulfide radicals formed by partioning of hydroxyl radicals between phenol and dithiothreitol.[15] With the concentrations used, the $DS_2^{\bullet-}$ absorbance increased after the pulse as a result of reaction 1 in a pseudo-first-order manner. Figure 1b shows the plot of the pseudo-first-order rate constants k_1 versus dithiothreitol concentration observed at pH 8. The $DS_2^{\bullet-}$ concentration was monitored at 480 nm, where the phenoxyl radical absorbance can be neglected.[16,17] In accord with the mathematical expressions for reversible reactions,[18] the slope of the linear plot

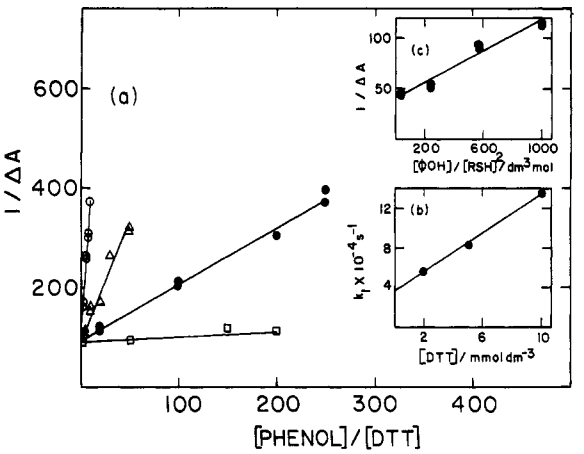

Figure 1. (a) Plot of $1/\Delta A$ at 480 nm versus ratio of [phenol] to [DTT] at pH 7 (○), 8 (△), 9 (●), and 10 (□). (b) Plot of pseudo-first-order rate constants for absorbance growth at 480 nm versus [DTT] at pH 9. (c) Plot of $1/\Delta A$ versus ratio of [phenol] to $[RSH]^2$ at pH 9.

corresponds to the second-order rate constant for the reaction in the forward direction. The intercept corresponds to the pseudo-first-order rate constant k_{-1} for the back reaction, in this case a complex third-order rate constant k_b, multiplied by $[H^+]$ and [PhOH]. The effective equilibrium quotient at pH 8 calculated from the ratio of the forward and reverse rate constants was 1.8×10^{-7}.

Starting with the equilibrium expression for reaction 1, one can derive an equation for the dependence of the absorbance of $DS_2^{\bullet-}$ anion on the ratio of the concentrations of phenol to dithiothreitol:

$$\frac{1}{\Delta A} = \frac{1}{\Delta A_\infty} + \frac{[PhOH][H^+]}{[D(SH)_2]K_1\Delta A_\infty} \qquad (3)$$

The quantity ΔA_∞ corresponds to the absorbance of $DS_2^{\bullet-}$ in the absence of phenol, when it is the only species formed. Figure 1a shows plots of $1/\Delta A$ versus the ratio of phenol to dithiothreitol concentration at pH 7, 8, 9, and 10. The equilibrium quotients, were 2.4×10^{-7}, 2.0×10^{-7}, 6.3×10^{-8}, and 6.4×10^{-8} for pH 7, 8, 9, and 10, respectively. The value for pH 8 is in good agreement with the result obtained from the kinetic data in Figure 1b.

The pH dependence for the free energy change in the half-reaction

$$PhO^\bullet + e^- + H^+ = PhOH \qquad (4)$$

is

$$\Delta G_4 = \Delta G^\circ_4 + RT \ln \frac{[PhOH]}{[PhO^\bullet]\{[H^+] + K_{iPhOH}\}} \qquad (5)$$

where K_{iPhOH} represents the ionization constant of phenol ($=10^{-10.0}$).[19,20] The corresponding expression for the half-reaction of $DS_2^{\bullet-}$

$$DS_2^{\bullet-} + 2H^+ + e^- = D(SH)_2 \qquad (6)$$

is

$$\Delta G_6 = \Delta G^\circ_6 + RT \ln \frac{[D(SH)_2]}{[H^+][DS_2^{\bullet-}]} - RT \ln \left\{[H^+] + K_a + \frac{K_aK_b}{[H^+]}\right\} \qquad (7)$$

The constant K_a refers to the ionization constant of the first SH group ($pK_a = 9.26$)[21] and K_b to the ionization constant of the

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second (pK_b = 10.34).[21] Combination of expressions 7 and 5 for the equilibrium condition $\Delta G_4 = \Delta G_6$ leads to

$$-\Delta G^\circ_1 = RT \ln K_1 - RT \ln \{[H^+] + K_{iPhOH}\} + RT \ln \left\{[H^+] + K_a + \frac{K_a K_b}{[H^+]}\right\} \quad (8)$$

Here it should be emphasized that [PhOH] and [$D(SH)_2$] refer to the total concentrations of ionized and un-ionized molecules. This formalism applies throughout the present work. Using expression (8) and correcting K_1 to zero ionic strength (see below), ΔG°_1 was found to be 37 280, 38 100, 40 300, and 38 600 J mol^{-1} from the data at pH 7, 8, 9, and 10, respectively, with an average value of 38 600 ± 1000 J mol^{-1}.

The equilibrium for the reaction between phenoxyl radicals and the disulfhydryl lipoamide was also investigated:

$$PhO^\bullet + L(SH)_2 = PhOH + LS_2^{\bullet -} + H^+ \quad (9)$$

Since the $L(SH)_2$ system is completely analgous to $D(SH)_2$, the half-reaction is similar to reaction 6 and the expression for the overall free energy change is similar to expression 8.

Experiments were performed with $L(SH)_2$ only at pH 9. The value of K_1 at that pH was 3.2×10^{-8}. The average of K_a and K_b is 10.7, and the value used for K_a was 10.4.[22,23] Substituting these and the above value of K_1 into expression 8, ΔG°_9 was found to be 43 100 J mol^{-1}.

2. Reaction between Phenoxyl Radicals and RSH. The equilibrium conditions for the reaction between phenoxyl radical and RSH molecules was studied at pH 9. The reaction involved is

$$PhO^\bullet + 2RSH = PhOH + RS\dot{S}R^- + H^+ \quad (10)$$

Again, the absorbance of $RS\dot{S}R^-$ was followed at 480 nm. As in all cases reported in this study, the equilibrium quotients were determined from the changes in absorbance measured when the growth of the anion was complete and before significant decay occurred. In this instance the equilibrium expression involves the square of RSH concentration. A plot of $1/\Delta A$ versus phenol concentration divided by the square of RSH concentration is shown in Figure 1c.

The half-reaction for the reduction of $RS\dot{S}R^-$ is

$$RS\dot{S}R^- + e^- + 2H^+ = 2RSH \quad (11)$$

The expression for the free energy change in reaction 10 is

$$-\Delta G^\circ_{10} = RT \ln K_{10} - RT \ln \{[H^+] + K_{iPhOH}\} + RT \ln \left\{[H^+] + 2K_c + \frac{K_c^2}{[H^+]}\right\} \quad (12)$$

where the parameter K_c is the ionization constant for the RSH molecule, which in the case of Figure 1c was β-mercaptoethanol with a pK of 9.7.[21] The magnitude of K_{10} determined from the linear plot in Figure 1c was 3.9×10^{-7}, from which with the aid of eq 12 we calculated ΔG°_{10} = 36 100 J mol^{-1}, after correction to zero ionic strength.

3. Reactions between RS• Radicals of Different Structure and Dithiothreitol. Several reactions of this kind, represented by eq 13, were studied at pH 7:

$$RS^\bullet + D(SH)_2 = RSH + DS_2^{\bullet -} + H^+ \quad (13)$$

The expression relating the absorbance of $DS_2^{\bullet -}$ anion to the equilibrium constant and the ratio of RSH to the $D(SH)_2$ concentrations can be shown to be entirely analogous to expression 3, with the concentration of RSH replacing the concentration of phenol. Some of the RSH systems used have appreciable absorbances of their RS• and $RS\dot{S}R^-$ anion radicals at pH 7,[24] and checks were made to ensure that these did not contribute significantly to the total absorption at 400 nm, where $DS_2^{\bullet -}$ was monitored. In addition, changing the absorption measurement to 430 nm, where RS• does not absorb, did not alter the magnitude of K_{13} calculated.

TABLE I: Free Energy Changes and Rate Coefficients[a] at 296 K

equilibrated reaction	ΔG°/J	pH	$k_f/10^7$ M^{-1} s^{-1}	$k_b/10^{13}$ M^{-2} s^{-1}
(1)	38 600[b]	8.0	1.0	5.0
		9.0	4.5	53.0
(9)	43 100	9.0		
(10) with RSH = β-ME	36 100	9.0		
	35 300	10.0		
(13) for DTT form of $D(SH)_2$ with				
RSH = MAA	36 700	7.0	1.5	6.0
RSH = β-MPA	35 500	7.0	0.9	2.4
RSH = β-ME	35 400	7.0	1.2	1.8
RSH = PenSH	36 900	7.0		
(13) for DTE form of $D(SH)_2$ with RSH = β-MPA	35 200	7.0	0.9	2.2
(13) for $L(SH)_2$ with RSH = β-ME	36 100	7.0	3.3	10.0
(15)	−48 050	6.0	17.5	
(14)	−50 330	6.0		

[a] Subject to uncertainties of ±1000 J in ΔG° and ±15% in k_f and k_b. [b] Average of pH 7, 8, 9, and 10.

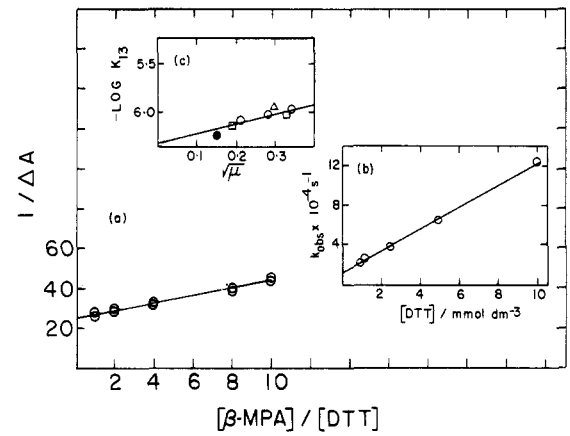

Figure 2. (a) Plot of $1/\Delta A$ versus ratio of [β-MPA] to [DTT] in 60 mM phosphate buffer at pH 7. (b) Plot of pseudo-first-order rate constants measured from absorbance changes at 430 nm versus [DTT]. (c) Plot of $-\log_{10} K_{13}$ versus $\sqrt{\mu}$: [β-MPA] = (●) 0.005, (○) 0.010, (□) 0.020, and (△) 0.010 M plus 0.071 M $NaClO_4$.

Equilibrium quotients were determined for the RSH molecules listed in Table I. A typical example of a plot of $1/\Delta A$ against the ratio of [RSH] to [$D(SH)_2$] is shown in Figure 2a. In several instances the kinetics of the approach to equilibrium were also followed. A plot of the pseudo-first-order rate constant for the growth of the $DS_2^{\bullet -}$ absorbance against dithiothreitol concentration is shown for β-mercaptopropionic acid in Figure 2b. The equilibrium quotient determined from the ratio of the slope of the line to the intercept divided by the RSH molecule concentration was 3.9×10^{-7}. This is to be compared with the value of 5.9×10^{-7} from the $1/\Delta A$ plot shown in Figure 2a.

For the case of β-mercaptopropionic acid, the effects of concentration of RSH and of buffer were also explored. Since the carboxylate group of this RSH molecule is ionized at pH 7, it and the buffer alter the ionic strength (μ) of the medium as their concentrations are changed. The results are therefore summarized in terms of a plot of log K_{13} versus $\sqrt{\mu}$ in Figure 2c. The line through the points has a slope of 1.02, which is the theoretical,

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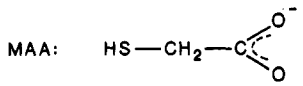
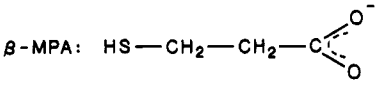
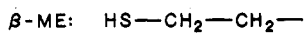
TABLE II: $\Delta E°$ Values Relative to ClPz Half-Reaction

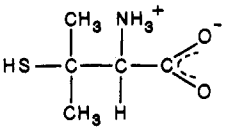
half-reaction	$\Delta E°$	conditions of measurement: pH	conditions of measurement: comparison couple[a]
(6) $DS_2^{\bullet-} + 2H^+ + e^- = D(SH)_2$	0.92_2	7–9	a
	0.87_2	7.0	b
(6) $LS_2^{\bullet-} + 2H^+ + e^- = L(SH)_2$	0.96_9	9.0	a
	0.87_9	7.0	b
(11) $RS\dot{S}R^- + 2H^+ + e^- = 2RSH$ (=β-ME)	0.89_6	9.0	a
	0.88_8	10.0	
(4) $PhO^\bullet + e^- + H^+ = PhOH$	0.52_2	6.0	c
	0.51[b]	6.0	c
(18) $RS^\bullet + e^- + H^+ = RSH$			
RSH = MAA	0.51	7.0	d
RSH = β-MPA	0.52	7.0	d
RSH = β-ME	0.49_8	6.0	c
	0.51_2[c]	9.0	a
RSH = PenSH	0.51	7.0	d

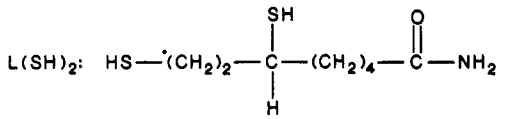
[a] a = $PhO^\bullet/PhOH$; b = $RS^\bullet/RSH$; c = $ClPz^{\bullet+}/ClPz$; d = $DS_2^{\bullet-}/D(SH)_2$. [b] From ref 29 (see text). [c] Via $RS\dot{S}R^-$ (see text).

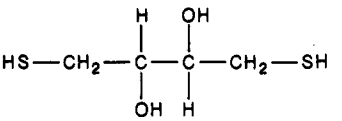
MAA: 

β-MPA: 

β-ME: 

PenSH: 

$L(SH)_2$: 

$D(SH)_2$: 

based on the Debye–Hückel limiting law as applied to the ionic species in reaction 13,[25] viz.

$$\log K°_{13} = \log K_{13} + \log [\gamma RSH\, \gamma DS_2^{\bullet-} \gamma H^+ / (\gamma RS^\bullet\, \gamma D(SH)_2)]$$

Since γRSH and $\gamma RS^\bullet$ are identical within that theoretical frame work and γ for the uncharged $D(SH)_2$ species can be taken as unity, this simplifies to

$$\log K_{13} = \log K°_{13} + 2(0.509\sqrt{\mu})$$

Within experimental error, the data are seen to conform to this treatment. The buffer concentration was varied from 10 to 60 mM, and [RSH] from 0.50 to 2.0 mM. The demonstration of an influence of ionic strength led us to correct equilibrium quotients to equilibrium constants at zero ionic strength, using the Debye–Hückel limiting law throughout this work. However, it should be noted that these corrections have only a small effect on the ΔG values.

For the RSH system where kinetics were followed, the equilibrium constants determined by the kinetics and the absorbance techniques were in good agreement. Values of the rate constants for reaction 13 are given in Table I with the free energy changes. It should be noted that in the case of β-mercaptoethanol the exchanges with $L(SH)_2$ and dithioerythritol, the diastereoisomer isomer of dithiothreitol, were also studied.

4. Reactions of $PhO^\bullet$ and $RS^\bullet$ with Chlorpromazine. The exchange reaction between phenoxyl radical and chlorpromazine, viz.

$$PhO^\bullet + ClPz^+ + H^+ = PhOH + ClPz^{\bullet 2+} \quad (14)$$

and the analogous process for $RS^\bullet$ of β-mercaptoethanol

$$RS^\bullet + ClPz^+ + H^+ = RSH + ClPz^{\bullet 2+} \quad (15)$$

were both investigated at pH 6. Data on the redox potential of the half-reaction

$$ClPz^{\bullet 2+} + e^- = ClPz^+ \quad (16)$$

are available (ref 26 and below), and there is evidence that it is independent of pH below pH 7.[27] In the calculations of redox potentials, which follow later, this reaction was therefore used as a standard.

The values of the equilibrium constants for reaction 14 were obtained by measuring the absorbance of the $ClPz^{\bullet 2+}$ radical at 525 nm.[26,28] The expression relating this absorbance to K_{14} and concentrations was

$$\frac{1}{\Delta A} = \frac{1}{\Delta A_\infty} + \frac{[PhOH]}{[ClPz^+][H^+]\, K_{14} \Delta A_\infty} \quad (17)$$

There is an analogous expression for reaction 15. The value of ΔA_∞ and the shape of the $ClPz^{\bullet 2+}$ absorption spectrum formed by the $RS^\bullet$ radicals were compared with the results obtained by oxidation of ClPz by $Br_2^{\bullet-}$ and shown to be the same. Values of $\Delta G°_{14}$ and $\Delta G°_{15}$ are given in Table I. Rates of oxidation at higher chlorpromazine concentrations were generally too fast to be measured accurately on our system and are not reported. The values of k_f and k_b at pH 6 for reaction 15 were 1.8×10^8 and 1.2×10^6 M^{-1} s^{-1}, respectively.

Values of $\Delta E°$, the redox potentials of half-reactions relative to $E°_{16}$, are given in Table II. It is worth noting that our value of $\Delta E°_4$ (=0.52_2) is in good agreement with 0.51 V calculated with the aid of expression 5 above from data at pH 11.1 given by Huie and Neta.[29] This later value was obtained indirectly by first comparing the redox potentials of $ClPz^{\bullet 2+}$ and $^\bullet SO_3^-$ at pH 3.6 and then those of $^\bullet SO_3^-$ and phenoxyl at pH 11.1. One also obtains from $\Delta G°_{15}$ a value of $\Delta E°_{18}$ at pH 6 for the half-reaction

$$RS^\bullet + e^- + H^+ = RSH \quad (18)$$

The result, 0.49_8 V, is discussed further below.

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The magnitude of ΔG°_{10} in Table I was used in conjunction with ΔE°_4 to obtain $\Delta E^\circ_{11} = 0.89_6$ V. The magnitude of ΔE°_{18} can also be obtained from this by adding (after conversion to E° units) ΔG°_{19a} and ΔG°_{19b}, viz.

$$RS^\bullet + RS^- = RS\dot{S}R^- \quad (19a)$$

$$RSH = RS^- + H^+ \quad (19b)$$

The magnitude of K_{19b} is known,[21] and K_{19a} has been determined.[30] We did not reinvestigate this reaction but found that absorbances of $RS\dot{S}R^-$ at 420 nm per 10 Gy of dose were in excellent agreement with those calculated from the equilibrium constant K_{19a} (=(1.7 ± 0.1) × 10^3 M) and the absorbance coefficient of $RS\dot{S}R^-$ (8300 M^{-1} cm^{-1}) given by Karman et al.,[30] viz., at pH 9 for 2 and 5 mM RSH, 0.016 and 0.027 measured versus 0.017 and 0.029 calculated; and at pH 8 for 4 and 10 mM RSH, 0.0055 and 0.0111 measured versus 0.0053 and 0.0114 calculated, respectively. Using ΔG°_{19a} derived from that value of K_{19a} and ΔG°_{19b}, we calculated $\Delta E^\circ_{18} = 0.51_2$ V. The agreement with the magnitude obtained by direct comparison with chlorpromazine (0.49_8 V, above) establishes the self-consistency of the present data.

The average value of ΔE°_{18} for β-ME from the two results above ($0.50_5 \pm 0.00_7$ V) and our own value of ΔE°_4, 0.52_2 V, have been used in conjunction with the ΔG°_1, ΔG°_9, and ΔG°_{13} values in Table I to determine the magnitudes of ΔE°_6 shown in Table II. The mean value of ΔE°_6 for $DS_2^{\bullet-}/D(SH)_2$ from those calculations (0.89_7 V) was then employed to find ΔE°_{18} for the other sulfhydryl molecules β-MPA, MAA, and PenSH.

5. The standard free energy change of the electrode half-reaction 4 can be equated with that of

$$^1/_2H_{2(g)} + PhO^\bullet = PhOH \quad (4')$$

In principle the magnitude of $\Delta G^\circ_{4'}$ can be found from the sums of the standard free energy changes of

$$^1/_2H_{2(g)} = H^\bullet_{(g)} \quad (4a)$$

$$H^\bullet_{(g)} + PhO^\bullet_{(g)} = PhOH_{(g)} \quad (4b)$$

$$PhOH_{(g)} = PhOH \quad (4c)$$

$$PhO^\bullet = PhO^\bullet_{(g)} \quad (4d)$$

The value of ΔH° for reaction 4b is 356 ± 12 kJ mol^{-1},[31] and entropies of the species (S° = 115, 308, and 314 J mol^{-1} K^{-1} for $H^\bullet$, $PhO^\bullet$, and PhOH, respectively from ref 32 and 33) can be found from tables of standard data, as can ΔG°_{4a} (203.3 kJ mol^{-1}, ref 32). The simplest approach to finding the combined free energy of the solution processes 4c and 4d is to assume that phenol behaves in the same way as the aliphatic alcohols, whose solution free energies were treated recently in terms of hydrogen-bond formation by Schwarz and Dodson.[34] Under their treatment $PhO^\bullet$ would have one less hydrogen bond than PhOH. The free energy of each H bond is 9.2 kJ mol^{-1} and $\Delta G^\circ_{4c} + \Delta G^\circ_{4d}$ would correspond to a net change in free energy of solution of −9.2 kJ mol^{-1}. On combination with the standard free energies for reactions 4a and 4b this result yields $\Delta G^\circ_{4'}$ = 129 kJ mol^{-1} or E°_4 = 1.34 V.

Discussion

The free energy changes in Table I and redox potential data in Table II represent the primary focus of this work. However, several points relating to the kinetics of the reactions deserve comment. In the first place, for those cases where RSH was the radical source, one has the possibility of $^\bullet OH$ abstracting H from sites other than SH. However, with the RSH concentrations used here, the alkyl radicals so formed should be converted to the desired thiyl radicals by reaction 1a on a time scale of 1 μs or

$$R^\bullet + R'SH = RH + R'S^\bullet \quad (1a)$$

less. This is because rate constants for typical alkyl radical H abstraction from RSH[35] are on the order of 10^8 M^{-1} s^{-1}. A second potential complication arises in connection with phenoxyl radicals. The initial products of $^\bullet OH$ attack on phenoxyl are in fact cyclohexadienyl OH adducts.[16] These decay to phenoxyl radical in an acid–base-catalyzed process. From the data of Land and Ebert[16] the rate of this decay at pH 7 for the 10 mM phosphate buffer used in this study would have been ≥10^5 s^{-1}, and we confirmed that. Since that rate is comparable to the forward rate for reaction 1, it is possible that some cyclohexadienyl radicals reacted directly with $D(SH)_2$ to form $D(SH)S^\bullet$. From the work of Akhlaq and von Sonntag[36] this species would cyclize and ionise to $DS_2^{\bullet-}$ at a rate of about 10^6 s^{-1} for our buffer concentrations. Formation of an equilibrium concentration of $PhO^\bullet$ could then require reaction of the $D(SH)S^\bullet$ species or $DS_2^{\bullet-}$ with PhOH. The fact that ΔG°_1 from the results at pH 7 was in agreement with values at pHs of 8, 9, and 10, where the rate of conversion of cyclohexadienyl OH adducts to $PhO^\bullet$ was found to be progressively faster, implies that the initial production of $^\bullet OH$ adducts did not prevent equilibrium being reached. The fact that all radical products of $^\bullet OH$ + PhOH were capable of producing $DS_2^{\bullet-}$ was established from the intercept of the plots in Figure 1a in accord with eq 3, which gave a value of ΔA_∞ in agreement with the known absorbance of $DS_2^{\bullet-}$ in the absence of PhOH. The absence of any effect of concentration of phosphate buffer on reaction 13 showed that acid- or base-catalyzed processes, such as the recently studied protonation–deprotonation equilibrium of the cyclic $DS_2^{\bullet-}$ anion,[36] were not inhibiting the approach to equilibrium. Furthermore, the generally good agreement between equilibrium constants from plots of $1/\Delta A_\infty$ versus concentration ratios (as in Figure 1a) and those from the ratios of rate constants for the forward and backward reactions strongly supports the conclusion that equilibria were established. This is also born out by the self-consistancy of the results from comparisons of ΔE° values obtained with $RS^\bullet/RSH$ and $PhO^\bullet/PhOH$ couples in Table II (see further below).

The kinetics of reactions 1, 9, 10, and 13 are complex, since a minimum of three reactant species are involved in the backward direction. One obvious mechanism in that direction would be preliminary protonation of the anion, ring opening to form a thiyl radical, and hydrogen abstraction by it from PhOH for reactions 1, 9, and 10 or from RSH for reaction 13. However, alternatives such as the formation of a complex between the disulfide anion and PhOH (or RSH) cannot be precluded at present.

The pK values of DTT in ref 21 lead to the conclusion that at pHs 7, 8, and 9 $D(SH)S^-$ constitutes 0.6, 5, and 35% respectively of the total DTT present. The major component at these pHs is $D(SH)_2$, and at pH 9 the percentage of S^-DS^- is still only 1.6. The rate-controlling step of the forward reaction 13 cannot be known until the mechanism is understood. However, from the above parameters it is evident that the reactants for the forward step of reaction 13 at pH 7 are $RS^\bullet$ and $D(SH)_2$. The second-order rate constants for (13) in Table I are in accord with those reported by Akhlaq and von Sonntag[36] for similar reactions. They are much smaller than k_{1a} for typical aliphatic radicals.

Because $RS\dot{S}R^-$ anions formed in reaction 19a become important at pH 8 and 9 and cannot be distinguished from $DS_2^{\bullet-}$, reaction 13 could not be studied at those pHs. However, reaction 1 was. Setting up simultaneous equations for the k_f values of this reaction from Table I at pH 8 and 9 with the fractions of $D(SH)_2$ and $D(SH)S^-$ given above, one finds second-order rate constants

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of $(3.4 \pm 0.3) \times 10^6$ and $(1.2 \pm 0.4) \times 10^8$ M^{-1} s^{-1} for the reaction of $PhO^•$ with $D(SH)_2$ and $D(SH)S^-$, respectively. The larger rate for the ionized species is not surprising for $PhO^•$. One may note, however, that the reverse trend should have occurred, if the cyclohexadienyl adducts refered to above were important at these pHs.[35]

Another interesting kinetic feature is the factor of 2–3 increase in the forward rate constant for β-ME $RS^•$ in going from $D(SH)_2$ to $L(SH)_2$ (see Table I). The presence of –OH groups in the former compound could possibly create tighter binding of water around the molecule and also have a small influence on the bond dissociation energy of the S–H group, causing the reaction to slow down. In contrast, the rates of attack of $RS^•$ of β-MPA on DTT and DTE are identical, as are those of the back-reactions. Thus within our experimental error the two stereoisomers react identically.

The rate constant for oxidation of $ClPz^+$ by $RS^•$ (see Results, section 4) is seen to be somewhat larger than those for $RS^•$ attack on $D(SH)_2$, but it is not as fast as for $ClPz^+$ oxidation by $PhO^•$. The central ring for $ClPz^+$ contains a thioether linkage and an N atom, each of which might be capable of forming transitory three-electron-bonded complexes with $RS^•$.[37] The present data do not preclude such complexes, but the only absorption seen on a 10-μs time scale was that of $ClPz^{•2+}$.

One objective of the present work was to determine how sensitive the redox potentials were to the structures of the aliphatic groups attached to the S atoms. (See structures at foot of Table II.) The approximately 1.2-kJ smaller values of ΔG°_{13} for β-MPA and β-ME in Table I were in excess of the experimental uncertainty of measurement of ΔG°_{13} within the series of RSH molecules (±500 J). This translates in larger values of $-\Delta G^\circ_{18}$ (or E°_{18}) for these two RSH species than for MAA and PenSH. That finding can be rationalized in terms of the close proximity to the sulfur atom of the electron-rich carboxylic group in MAA and of the two methyls in PenSH. However, the effects of these substitutions were not very large relative to the overall uncertainties for all intercomparisons, which are ±2 kJ mol^{-1} for ΔG° values in Table I and ±0.02 V in ΔE° values in Table II. The ΔE° results for the $DS_2^{•-}$ and $LS_2^{•-}$ species in Table II (their average values are 0.90 ± 0.02 and 0.92 ± 0.04 V, respectively) tend to exhibit a slight dependence on pH. Further study is required to determine whether this is caused by inaccuracies in the ionization constants or the presence of some dimer anions (e.g., $SDS^- \div SDSH^-$) at pH 9 and 10. However, the present data established ΔE°_6 for the two cases within ±0.04 V. These are very similar to ΔE°_{11} for the open-chain sulfhydryl β-ME (average value of $\Delta E^\circ_{11} = 0.89_2 \pm 0.004$ V).

A parameter not previously accessible is ΔG° for the ring closure of species such as $DS_2^{•-}$ (or $LS_2^{•-}$), viz.

$$^-S\text{–}D\text{–}S^• = DS_2^{•-} \qquad (20)$$

If one takes ΔE°_{18} for the $^-S\text{–}DS^•/^-S\text{–}D\text{–}SH$ couple to be the same as ΔE°_{18} (i.e., 0.50 V), then

$$\Delta G^\circ_{20} = -(\Delta E^\circ_{18} - \Delta E^\circ_6) \times 96487 - \Delta G^\circ_{19b}$$

Substituting $\Delta E^\circ_6 = 0.90$ V, one obtains $\Delta G^\circ_{20} \sim -17$ kJ mol^{-1}. This means that $K_{20} \sim 1000$, and therefore relatively little open-chain radical should exist. Comparison of ΔG°_{20} with the ΔG°_{19a} for β-ME[30] of −18.4 kJ mol^{-1} shows that, as expected, there is relatively little strain in the six-membered ring of $DS_2^{•-}$.

At this point the relative potentials in Table II can be placed on an absolute scale using the best available data for $ClPz^+$. A determination of the half-wave potential in 0.1 N H_2SO_4 has been reported by Kabaskalian and McGlotten.[26] Correction to zero ionic strength leads to $E^\circ_{16} = 0.84$ V. A more recent determination by cyclic voltammetry at pH 6 in 0.01 M phosphate buffer with precautions to minimize junction potentials[38] gave 0.83 V, in good agreement with the above. The observation that E°_{16} is higher than the formal potential 0.78 V reported by Pelizzetti and Mentasi in 1 M $HClO_4$[39] is not unexpected in view of the charges borne by the oxidized and reduced forms.

Combination of $E^\circ_{16} = 0.83$ V with $\Delta E^\circ_4 = 0.52_2$ V gives $E^\circ_4 = 1.35$ V. The agreement with 1.34 V calculated from gas-phase thermodynamic data in section 5 under Results is excellent. This supports the application of Schwarz and Dodson's result[34] to phenol, but this must be treated with some caution because of the uncertainties in the bond energy data.

Combination of E°_{16} with $\Delta E^\circ_{18} = 0.50$ V and $\Delta E^\circ_{11} = 0.90$ V leads to $E^\circ_{18} = 1.33$ V and $E^\circ_{11} = 1.72$ V for β-ME. These results may be compared with 1.42 and 1.80 V, respectively, derived from our calculations in ref 7 for sulfhydryl molecules of simple structure. The differences of 0.09 and 0.08 are within the overall uncertainty for the heats of formation and other quantities used in the calculations. Here one may note that the differences of 0.09 V in E°_{18} is of course carried forward into other potentials calculated from the original data (cf. E°_{21} and E°_{22} for β-ME below and in ref 7).

Taking $\Delta G^\circ_{19b} = 55\,480$ J mol^{-1} from ref 21, one obtains $E^\circ_{21} = 0.75$ V, which is similar to a value of 0.73 V reported by Prutz,

$$RS^• + e^- = RS^- \qquad (21)$$

Butler, Land, and Swallow[40] for cysteine equilibrated with tyrosyl radicals.

Finally one may estimate the magnitude of E°_{22}, a reaction

$$RSSR + e^- = RS\dot{S}R^- \qquad (22)$$

which is of some interest because of the suggestion that –S–S– groups in proteins may participate in electron transport.[41,42] This can be obtained by subtracting ΔG°_{11} from ΔG°_{23} or ΔG°_6 from

$$RSSR + 2e^- + 2H^+ = 2RSH \qquad (23)$$

the counterpart of (23) for disulfhydryl molecules. Utilizing the values of E°_{23} given by Clark[43] for open-chain disulfides (0.08_0 V) and lipoamide (0.14_4 V), one finds $E^\circ_{22} = -1.57$ V for β-ME and −1.60 V for LS_2. It is evident from these negative potentials that the –S–S– group will not readily accept electrons from other sites, unless their E° values are quite negative or there is strong stabilization of the $-S\dot{-}S-^-$ by neighboring groups.

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