Polarographic Reduction of the Azines[1]

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Abstract: The polarographic reduction of many azines is irreversible with the initial anion radical undergoing rapid reaction to give other products. Cyclic voltammetry is able in most cases to effect reoxidation of the anion radical before it reacts further, and gives the reversible electrode potential. The observed energy differences between azine and radical anion are well correlated with the results of CNDO and self-consistent field π-electron calculations.

The azines provide a group of small related compounds which should be useful in comparing the results of various approximate quantum chemical methods with experimental data. One quantity which was of interest to us is the energy required to perform a one-electron reduction and form an anion radical. As a result we have investigated the polarographic reductions of the azines.

Streitwieser[3] has summarized previous work which indicates for aromatic hydrocarbons a correlation of polarographic half-wave potentials with the energies of the lowest unfilled π-molecular orbitals calculated using the simple Hückel method. Streitwieser and Schwager[4] have made a comprehensive investigation of the polarographic reduction of aromatic hydrocarbons in dimethylformamide. It should be noted that benzene itself is not reducible at accessible polarographic potentials. With aza substitution into the aromatic ring, however, molecular orbital levels will be reduced in energy as a result of the lower energy of an electron in a nitrogen orbital as compared to carbon. Indeed, the azines are reducible, and we are concerned in this study with the effect of successive aza substitution both on the ease and on the nature of the reductions. We are particularly interested in whether the order of reduction potentials may be anticipated by molecular orbital methods, and we shall utilize methods more sophisticated than Hückel theory in this connection.

Polarography

The polarographic reduction of the azines in aqueous medium is a complex process which among other things is pH dependent. Our work has, therefore, been carried out in anhydrous dimethylformamide to ensure that we only observe reduction of the aromatic ring. Given[5] has shown that dimethylformamide is a particularly useful solvent for the reduction of pyridine and quinoline. Anthoine, *et al.*,[6] and Parkanyi and Zahradnik[7] also have examined the reduction of pyridine, quinoline, and isoquinoline in this solvent. The relatively high diffusion currents observed by these workers caused them to suggest that the electrode processes involved two electrons. Stone and Maki[8] have examined some of these compounds in dimethyl sulfoxide, but not in a quantitative fashion.

Polarographic data for a series of azaaromatic compounds were obtained. The results are summarized in Table I. It can be seen that the diffusion-current constants vary considerably, especially for the azabenzenes. The variation suggests that some compounds react *via* a net two-electron process and others *via* a net one-electron process. The two-electron process need not involve the simple transfer of two electrons, but more likely could involve an ECE mechanism in which two charge transfers are separated by a chemical change.

Table I. Polarographic Data for Reduction of Azaaromatic Compounds

Compd	$E_{1/2}$[a]	A[b]	Log plot slope[c]
Pyridine	−2.15	5.31	64
Pyrimidine	−1.78	3.51	55
Pyridazine	−1.61	5.59	76
Pyrazine	−1.57	4.61	59
s-Triazine	−1.47	2.43	53
as-Triazine	−1.04	3.89	60
s-Tetrazine	−0.29	2.18	60
(Naphthalene	−1.98	3.14	58)
Isoquinoline	−1.62	3.16	63
Quinoline	−1.59	3.83	56
Phthalazine	−1.41	3.82	50
Quinazoline	−1.22	2.91	46
Quinoxaline	−1.09	3.22	53
Cinnoline	−1.06	3.30	68
Pyrido[2,3-*b*]-pyrazine	−0.85	3.10	62
Pteridine	−0.52	2.14	52
(Biphenyl	−2.03	2.36	57)
2,2′-Bipyridyl	−1.59	3.59	60
4,4′-Bipyridyl	−1.31	2.60	56

[a] Potentials are in volts *vs.* a mercury pool. [b] A, the diffusion-current constant, $= i_d/m^{2/3}t^{1/6}c$ where i_d is the diffusion current in microamperes (taken at the maximum of the drop oscillations), m is the rate of mercury flow in milligrams/second, t is the drop time in seconds, and c is the concentration of substrates in millimoles/liter. [c] The slope of the plot of $-E$ (mV) *vs.* log $(i/(i_d - i))$.

The log plot slopes in Table I are all in the neighborhood of the 59-mV value one would expect for a reversible one-electron process. There are substantial deviations in some cases, however, and values of this

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[2] National Institutes of Health Predoctoral Fellow, 1965–1968; taken from part of the Ph.D. Thesis of T. P. L., Yale University, 1968.

[3] A. Streitwieser, Jr., "Molecular Orbital Theory for Organic Chemists," Wiley, New York, N. Y., 1961.

[4] A. Streitwieser and I. Schwager, *J. Phys. Chem.*, **66**, 2316 (1962).

[5] P. H. Given, *J. Chem. Soc.*, 2684 (1958).

[6] G. Anthoine, G. Coppens, J. Nasielski, and E. Vander Doncht, *Bull. Soc. Chim. Belg.*, 74, 65 (1964).

[7] C. Parkanyi and R. Zahradnik, *Abh. Deut. Akad. Wiss. Berlin, Kl. Chem., Geol.*, 363 (1964).

[8] E. W. Stone and A. H. Maki, *J. Chem. Phys.*, **39**, 1635 (1963).

magnitude may also be observed for irreversible two-electron processes. Considerable doubt exists, therefore, as to the number of electrons involved in the electrode processes and also as to the reversiblity of the reductions.

Pyridine and pyridazine were the only compounds to demonstrate substantial concentration dependence of the half-wave potential. With increasing concentration around the region of 10^{-3} M, the polarographic waves moved to more cathodic potentials. Moreover, pyridine and pyridazine are the only azines which do not demonstrate a linear relation between the diffusion current and $h^{1/2}$, the square root of the height of the mercury column above the electrode capillary. Such a linear dependence is characteristic of a process where the only mode of mass transfer to the electrode is diffusion. For pyridine and pyridazine, the diffusion current is proportional to a power of h intermediate between one-half and one. These characteristics are clearly indicative of the presence of adsorption prewaves before the main reduction waves.[9]

The log plot slope for pyridazine is 75 mV, and, therefore, larger than the anticipated 59 mV. Likewise, the pyridine slope (64 mV) is somewhat high. Large slopes are indicative of spread-out waves, and it seems clear that for both of these compounds, there is a prewave adsorption masked by the main reduction wave. The diffusion-current constants for pyridine and pyridazine were larger than for any of the other azines, and it is possible that these currents are enhanced by the adsorption effects. Although the adsorption prewaves could not be resolved from the main diffusion-controlled reduction by the polarographic experiment, as shall be seen, resolution proved possible using the cyclic voltammetric techniques. It may be added that pyridine is well known to show adsorption prewaves in aqueous solution.[10] Evidently the same holds in anhydrous aprotic media, and apparently pyridazine behaves similarly.

The polarographic results may be taken to indicate that the electrode processes are not generally simple one-electron reductions. Further, the significant differences in behavior among the azines suggest that the observed potentials may not be reasonable approximations to reversible electrode potentials. These reversible potentials, of course, are what we wish to utilize for correlations with molecular orbital calculations since they will be good approximations to the formal half-reaction potentials for the one-electron reductions.[11]

Cyclic Voltammetry

It seemed impractical to sort out the problems associated with the simple polarographic reduction and, therefore, most of our work was devoted to studying the reductions *via* cyclic voltammetry. The theory of cyclic voltammetry for various electrochemical processes has been developed and reviewed in the literature.[12] This method involves the application of a triangular wave potential to a stationary electrode with the current recorded as a function of voltage. For a reduction, one scans past the reduction potential and then back to the starting point. In the negative voltage scan the substrate is reduced, and in the positive scan the reduced form is reoxidized if it has not reacted chemically in the interim. A reversible one-electron process leads to a curve such as the one shown for pyrazine in Figure 1. The separation of anodic and cathodic peak currents is 57 mV, and the magnitude of the peak currents should be equal for the anodic and cathodic portions. The potential corresponding to the polarographic half-wave potential occurs 85.17% of the way up the wave and is at the same place for both the reduction and the oxidation.

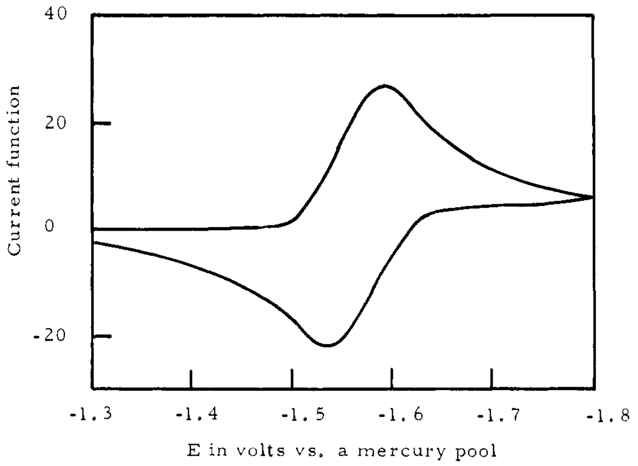

Figure 1. Cyclic voltammogram of pyrazine at 0.05 V/sec which illustrates reversible behavior. The current function is $i/v^{1/2}c$, where i is the current in milliamperes, v is the scan rate, and c is the concentration in millimoles/liter.

Cyclic voltammetry is useful because one can vary the rate of potential scan over a wide range and, therefore, obtain information on the behavior of the reduction product in the interim between reduction and oxidation. Thus, for example, if the reduction product reacts chemically, one may not see an oxidation wave at slow scans. On increasing the scan, however, one may be able to reoxidize the reduced species before it reacts. As one goes through the intermediary stages, one examines the ratio of the anodic to cathodic peak currents which will reach the limit of unity for the completely reversible case. As one eliminates chemical reactions, the whole wave form will also move somewhat toward more negative potentials.

The ECE situation is easily recognizable using cyclic voltammetry As one increases the scan rate and eliminates chemical reactions of the reduction product, one should also eliminate all subsequent charge transfers. The result will be a net decrease in the cathodic current with increasing scan rate for the ECE case.

We have performed cyclic voltammetric measurements on four N-methylazinium iodide salts as well as on the azabenzenes, azanaphthalenes, and azabiphenyls on which polarographic measurements had been made. In no case do we observe evidence for ECE behavior. For scan rates ranging from 0.02 to 10 V/sec, we observe anodic peaks at some scan rate

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(9) C. L. Perrin, *Progr. Phys. Org. Chem.*, 3, 177 (1965).

(10) M. K, Polievktov and S. G. Mairanovskii, *Izv. Akad. Nauk SSSR, Ser. Khim.*, 413 (1965), and earlier papers by Mairanovskii and coworkers.

(11) L. Meites in "Treatise on Analytical Chemistry," Part I, Vol. 4, I. M. Kolthoff and P. J. Elving, Ed., Wiley, New York, N. Y., 1963, p 2325.

(12) R. S, Nicholson and I. Shain, *Anal. Chem.*, 36, 706 (1964); 37, 178, 190 (1965); R. H. Wopschall and I. Shain, *ibid.*, 39, 1514, 1527, 1535 (1957); R. S. Nicholson, *ibid.*, 37, 1351 (1965); 39, 1406 (1967).

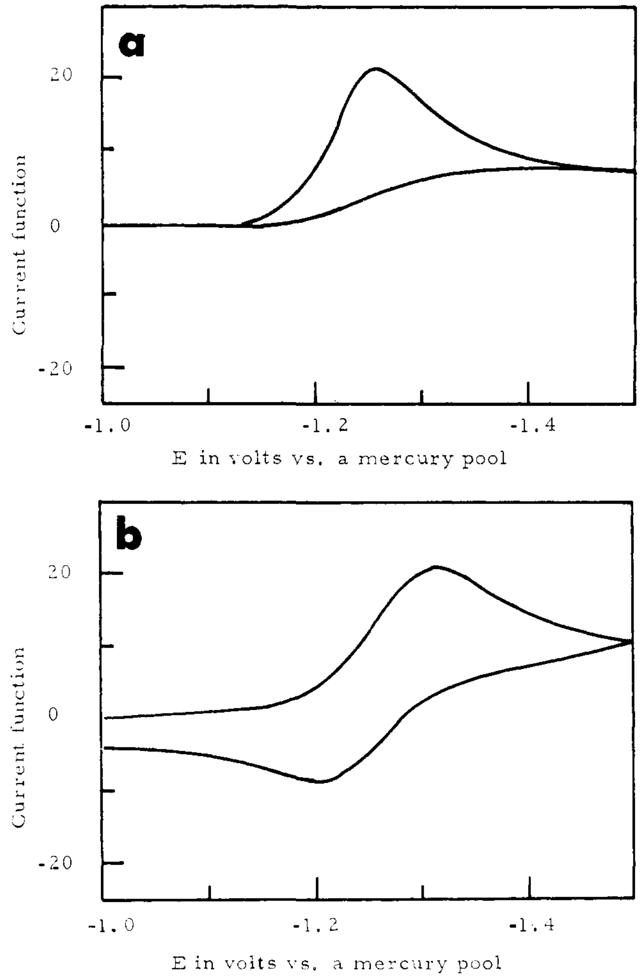

Figure 2. (a) Cyclic voltammogram of quinazoline at 0.05 V/sec. (b) Cyclic voltammogram of quinazoline at 10.0 V/sec.

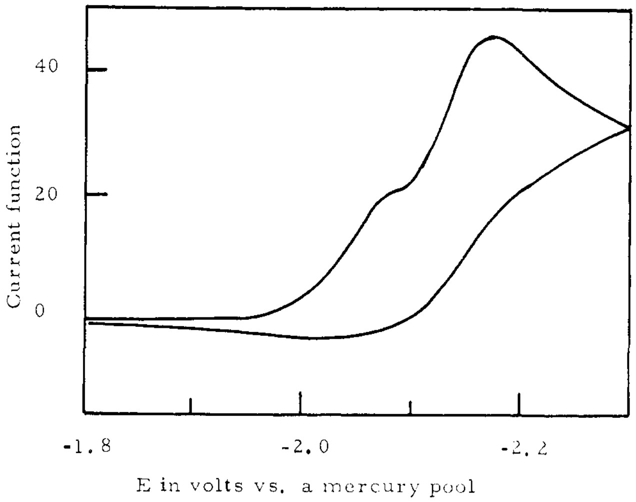

Figure 3. Cyclic voltammogram of pyridine at 0.05 V/sec illustrating prewave adsorption.

for all the neutral aza aromatics except for pyridine. The failure to observe an anodic peak at a particular scan rate is attributable to chemical reaction of the reduction product. Since no ECE character is observed, charge transfer does not follow chemical reaction. The strongest chemical effects are observed for pyridine and for three of the *N*-methylazinium salts for which anodic peaks could not be observed even at the fastest scans. Most of the neutral azines display lesser chemical effects which can be eliminated at some accessible scan. The elimination of chemical reaction of the anion radical reaction product with increasing scan rate is shown for quinazoline in Figure 2.

The prewave adsorptions for pyridine and pyridazine could be resolved from the main reduction by a voltammetric scan. This is illustrated for pyridine in Figure 3. Resolution is possible in the voltammetric experiment but not in the polarographic experiment because voltammetric currents reach a maximum and then decay, allowing better resolution of closely spaced peaks. Prewaves were not observed for any other compounds.

The cyclic voltammetric data indicate, therefore, that in no case is there transfer of more than one electron although in many cases there is chemical reaction of the anion radical reduction product. For almost all compounds the completely reversible one-electron reduction situation could be attained and reversible reduction potentials derived. In Table II these reversible half-wave potentials are listed together with a description of the reactivity of the one-electron anion radical reduction products.

Table II. Cyclic Voltammetric Reduction of Azaaromatic Compounds

Compd	Reversible $E_{1/2}$[a]	τ, sec[b]
Pyridine	-2.20[c]	<0.01
Pyrimidine	-1.822	0.06
Pyridazine	-1.657[d]	0.4
Pyrazine	-1.569	>25
s-Triazine	-1.527	0.02
as-Triazine	-1.057	30
s-Tetrazine	-0.292	20
(Naphthalene	-1.982	>20)
Isoquinoline	-1.642	1
Quinoline	-1.602	3
Phthalazine	-1.432	15
Quinazoline	-1.255	0.25
Quinoxaline	-1.097	>20
Cinnoline	-0.982	2
Pyridopyrazine	-0.857	6
Pteridine	-0.527	30
(Biphenyl	-2.032	>30)
2,2′-Bipyridyl	-1.602	10
4,4′-Bipyridyl	-1.305	>20
Pyridinium iodide	-0.750[c]	<0.01
Pyrimidinium iodide	-0.465[c]	<0.02
Pyridazinium iodide	-0.250[c]	<0.02
Pyrazinium iodide	-0.192	2

[a] Potentials are with reference to a mercury pool. The potentials were taken at the voltage corresponding to 85.2% of the maximum cathodic current for reversible waves and should be accurate to ±3 mV. [b] The ratio of the anodic to cathodic current was estimated using the semiempirical method of Nicholson.[12] The time from the half-wave potential to the switching potential for $i_a/i_c = 0.5$ is given as τ. This is an indication of the lifetime of the reduced species under the reaction conditions (approximately 5×10^{-4} *M* solutions) When τ is given as a minimum value, $i_a/i_c = 1.0$ even at the slowest sweep rates (0.01 V/sec). When τ is given as a maximum value, no anodic wave was found even at the fastest sweep rates (10–20 V/sec). [c] The reversible reduction potential for this compound was estimated from the peak positions at high scan rates. [d] Adsorption effects may lead to a small error for this value which applies to the diffusion-controlled reduction.

Energy Correlations with Molecular Orbital Calculations

We shall now compare our reversible half-wave potentials with the results of molecular orbital calculations. As has already been noted, the reversible half-wave potential provides a good estimate of the formal electrode potential which represents the energy difference between reactants and products for a particular electrode half-reaction. These reversible half-wave potentials differ somewhat from the original polarographic potentials for all cases where substantial chemical effects were observed. One of the results of eliminating the chemical complications is to shift the reduction potential in a cathodic direction.

Given, Peover, and Schoen[13] have noted that the electrocapillary maximum for dimethylformamide with 0.1 *M* tetraethylammonium iodide as the supporting electrolyte is approximately 0 V with reference to a mercury pool. Our cell-rest potential may, therefore, be taken as 0 V, and the reversible half-wave potentials will be direct measurements of the difference in energy in electron volts between azines and anion radicals.

We have calculated the energies of the azines and of their anion radicals using self-consistent field π-electron methods and also the Pople CNDO method[14] in which all of the valence electrons are considered. These procedures appear preferable to correlations using the lowest unoccupied Hückel orbitals because the interelectronic interactions arising from the addition of an electron will be included.

Table III compares the calculated difference in energy between the azine and its anion radical with the observed values. The π-electron self-consistent field calculations were carried out using both the Nishimoto–Mataga[15] and Pariser–Parr[16] integrals. *s*-Triazine is not included in the calculations since it has degenerate lowest unoccupied orbitals. The addition of an electron may then lead to two different electronic configurations of equal energy. In this case, the molecule may undergo a distortion on addition of an electron so as to eliminate the degeneracy of the orbitals. For configurations of nearly equal energy a simple ground state molecular orbital treatment is a poor approximation, and configuration interaction should be included in the calculations.

Table III. Comparison of Reversible Electrode Potentials for the Azines with Calculated Differences in Energy between Anion Radicals and Neutral Molecules

Compd	$-E$ (V), obsd	ΔE, calcd		
		CNDO[a]	NM int[b]	PP int[c]
Pyridine	2.20	2.10	1.28	2.79
Pyrimidine	1.822	1.90	1.13	2.75
Pyridazine	1.657	1.43	0.99	2.57
Pyrazine	1.569	1.21	0.87	2.52
as-Triazine	1.057	0.92	0.66	2.37
s-Tetrazine	0.292	0.43	0.29	2.07

[a] The calculations were carried out using the parameters suggested by Pople and Segal[14] and the geometry was taken as a regular hexagon. [b] The Nishimoto–Mataga integrals were used along with the valence state ionization potentials suggested by Nishimoto and Forster (*Theor. Chem. Acta*, **4**, 155 (1966); $W_{2p} = -11.42$ eV for N, -14.23 eV for C). The off-diagonal core elements were those suggested by Nishimoto and Mataga[15] ($\beta_{CC} = -2.388$, $\beta_{CN} = -2.576$, and $\beta_{NN} = -2.7$). [c] The Pariser–Parr integrals[16] were used along with the other parameters given above.

The CNDO results agree well with the observed values. The slope and intercept of a plot of observed against calculated values are 1.04 and 0.05, respectively, and the correlation coefficient is 0.951. The π-electron calculations using the NM integrals gave the correct ordering but too large a slope. Here

$$-E_{\text{obsd}} = 1.88\Delta E - 0.20 \qquad R = 0.991$$

Finally, the use of the PP integrals gave values which were too large, and had too large a slope

$$-E_{\text{obsd}} = 2.49\Delta E - 4.80 \qquad R = 0.981$$

Thus, the trend and magnitudes of the values were reproduced fairly well by the CNDO calculations, and the π-electron calculations using the NM integrals also were satisfactory except for the slope. The large slope and intercept found using the PP integrals suggest that they are not as satisfactory.

Calculations also were carried out for the azanaphthalene and azabiphenyl series. For these compounds, the CNDO calculations were prohibitively expensive. Thus, we have only used the π-electron method and the NM integrals. The results are given in Table IV. A plot of the observed values against calculated energies gives

$$-E_{\text{obsd}} \doteq 1.50\Delta E + 2.10 \qquad R = 0.958$$

The trend is reproduced quite well, although here an absurd intercept is found. The energy of the π-electron system is calculated to be decreased on the addition of an electron whereas the energy actually increases. The difficulty with the π-electron calculation probably arises from inadequate estimation of the electron correlation effects arising from addition of an electron. This situation could probably be improved by better semiempirical parametrization.

Table IV. Comparison of the Reversible Electrode Potentials for the Azanaphthalenes and Azabiphenyls with the Calculated Energies

Compd	$-E$, obsd	ΔE, eV
Naphthalene	1.982	−0.141
Isoquinoline	1.642	−0.271
Quinoline	1.602	−0.394
Phthalazine	1.432	−0.409
Quinazoline	1.255	−0.374
Quinoxaline	1.097	−0.609
Cinnoline	0.982	−0.693
Pyridopyrazine	0.857	−0.919
Pteridine	0.527	−1.035
Biphenyl	2.032	−0.113
2,2′-Bipyridyl	1.602	−0.349
4,4′-Bipyridyl	1.305	−0.594

It might, therefore, be concluded that successive aza substitution produces an irregular continual de-

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[13] P. H. Given, M. E. Peover, and J. Schoen, *J. Chem. Soc.*, 2764 (1958).

[14] J. A. Pople, D. P. Santry, and G. A. Segal, *J. Chem. Phys.*, **43**, S129 (1965); J. A. Pople and G. A. Segal, *ibid.*, **43**, S136 (1965); **44**, 3289 (1966).

[15] K. Nishimoto and N. Mataga, *Z. Phys. Chem.*, **12**, 335 (1957); N. Mataga and K, Nishimoto, *ibid.*, **13**, 140 (1957).

[16] R. Pariser and R. G. Parr, *J. Chem. Phys.*, **21**, 767 (1953).

crease in reduction potential which can be anticipated quite well by molecular orbital calculations.

It is also of interest of examine the reduction potentials for the *N*-methylazinium iodides. Values are shown in Table II for salts of pyridine and the three diazines, although for three of these salts reversible potentials were estimated since no anodic peaks could be observed. The same potential orderings were maintained as for the neutral compounds, and, in fact, the effect of quaternization seems merely to be a decrease in the reduction potential of about 1.4 V.

We should also mention some use of our electrochemical data by Ford, *et al.*[17] These workers have studied the metal to ligand charge-transfer bands of some pentaammineruthenium(II) complexes of azabenzenes which may be represented by the formula below where L

$$[(NH_3)_5Ru^{II}L]^{2+}(ClO_4^-)_2$$

is one of the azabenzenes. These authors have found that a good correlation exists between the energy of this charge-transfer band and our polarographic half-wave potentials. This is indicated in Table V where we have also shown the reversible half-wave potentials for which the correlation is slightly better. Evidently the charge transfer is to the azabenzene ligand, and the processes of metal to ligand charge transfer and electrochemical reduction involve similar energetics.

Table V. Correlation of Azabenzene Half-Wave Potentials with Energy of Charge-Transfer Band of Pentaammineruthenium(II)–Azabenzene Complexes

Ligand	$-E_{1/2}$[a]	$-E_{1/2}$[b]	ν, cm^{-1} × 10^4
Pyridine	2.15	2.20	2.45
Pyrimidine	1.78	1.822	2.25
Pyridazine	1.61	1.657	2.14
Pyrazine	1.57	1.569	2.12
s-Triazine	1.47	1.527	2.13

[a] Polarographic half-wave potential in volts. [b] Reversible half-wave potential in volts.

Chemical Behavior of Anion Radicals

It is clear from cyclic voltammetry that a number of the anion radicals produced on one-electron reduction react rapidly with themselves or with some other species. The most likely modes of reaction are: (1) proton abstraction by the anion radical from the solvent or supporting electrolyte; (2) addition of the anion radicals to the carbonyl group of dimethylformamide; (3) dimerization of the anion radicals; and (4) the addition of the anion radical to the parent azine.

We conducted some controlled potential electrolysis experiments on pyrazine and pyridazine hoping originally to obtain evidence for one-electron reductions by integration of the current-time curves to get the total charge transferred. We found, however, that electrolysis was complete after only about 0.25 electron per molecule was transferred and a black polymeric precipitate had formed. It was demonstrated by vapor-phase chromatographic analysis using a flame ionization detector that no azine was left in solution at this point. Both pyrazine and pyridazine displayed identical behavior. This result suggested that our fourth possibility was the most likely, since only this mechanism would explain the completion of electrolysis before transfer of at least one electron.

Addition of the anion radical to the parent azine may be best represented by the scheme below. The

$$A + e^- \longrightarrow A^- \xrightarrow{A} A\text{–}A^- \text{ products}$$

anion radical performs a nucleophilic attack on the parent, and as in the case of any bimolecular reaction, the reactivity of each of the participating species will contribute toward the reaction energetics. The rate of loss of the anion radical should, therefore, depend on its nucleophilicity and also on the susceptibility of the azine to nucleophilic attack. It is well known that aza substitution greatly activates an aromatic ring for nucleophilic attack, and that the effects appear predominantly in the positions ortho and para to the nitrogen atom.[18] On the other hand, aza substitution should stabilize the anion radical reduction product and decrease its nucleophilicity. The fact that reduction potentials are reduced with successive aza substitution reflects this increased stabilization.

It should be noted that neither of the two hydrocarbons (naphthalene and biphenyl) listed in Table II display chemical effects, and it is only with aza substitution that these phenomena manifest themselves. The data also indicate that except for pyridine, the largest chemical effects are displayed by compounds in which nitrogen atoms are in a 1,3 arrangement. The examples are pyrimidine, *s*-triazine, and quinazoline. The 1,3 arrangement of nitrogen atoms permits the nitrogens to simultaneously activate positions ortho and para to each for attack by a nucleophilic agent. Further nitrogen substitution as with triazine, pteridine, and *s*-tetrazine decreases the rate of reaction of the anion radical. The added nitrogen acts by decreasing the nucleophilicity of the anion radical. The high reactivity of pyridine probably results from the small stabilization of the anion radical making it particularly reactive in nucleophilic attack.

The observations for the *N*-methylazinium salts also fit with this hypothesis. Nitrogen alkylation should greatly activate the aromatic ring for nucleophilic attack, and although more stable anion radicals should be produced, the ring activating effect appears to predominate.

The arguments become even clearer when one examines the CNDO charge densities shown in Table VI for the azabenzenes and the *N*-methylazinium salts. It is reasonable to assume that the greater the positive charge density on a position in the neutral molecule, the more susceptible it is to nucleophilic attack. The larger the negative charge density on a position in the anion radical the greater the tendency for it to perform a nucleophilic attack.

Examination of Table VI reveals that *s*-triazine is the parent azine with the most positively charged carbon position. It displays the greatest chemical effects of any neutral compound except pyridine. The second most positively charged parent position is the pyrimidine 3 position. The pyrimidine 4 and 6 positions in the anion radical retain considerable negative charge,

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[17] P. Ford, DeF. P. Rudd, R. Gaunder, and H. Taube, *J. Amer. Chem. Soc.*, 90, 1187 (1968).

[18] G. Illuminati, *Advan. Heterocycl. Chem.*, 3, 285 (1964).

Table VI. CNDO Charge Densities for Azabenzenes and Azinium Salts[a]

Compd	Position	q, NA[b]	q, AR[c]
Pyridine	2,6	+0.083	−0.011
	3,5	−0.023	−0.018
	4	+0.035	−0.147
Pyrazine	2,3,5,6	+0.055	+0.007
Pyrimidine	2	+0.156	+0.158
	4,6	+0.115	−0.068
	5	−0.043	−0.007
Pyridazine	3,6	+0.061	+0.052
	4,5	+0.016	+0.097
s-Triazine	2,4,6	+0.183	*d*
as-Triazine	3	+0.135	+0.130
	5	+0.091	−0.010
	6	+0.035	+0.013
s-Tetrazine	3,6	+0.110	+0.116
Pyridinium ion[e]	2,6	+0.163	+0.012
	3,5	−0.013	+0.022
	4	+0.122	−0.058
Pyrazinium ion	2,6	+0.140	*f*
	3,5	+0.063	*f*
Pyridazinium ion	3	+0.082	*f*
	4	+0.094	*f*
	5	+0.018	*f*
	6	+0.154	*f*
Pyrimidinium ion	2	+0.237	*f*
	4	+0.197	*f*
	5	−0.036	*f*
	6	+0.190	*f*

[a] All charge densities are defined as the core charge on the atom minus the calculated electron density. [b] This column is for the neutral azine charge densities. [c] This column is for the anion radical charge densities. [d] No valid calculation could be performed on the *s*-triazinium anion radical. [e] The numbering scheme starts from the quaternary nitrogen. [f] Calculations were not performed for this anion radical.

and pyridimine displays large chemical effects. *as*-Triazine and *s*-tetrazine also have nitrogen atoms in a 1,3 orientation. Their carbon ring positions are relatively positively charged, but not as much as for *s*-triazine or pyrimidine. Negative charges in the anion radicals are much smaller, however, and this is probably the main reason why very small and no chemical effects are observed for *as*-triazine and *s*-tetrazine, respectively.

Pyridine has by far the largest negative charge in its anion radical, and its large chemical effects are attributable mainly to this, It, however, also has relatively positively charged 2 and 6 ring positions in the parent favoring nucleophilic attack. The complete absence of chemical effects for pyrazine is also easily explained. It has the least positively charged ring positions of any of the azabenzenes, and its anion radical has no position with negative charge density. The moderate chemical effects observed for pyridazine are also in accord with the charge densities.

In passing on to the four *N*-methylazinium iodide salts, one notes that very large positive charge densities are predicted in the parents and, indeed, very strong chemical effects are observed for these species. The only anion radical calculation performed for these compounds was on the anion radical of the pyridinium cation. This calculation can be seen to predict a reasonably large negative charge density in the 4 position, and the large chemical effects observed for the *N*-methylpyridinium salt seem reasonable. The pyrimidinium salt displays the largest chemical effects of any species we have examined, and this is in accord with the very large positive charge densities predicted by the calculations. The pyrazinium ion shows only moderate chemical effects in accord with its reduced charge densities.

The same arguments can be invoked for the azanaphthalene and azabiphenyl series using the π-electron charge densities shown in Table VII. One sees that quinazoline should show the strongest chemical behavior, as indeed it does.

Table VII. Extremes of π-Electron Charge Densities for Azanaphthalenes and Azabiphenyls[a]

Compd	Position	q, NA[b]	q, AR[b]
Naphthalene	All	0.000	
	1,4,5,8		−0.178
Quinoline	2	+0.106	
	5,8		−0.165
Isoquinoline	1	+0.106	
	4		−0.180
Cinnoline	3	+0.061	
	4		−0.158
Phthalazine	1,4	+0.096	
	5,8		−0.158
Quinoxaline	2,3	+0.095	
	5,8		−0.149
Quinazoline	2	+0.174	
	5		−0.161
Pyridopyrazine	6	+0.117	
	8		−0.126
Pteridine	2	+0.183	
	4		−0.047
Biphenyl	All	0.000	
	4,4′		−0.174
2,2′-Bipyridyl	3,3′	+0.087	
	4,4′		−0.168
4,4′-Bipyridyl	3,5,3′,5′	+0.088	
	2,6,2′,6′		−0.074

[a] All charge densities are calculated using a π-electron SCF MO method with Nishimoto–Mataga repulsion integrals. Charge densities are defined in footnote *a* in Table VI. [b] See footnotes *b* and *c* in Table VI.

The chemical behavior which was observed is in agreement with esr studies. It has been reported that pyridine, pyrimidine, and *s*-triazine do not form stable anion radicals in dimethyl sulfoxide which may be observed by esr.[6] Stone and Maki found the *s*-tetrazine anion radical more difficult to observe than the pyrazine one and the pyridazine anion radical significantly more difficult to observe than either. This is in agreement with our orderings for the reactivity of the anion radicals. Anion radicals have been observed by esr for phthalazine,[8] quinoxaline,[19] and 4,4′-bipyridyl.[20] Our data indicate that chemical effects are virtually nonexistent for these species, and it is not surprising that anion radicals may be observed.

After this work was completed, Chaudhuri, *et al.*,[21] reported observation in the esr of anion radicals of pyridine, quinoline, and isoquinoline. These species were generated by chemical reduction, and the authors report that the anion radicals are stable only in hexamethylphosphoramide with dimerization to dianions

(19) R. L. Ward, *J. Amer. Chem. Soc.*, **84**, 332 (1962); J. C. M. Henning, *J. Chem. Phys.*, **44**, 2139 (1966).

(20) A. Carrington and J. dos Santos-Veiga, *Mol. Phys.*, **5**, 22 (1962); R. L. Ward, *J. Amer. Chem. Soc.*, **83**, 3623 (1961).

(21) J. Chaudhuri, S. Kume, J. Jagur-Grodzinski, and M. Szwarc, *ibid.*, **90**, 6421 (1968).

occurring in other solvents. This dimerization, however, might just as easily involve nucleophilic attack of anion radical on parent as in our proposed mechanism. Indeed, these authors have given no direct evidence that dianions, rather than dimeric anion radicals, are involved. However, it should be noted that in multiple scan cyclic voltammograms we have not been able to detect dimeric products, although we have looked for them.

Experimental Section

Chemicals. Pyridine, pyrazine, pyrimidine, pyridazine, *s*-triazine, quinoline, isoquinoline, naphthalene, quinoxaline, cinnoline, phthalazine, quinazoline, pyrido[2,3-*b*]pyrazine, biphenyl, 2,2′-bipyridyl, and 4,4′-bipyridyl were obtained from commercial sources and were further purified by either distillation or sublimation.

s-Tetrazine was synthesized according to the directions of Spencer, Cross, and Wiberg;[22] *as*-triazine was synthesized according to the method of Paudler and Barton;[23] pteridine was synthesized by the procedure of Albert, Brown, and Cheeseman.[24]

N,N-Dimethylformamide obtained from Matheson Coleman and Bell was fractionally distilled from calcium hydride at atmospheric pressure. A middle fraction of the constant boiling distillate was used.

Tetraethylammonium iodide was prepared by mixing equimolar amounts of triethylamine and ethyl iodide in nitrobenzene. The mixture was heated on the steam bath for several hours, cooled, and filtered. The precipitated material was washed with ether and recrystallized at least five times from absolute ethanol. It was dried in an oven at 100°. Tetrabutylammonium iodide was purchased from Eastman Organic Chemicals.

Methiodide salts of pyridine and the diazines were prepared by mixing equimolar amounts of the azine and methyl iodide either neat or in ethanol. Products were recrystallized two to three times from ethanol or ethanol–hexane. Melting points were the same as those reported in the literature, and elemental analyses were satisfactory.

Electrochemical Instrumentation. Polarograms were performed in a Sargent No. S-29390 water-jacketed cell at 25.00 ± 0.01°. The cyclic voltammetric cell was a Pyrex No. 50 O-ring joint (i.d. = 50 mm) closed off at the bottom to accommodate a cell volume of about 60 ml. A platinum wire was sealed through the bottom to establish contact with a mercury pool used as the counterelectrode during operation. A Teflon top was constructed to fit snugly into the O ring with three holes drilled to provide for insertion of working and reference electrodes and a gas-inlet tube for degassing with nitrogen.

All investigations were conducted using a three-electrode configuration. The working electrode for polarography was, of course, a dropping mercury electrode, and the working electrode for cyclic voltammetry was a platinum wire hanging mercury drop electrode of the type described by Shain.[25] The reference and counterelectrodes for both polarography and cyclic voltammetry were mercury pools. The counterelectrode was the pool at the bottom of the cell. The reference electrode was at the bottom of a tube inserted in the cell and containing cell solution. A capillary probe came off the tube above the pool and could be adjusted so that it was about a millimeter from the working electrode. Contact with this pool was established through a platinum wire inside a thin glass tube.

Some cyclic voltammetric runs were also conducted using a saturated calomel electrode reference. In this arrangement, a Beckman fiber junction saturated calomel electrode was placed in a tube inserted in the cell. The tube consisted of two parts: a solution of 1 *M* sodium nitrate separated by an ultrafine fritted disc from a capillary probe compartment containing cell solution. The runs using the sce gave the same currents and same reversibility behavior as the mercury pool. The mercury pool reference, however, was used predominantly through this work because of its convenience. Given and Peover[26] have gone into some detail on the stability of the mercury pool in dimethylformamide under a solution of tetraethylammonium iodide. They note that the mercury pool anode assumes a well-regulated value suitable for reference purposes. We have always observed full stability and reproducibility for the mercury pool reference under these conditions. As has been noted, results using an external saturated calomel electrode are entirely equivalent to those at the mercury pool. The mercury pool is 516-mV cathodic of the sce.

All investigations were conducted using a Wenking 61RS Fast Rise potentiostat. The triangular wave for cyclic voltammetry as well as the single voltage scan for polarography were generated using an EAI TR-20 analog computer. The cyclic voltammograms were recorded on a Tektronix Model RM 564 storage oscilloscope with two Type 2A63 differential amplifiers as vertical and horizontal inputs. The oscilloscope was equipped with a Polaroid camera. Scans were taken and recorded on the screen using the oscilloscope storage faculty. Desired runs were photographed.

Solutions were prepared by careful dilution. Dimethylformamide with 0.1 *M* supporting electrolyte was used to dissolve and dilute carefully weighed samples of azines. The supporting electrolyte was tetraethylammonium iodide except for scans above -2.0 V *vs.* a mercury pool when the tetrabutylammonium iodide salt was used. The general procedure for a run was as follows. A solution of azine (25 ml) was placed in the cell which was fitted with the electrodes and gas inlet tube. The cell was degassed by bubbling dry oxygen-free nitrogen through the solution. After 15–20 min of degassing, the electrochemical operations were commenced with nitrogen now blown over the solution. Solutions for cyclic voltammetry were generally of the range $2\text{–}5 \times 10^{-4}$ *M*. Cyclic voltammetric investigations were conducted at room temperature of 23–25°. Solutions for polarography were of the order of 1×10^{-3} *M*. The electrochemical scan was started by beginning the integration on the analog computer. A fresh mercury drop was hung for each scan in the cyclic voltammetry. Scan rates ranging from 20 mV/sec to 10 V/sec were used.

Cell resistances were measured using a General Radio Company Type 1650-A impedance bridge. Resistance between the hanging drop electrode and the counterelectrode was 425 ohms, and the resistance between the polarographic dropping mercury electrode and its counterelectrode was 635 ohms.

Some coulometric (controlled potential electrolysis) determinations were also carried out on 10^{-3} *M* solutions of pyrazine and pyridazine in 0.1 *M* tetraethylammonium iodide in dimethylformamide. The measurements were made in the voltammetric cell, and currents and potentials were controlled by the Brinkmann potentiostat. The working electrode was the mercury pool at the bottom of the cell, and another mercury pool was the reference electrode. The counterelectrode was a silver wire placed in a compartment separated from the cell by a fine fritted disk and containing cell solution. A constant potential above the reversible half-wave potential was maintained by the potential control circuit of the potentiostat, and currents were read on the potentiostat meter and recorded as a function of time. Solutions were, of course, carefully degassed before all measurements and were stirred mechanically during electrolysis. Vapor-phase chromatographic analyses of the electrolysis solutions were performed on the Aerograph Hy-Fi Model 600-C apparatus with a flame ionization detector and a Model 650 hydrogen generator. This instrument allowed observation of the minute quantities of materials involved.

Molecular Orbital Calculations. π molecular orbital calculations were performed using a revised Fortran IV version of a program written by H. M. Gladney and obtained from the Quantum Chemistry Program Exchange, Bloomington, Ind. (QCPE Program No. 17).

CNDO calculations were carried out using a modified version of the CNDO/2 program written by G. A. Segal and also obtained through QCPE (QCPE Program No. 91).

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