A Method for Estimating Reduction and Standard Potentials of Unconjugated Alkyl Radicals
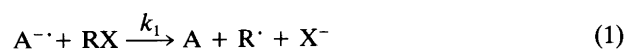
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Occhialini, D., Pedersen, S. U. and Lund, H., 1990. A Method for Estimating Reduction and Standard Potentials of Unconjugated Alkyl Radicals. – Acta Chem. Scand. 44: 715–719.

Anion radicals of aromatic compounds react with alkyl halides by a dissociative electron transfer (ET) reaction. The alkyl radical thus formed might react with the aromatic anion radical either by coupling (k_2) or in a second ET reaction (k_3). The competition which can be expressed by the dimensionless parameter $q = k_3/(k_3 + k_2)$, may be used to estimate the standard potential of the alkyl radical, when it is assumed (i) that the rate constant k_2 is the same for all combinations of aromatic anion radicals and all alkyl radicals (the value 1×10^9 M^{-1} s^{-1} is assumed here) and (ii) the rate constant k_3 for the ET reaction from the anion radical to the alkyl radical can be expressed according to the Marcus theory by the reorganisation energy and the free energy change of the ET reaction. A characteristic reduction potential $E_{1/2}^q$ can be determined from plots of q versus the standard potential of the mediator, E_A°. λ, the reorganisation energy, can be estimated from fitting the experimental curves depicting q vs. E_A° to calculated curves. The standard potential of the redox pair alkyl anion/alkyl radical, E_R°, can readily be calculated when $E_{1/2}^q$ and λ are known. The average reorganisation energy of the ET from the aromatic anion radicals to the alkyl radicals is estimated to be 30 kcal mol^{-1} which means that the self-exchange reorganisation energy for the unconjugated alkyl radicals is about 50 kcal mol^{-1}. The following standard potentials were found for the n-, sec- and $tert$-butyl radicals, -1.62, -1.72, -1.77 V vs. SCE, respectively.

The reaction between electro-generated anion radicals of aromatic compounds ($A^{-\cdot}$) and alkyl halides (RX) has been studied over the last 15 years;[1–4] reactions (1)–(3) describe the mechanism. Reaction (1) is the rate-determining electron transfer (ET) to the alkyl halide. This ET reaction is generally regarded as being dissociative, i.e., the ET step and C–X-bond breakage are concerted, and the ET reaction leads immediately to the alkyl radical and the halide ion.[2]

$$A + e^- \rightleftarrows A^{-\cdot} \qquad E_A^\circ$$

$$A^{-\cdot} + RX \xrightarrow{k_1} A + R^\cdot + X^- \qquad (1)$$

$$A^{-\cdot} + R^\cdot \xrightarrow{k_2} A\text{-}R^- \qquad (2)$$

$$A^{-\cdot} + R^\cdot \xrightarrow{k_3} A + R^- \qquad (3)$$

It has been noticed in many investigations that the overall reaction of aromatic anion radicals with benzylic halides is a reduction of RX,[3] whereas the overall reaction of unconjugated alkyl halides was an alkylation of the mediator.[4] Recently this behaviour has been ascribed to the difference between the standard potentials of the alkyl radical and the mediator.[5] Here we describe the competition between (2) and (3) more explicitly by assuming that (i) the

rate constant for the coupling reaction (2), k_2, to a first approximation, is the same for all the combinations of aromatic anion radicals with all the alkyl radicals used in this investigation ignoring steric effects which seem to have only a minor effect on such reactions and (ii) the rate constant k_3 for the ET reaction between aromatic anion radicals and the alkyl radicals can be calculated according to the Marcus theory[6] from the reorganisation energy and the driving force.

Reaction (2) is a fast coupling of an anion radical with a neutral radical which involves no charge repulsion. Such reactions are generally believed to be very fast, probably near the diffusion limit. Garst[7] has published the value 2×10^9 M^{-1} s^{-1} for the rate constant for the coupling of sodium naphthalene anion radical and the 1-hexenyl radical in THF. The influence of the strong ion-pairing in THF on the measurement of k_2 is unknown. Another problem is that homogeneous conditions were assumed in his calculations but the initial reaction between the naphthalene anion radical and the 6-bromo-1-hexene seems much too fast (ca. 5×10^7 M^{-1} s^{-1}) for complete mixing before the reaction is finished.

Preliminary results of measurements of k_2 for the reaction between electrogenerated aromatic anion radicals and 'radical clocks' have indicated that k_2 is close to 1×10^9 M^{-1} s^{-1};[8] a value that will be used in the following for all

coupling reactions between aromatic anion radicals and all alkyl radicals. The steric influence on the rate of the coupling is neglected to a first approximation.

Reaction (3) is an outer-sphere ET reaction between the mediator and the alkyl radical. The rate constant k_3 can according to Marcus theory[9] be expressed [eqn. (4)] through the reorganisation energy, λ, and the driving force.

$$k_3 = \frac{k_d}{1 + A\left\{\left[\frac{\lambda}{4}\left(1 + \frac{\Delta G°}{\lambda}\right)^2\right]/RT\right\}} \tag{4}$$

Here k_d is the diffusion-limited rate constant, which is taken to be 1×10^{10} M^{-1} s^{-1} in DMF, A is the pre-exponential parameter for which a value of 0.2 is chosen.[9]

The reorganisation energy λ for the reaction can, according to Marcus' cross relation,[10] be expressed as the average of the self-exchange reorganisation energies for the electron donor, λ_{A^-} (0) and for the electron acceptor, λ_R (0) as shown in eqn. (5). The self-exchange reorganisation energy for the aromatic anion radicals is usually 10 kcal mol^{-1} with very small deviations.[9]

The driving force $\Delta G°$ can be expressed by the difference in redox potentials between the aromatic anion radical ($E_A°$) and the alkyl radical ($E_R°$) as shown in eqn. (6).

$$\lambda = \frac{1}{2}[\lambda_{A^-}(0) + \lambda_R(0)] \tag{5}$$

$$\Delta G° = -F(E_R° - E_A°) \tag{6}$$

The competition between reactions (2) and (3) can be expressed[11] by the dimensionless parameter $q = k_3/(k_2 + k_3)$; q is equal to 0 when the overall reaction is an alkylation of the mediator and equal to 1 when the overall reaction is a $2e^-$ reduction of RX. The competition parameter q can be measured to a precision of ± 0.1 by linear-sweep voltammetry (LSV) or by coulometry.

Kern and Federlin[12] have measured the anodic half-wave potentials at a stationary Pt electrode of several carbanions in DMSO and/or 90:10 dioxane–water and correlated the potentials with the pK_a of the conjugate acid. The correlation between the anodic half-wave potentials of the anions and the standard potentials of the anion/radical is uncertain, mainly due to lack of knowledge of the rate of the chemical follow-up reaction and the rate of the heterogeneous ET. However, within each family of enolate ions these factors seem to be reasonably constant since an acceptable correlation between the measured potentials and measured pK_a values was obtained, the correlation being $E_{1/2}(ox) = E_{1/2}(ox,0) - 0.085 \ pK_a$, where $E_{1/2}(ox,0)$ is the value of $E_{1/2}(ox)$ extrapolated to $pK_a = 0$.

Recently Wayner and Griller[13] measured reduction potentials for substituted benzyl radicals by using modulated photolysis for the generation of radicals and voltammetry tuned to the modulation of the photolysis for the measurement. In these studies the measured $E_{1/2}$ were not related to the standard potentials, but as pointed out by Savéant et

$al.$[14] the difference between $E_{1/2}$ and $E°$ is expected to be smaller for benzylic radicals than for unconjugated alkyl radicals. The method of photogeneration and electrochemical detection of the radicals is possible only for relatively stable radicals such as benzylic radicals, but it is not applicable for measuring potentials of the unconjugated alkyl radicals because the lifetime will be too short for detection.

Savéant et al.[14] have recently investigated the butyl iodides by CV in DMF. From the second reduction peak which is due to the direct reduction of the butyl radicals and by comparing experimental voltammograms with simulated voltammograms the reduction potentials, the self-exchange reorganisation energies, and the standard potentials could be estimated for tert-butyl and sec-butyl radical. For butyl iodide, no second reduction peak was observed in the voltammogram. The method of Savéant is applicable only for compounds where a reduction peak due to the reduction of the radical can be observed in the voltammogram. In the fitting procedure between experimental and digital simulated voltammograms, the transfer coefficients α_1 and α_2, the rate constants $k_1°$ and $k_2°$, and E^R are adjusted simultaneously. This procedure involves the danger of introducing errors but nevertheless, the values obtained for E^R agrees reasonably well with the results presented below.

The redox properties of the system radical/anion is of interest for many aspects of organic chemistry. Besides the connection between redox potentials and basicity mentioned earlier, the redox potentials of a radical and of the substrate are important for predicting the course of the reaction. A straightforward example is the reaction between anion radicals and benzylic radicals where strongly reducing anion radicals produce only toluene, whereas less strongly reducing anion radicals give coupling products.

In this investigation the following unconjugated alkyl chlorides have been included: butyl chloride (1), sec-butyl chloride (2), tert-butyl chloride (3), chlorocyclohexane (4), chlorocyclopentane (5), 1-chloro-2-methylpropane (6), and 2-chloropropane (7) representing primary, secondary and tertiary alkyl radicals.

Results and discussion

LSV was generally used to measure q except for the systems where the rate constant k_1 is so small that no significant catalytic effect can be observed. In these cases coulometry is the method of choice. The accuracy in q is not better than ± 0.1. The data for different combinations of mediators and alkyl chlorides are shown in Table 1. It is seen that for a given alkyl chloride q is a monotonically decreasing function of the standard potential of the mediator. For the compounds 2, 3, 5 and 7 the upper limit of q ($q = 1$) was not reached because the potentials were too close to the discharge of the medium.

In all cases, however, the redox potential of an imaginary compound which will cause $q = \frac{1}{2}$ can be determined by interpolation. This potential will in the following be referred to as $E^q_{1/2}$.

Table 1. Competition parameters (q), in the reaction between anion radicals ($A^{-\cdot}$) and aliphatic halides (RX) in DMF/0.1 M TBABF$_4$ measured by LSV.

RX	A	$-E_A^\circ / V^b)$	q
1	Anthracene	1.890	0.0[a1]
	p-Methoxyacetophenone	2.059	0.0[a2]
	Methyl benzoate	2.183	0.5
	Benzonitrile	2.230	0.6
	m-Toluonitrile	2.264	0.6
	o-Toluonitrile	2.277	0.7
	p-Toluonitrile	2.337	0.8
	Naphthalene	2.457	0.9
	Biphenyl	2.519	0.9
	1-Methoxynaphthalene	2.537	1.0
2	Anthracene	1.890	0.0[a3]
	p-Methoxyacetophenone	2.059	0.0[a4]
	Benzonitrile	2.230	0.2
	m-Toluonitrile	2.264	0.2
	o-Toluonitrile	2.277	0.2
	p-Toluonitrile	2.337	0.6
	Naphthalene	2.457	0.6
	Biphenyl	2.519	0.8
	1-Methoxynaphthalene	2.537	0.8
3	Anthracene	1.890	0.0[a5]
	2-Dimethylaminopyrene	2.067	0.2
	3-Methylisoquinoline	2.154	0.2
	Benzonitrile	2.230	0.2
	m-Toluonitrile	2.264	0.2
	o-Toluonitrile	2.277	0.2
	p-Toluonitrile	2.337	0.5
	Biphenyl	2.519	0.6
	1-Methoxynaphthalene	2.537	0.6
4	Anthracene	1.890	0.0[a6]
	Isoquinoline	2.131	0.2
	Methyl benzoate	2.183	0.3
	Benzonitrile	2.230	0.4
	m-Toluonitrile	2.264	0.5
	o-Toluonitrile	2.277	0.6
	p-Toluonitrile	2.337	0.9
	Biphenyl	2.519	1.0
	1-Methoxynaphthalene	2.537	1.0
5	Anthracene	1.890	0.0[a7]
	p-Methoxyacetophenone	2.059	0.0[a8]
	Benzonitrile	2.230	0.2
	m-Toluonitrile	2.264	0.3
	o-Toluonitrile	2.277	0.3
	p-Toluonitrile	2.337	0.6
	Naphthalene	2.457	0.6
	1-Methoxynaphthalene	2.537	0.7
6	Anthracene	1.890	0.0[a9]
	Isoquinoline	2.131	0.5
	Methyl benzoate	2.183	0.5
	Benzonitrile	2.230	0.8
	m-Toluonitrile	2.264	0.8
	o-Toluonitrile	2.277	0.8
	p-Toluonitrile	2.337	1.0
	Naphthalene	2.457	1.0
7	Anthracene	1.890	0.0[a10]
	Isoquinoline	2.131	0.2
	Benzonitrile	2.230	0.3
	o-Toluonitrile	2.277	0.3
	p-Toluonitrile	2.337	0.5
	Naphthalene	2.457	0.5
	Biphenyl	2.519	0.7
	1-Methoxynaphthalene	2.537	0.7

[a]From coulometric experiments: [a1] $n = 2.12$; [a2] $n = 2.06$; [a3] $n = 2.28$; [a4] $n = 2.12$; [a5] $n = 2.14$; [a6] $n = 1.94$; [a7] $n = 2.17$; [a8] $n = 2.02$; [a9] $n = 2.11$; [a10] $n = 1.94$. [b]vs. SCE.
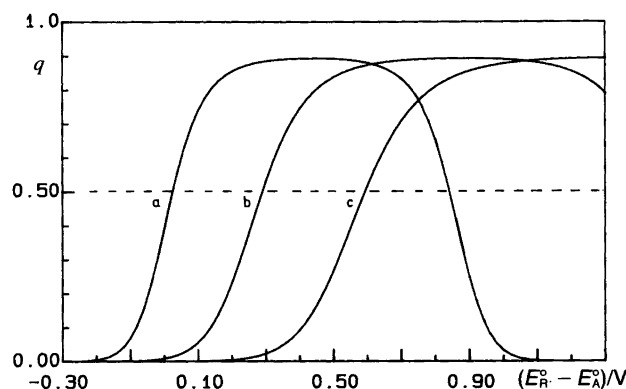

Fig. 1. q plotted versus the potential difference $E_R^\circ - E_A^\circ$ for three different values of λ using a constant value of $k_2 = 1 \times 10^9$ M^{-1} s^{-1}: a, $\lambda = 10$ kcal mol^{-1}; b, $\lambda = 20$ kcal mol^{-1}; c, $\lambda = 30$ kcal mol^{-1}.

In Fig. 1 q versus the potential difference $E_R^\circ - E_A^\circ$ is plotted for different values of λ and a constant value of $k_2 = 1 \times 10^9$ M^{-1} s^{-1}; the plot includes the Marcus 'inverted region.' The width of the S-shaped curves can be used to estimate λ, and the difference between $E^q_{1/2}$ and E_R° can also be read on the curve by following the dotted line for q equal to ½. The determination of $E^q_{1/2}$ is reasonably good and the uncertainty is estimated to be less than 50 mV.

An average value of 30 kcal mol^{-1} is found for λ of the ET to the alkyl radicals and this value is used to determine the potential correction for transforming the reduction potentials ($E^q_{1/2}$) to standard potentials. The potential correction is 580 mV for $\lambda = 30$ kcal mol^{-1} and $k_2 = 1 \times 10^9$ M^{-1} s^{-1}.

The uncertainty in the standard potentials is probably ±100 mV, but may be more if some of the assumptions are not met. The potential correction for transforming reduction potentials to standard potentials was calculated for $\lambda = 30$ kcal mol^{-1} for all reactions. This might not be quite correct as the measurements here are not sufficiently sensitive to detect small differences in λ between the alkyl radicals. There is, however, no doubt that the self-exchange reorganisation energy for alkyl radicals (50 kcal mol^{-1}) is substantially higher than for benzylic radicals (15–30 kcal mol^{-1}). *ab initio* calculations[15] show that the alkyl anion is pyramidal with the charge localized whereas the benzylic anion is flat and the negative charge is spread throughout the whole π-system. The calculations show also that the bond length and the bond angles in the alkyl anion are altered compared with the alkyl radical, which give rise to a larger inner reorganisation energy, λ_i, for the alkyl compared with the benzyl system. According to the frequently used dielectric continuum model the solvent reorganization energy λ_o depends only on the total charge and the radius of the solvated molecule, which means that λ_o should be approximately equal for the alkyl and benzyl systems. However, in order to take the microstructure of the solvent shell into account, a more elaborate model is necessary to quantify the effect of charge distribution on
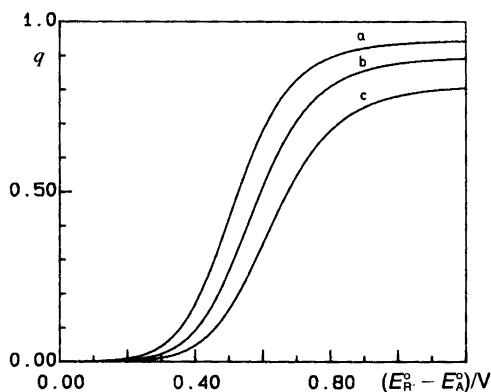
Fig. 2. q plotted versus the potential difference $E^{\circ}_{R} - E^{\circ}_{A}$ for three different values k_2/k_d with $\lambda = 30$ kcal mol^{-1}: a, $k_2/k_d = 0.05$; b, $k_2/k_d = 0.10$; c, $k_2/k_d = 0.20$.

the solvent reorganisation energy. Another problem is the contribution of entropy loss coming from the ordering of solvent molecules, which would be expected to lower λ_o for a molecule with delocalized charge compared with a molecule with localized charge as more solvent molecules are involved in the former case.

In Fig. 2 curves for q versus $E^{\circ}_{R} - E^{\circ}_{A}$ for $\lambda = 30$ kcal mol^{-1} are shown for different ratios of $k_2:k_d$. We have observed experimentally that q increases to at least 0.9 which rules out $k_2:k_d$ ratios higher than 0.1. The diffusion-controlled rate constant was assumed to be 1×10^{10} M^{-1} s^{-1} but the real value might be somewhat larger because the alkyl radicals are quite small. In Ref. 16 it was observed that k_d is nearly doubled on going from anthracene ($r = 3.9$ Å, $k_d = 3.9 \times 10^9$ M^{-1} s^{-1}) to naphthalene ($r = 3.5$ Å, $k_d = 7.7 \times 10^9$ M^{-1} s^{-1}). The radius of the *tert*-butyl radical[17] is between 2.0 and 2.6 Å so one could expect k_d to be somewhat higher than 1×10^{10} M^{-1} s^{-1}. The standard potential for the alkyl radicals will be shifted 70 mV to more negative values when the ratio k_2/k_d is halved.

The results in Table 2 show that primary alkyl radicals are easier to reduce than secondary alkyl radicals, which in turn are easier to reduce than tertiary alkyl radicals. This order of standard potentials is that expected as alkyl groups

Table 2. Reduction and standard potentials and reorganisation energies for different unconjugated alkyl radicals measured in DMF/TBABF at $T = 20\,°C$.

RX	$-E^q_i/V^a$	$-E^{\circ}/V^{a,b}$
Butyl chloride (1)	2.20	1.62
sec-Butyl chloride (2)	2.30	1.72
tert-Butyl chloride (3)	2.35	1.77
Chlorocyclohexane (4)	2.26	1.68
Chlorocyclopentane (5)	2.30	1.72
1-Chloro-2-methylpropane (6)	2.15	1.57
2-Chloropropane (7)	2.30	1.72

[a] versus SCE. [b] The E° was obtained by adding 580 mV to $E^q_{1/2}$. The value 580 mV was calculated for $\lambda = 30$ kcal mol^{-1} and $k_2 = 1 \times 10^9$ M^{-1} s^{-1}.

are generally considered to be electron donating. It is particularly gratifying to observe that the standard potentials seem to be divided into three distinct groups, one for primary alkyl radicals (1,6) between -1.57 and -1.62 V, one for secondary alkyl radicals (2,4,5,7) varying between -1.68 and -1.72 V and finally the tertiary alkyl radical (3) at -1.77 V vs. SCE. The values are slightly more negative than the standard potentials found by Savéant:[14] butyl radical (< -1.30 to -1.42), *sec*-butyl radical (-1.38 to -1.50) and *tert*-butyl radical (-1.48 to -1.60) V vs. SCE.

The concept of electrophilic and nucleophilic radicals is often used to explain the course of radical reactions. We expect the 'electrophilicity' or 'nucleophilicity' of a radical to be a function of the redox potentials of both the substrate and the radical, and a more thorough discussion can be made when more of these redox potentials are available.

How general is the method? The method includes three potentials which are characteristic for a given alkyl halide. These are the reduction potential of the alkyl halide, the reduction potential of the alkyl radical, and finally the standard potential of the alkyl radical. The method described here requires there to be some potential difference between $E^q_{1/2}$ and the reduction potential of the alkyl halide. This condition rules out, in most cases, halides other than the chlorides. Even with the chlorides there might be problems if λ exceeds 40 kcal mol^{-1} because the alkyl chloride is then reduced directly at $q < 0.9$. Another requirement of the method is that $E^q_{1/2}$ should not be more negative than -2.35 V (SCE) when medium discharge would begin to disturb the experiments.

The method outlined yields reduction potentials, $E^q_{1/2}$ with an estimated uncertainty of 50 mV, but the calculation of standard potentials is less reliable owing to an unsatisfactory determination of the reorganisation energy, λ. A better estimate of λ would improve the reliability of the standard potentials calculated from the reduction potentials.

Experimental

Materials. The mediators are the same as those used previously.[3–5] The alkyl chlorides were obtained commercially and distilled before use. The supporting electrolyte, Bu_4NBF_4. and the solvent, N,N-dimethylformamide were purified by standard procedures.

Equipment. Instrumentation and data treatment procedures have been described in detail elsewhere.[18] The working electrode in LSV was a mercury plated platinum disc electrode, $\varnothing = 0.6$ mm. The electrolyte solution was dried by passage through a column of activated alumina immediately before the experiments.

Procedures. Measurements of the competition parameter q by LSV were performed by the method described by Pedersen.[11]

The reaction between anthracene anion radical and butyl chloride was investigated by coulometry. Anthracene (0.15 g) was reduced in 50 ml of DMF/0.1 M TBABF$_4$ at a mercury pool electrode in an H-cell in the presence of 0.879 ml of butyl chloride. Prior to the experiment, the catholyte had been dried through a column of activated alumina. The charge consumed was measured, with a coulometer, to be 172 C ($n = 2.12$). The q value was calculated after the following considerations, $n = n_{coup} + n_{cat}$, $q = (n - 2)/n$. This formula is only valid when all the mediator has been reduced to the coupling product during the electrolysis. This means that the excess of the alkyl halide should always be larger than $q/(1 - q)$ to satisfy this condition.

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