Inorg. Chem. **2006**, *45*, 9181–9184

Inorganic Chemistry
Article

Iron-Only Hydrogenase Mimics. Thermodynamic Aspects of the Use of Electrochemistry to Evaluate Catalytic Efficiency for Hydrogen Generation

Greg A. N. Felton, Richard S. Glass, Dennis L. Lichtenberger, and Dennis H. Evans*

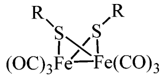
Department of Chemistry, The University of Arizona, Tucson, Arizona 85721

Received June 2, 2006

Voltammetry is widely used for the evaluation of iron-only hydrogenase mimics and other potential catalysts for hydrogen generation using various dipolar aprotic solvents. Effective catalysts show enhanced current in the presence of a proton donor at the potential where the catalyst is reduced. To facilitate the comparison of catalytic efficiencies, this paper provides a simple means of calculating the standard potential for reduction of the acid, HA, according to the half reaction $2HA + 2e^- \rightleftharpoons H_2 + 2A^-$. This standard potential depends on the pK_a of HA in the solvent being used. It is thermodynamically impossible for reduction of HA to occur at less negative potentials than the standard potential, and the most effective catalysts will operate at potentials as close as possible to the standard potential. In addition, direct reduction of HA at the electrode will compete with the catalyzed reduction, thus complicating evaluation of the rate of the catalyzed reaction. Glassy carbon electrodes, commonly used in such evaluations, show a quite large overpotential for direct reduction of HA so that the necessary corrections are small. However, catalysis at very negative potentials will be contaminated by significant direct reduction of HA at glassy carbon. It is demonstrated that direct reduction can be almost completely suppressed by using a mercury or amalgamated gold electrode, even at very negative potentials.

Introduction

The hydrogenases are very efficient catalysts for the reductive generation or oxidative uptake of dihydrogen.[1] The most efficient enzymes for hydrogen generation are the iron-only hydrogenases that contain a butterfly-shaped Fe_2S_2 core in the active site with ligands such as CO and CN^- that are unusual in biological systems. Many studies have been devoted to the synthesis and characterization of hydrogenase mimics containing this iron–sulfur core, as generalized in **1** (in which one or more CO may be replaced by CN^-, PR_3, or an *N*-heterocyclic carbene).[2] The most common method of evaluating the catalytic efficiency of the hydrogenase mimic is to obtain cyclic voltammograms of the potential catalyst in a nonaqueous solvent such as acetonitrile in the absence and in the presence of a source of protons. These proton sources have been variously perchloric acid,[2c,i] tetrafluoroboric acid,[2d,l] *p*-toluenesulfonic acid,[2h,n,o] trifluoromethanesulfonic acid,[2a] 2,4-dimethylpyridinium cation,[2e] and acetic acid.[2b,f,g,j,k,m]



1

Catalytic activity is indicated by a change in the peak(s) for reduction of the catalyst when the proton source is present. The changes include a growth in the height of the reduction peak and sometimes a small positive shift in its position. The best catalysts are taken to be those that produce the largest increase in peak height in the presence of acid and whose reduction potentials are not too negative, i.e., the catalysis occurs with minimal overpotential. The increase in peak height is interpreted as being due to a catalytic cycle that produces dihydrogen and the original oxidized form of the catalyst which is in turn reduced, giving more current.

The hydrogenases themselves are remarkably efficient. When immobilized on an electrode surface under the proper conditions, these natural catalysts can support the reversible, proton/dihydrogen reaction in aqueous media with low overpotential and at a diffusion-controlled rate.[3]

---

* To whom correspondence should be addressed. E-mail: dhevans@email.arizona.edu.

(1) Tye, J. W.; Hall, M. B.; Darensbourg, M. Y. *Proc. Nat. Acad. Sci. U.S.A.* **2005**, *102*, 16911–16912 and references therein.

10.1021/ic060984e CCC: $33.50

Two aspects of the electrochemistry of acids in dipolar aprotic solvents are of importance in obtaining realistic assessments of catalytic activity: knowledge of the standard potential for the reduction of the acid and knowledge of the contribution of the direct reduction of the acid at the electrode to the overall production of dihydrogen. These are the points to be addressed in this work.

Experimental Section

The source and treatment of solvent and electrolyte has been described.[4] $[(\eta^5\text{-}C_5H_5)Fe(CO)_2]_2$ (Aldrich) was purified by sublimation. Procedures and instrumentation for voltammetry have been reported.[4] Except for the hanging mercury drop electrode (EG&G Princeton Applied Research model 303A static mercury drop electrode), the electrodes were highly polished commercial disk electrodes.[4] The amalgamated gold electrode was prepared by dipping a gold disk electrode in mercury for 15−20 min followed by removal of excess mercury from the surface by manually dislodging the adhering drop of mercury. The potential of the silver reference electrode (Ag, 0.010 M $AgNO_3$, 0.10 M Bu_4NPF_6 in acetonitrile) was periodically measured vs the potential of the ferrocenium/ferrocene couple in acetonitrile, and all potentials are reported vs ferrocene.

Results and Discussion

Standard Potentials for the Reduction of Acids in Nonaqueous Solvents. The general acid, HA, enters into the following half-reaction with its conjugate base, A^-, and dihydrogen:

$$2HA(soln) + 2e^- \rightleftharpoons 2A^-(soln) + H_2(g) \quad E°_{HA} \qquad (1)$$

The standard potential for the solvated proton/dihydrogen couple is

$$2H^+(soln) + 2e^- \rightleftharpoons H_2(g) \quad E°_{H+} \qquad (2)$$

The acid dissociation constant of the acid, $K_{a,HA}$, corresponds to the reaction:

$$HA(soln) \rightleftharpoons A^-(soln) + H^+(soln) \quad K_{a,HA} \qquad (3)$$

Equations 1−3 can be combined to obtain $E°_{HA}$ as a function of $pK_{a,HA}$:

$$E°_{HA} = E°_{H+} - (2.303RT/F)pK_{a,HA} \qquad (4)$$

This expression is not exact when quite strong ($pK_a \approx 0$) acids are considered, but that complication has been included in a more general equation that was published earlier.[5]

The potential for the solvated proton/dihydrogen couple has been measured in a number of solvents ($E°_{H+}$, V vs the ferrocenium/ferrocene couple in the same solvent):[6a] nitromethane (−0.080), acetonitrile (−0.260), dimethoxyethane (−0.400), tetrahydrofuran (−0.440), *N*,*N*-dimethylformamide, DMF (−0.770), dimethyl sulfoxide (−0.830), and pyridine (−1.020). The value in acetonitrile was corrected for the incomplete dissociation of perchloric acid (pK_a = 2.1) giving −0.14 V.[6b]

The potential, *E*, of an active metal electrode (high-surface-area platinum, for example) with respect to the ferrocenenium/ferrocene couple in the solvent being used is given by the Nernst equation for reaction 1 and the activities, *a*, of HA and A^-, as well as the partial pressure of dihydrogen, $P°$ (eq 5). The potential will equal $E°_{HA}$ when the activities of HA and A^- are both unity and the partial pressure of hydrogen is 1 atm.

$$E = E°_{HA} + \frac{2.303RT}{2F}\log\frac{a_{HA}^2}{a_{A^-}^2 P^o} \qquad (5)$$

E, the reversible potential for reduction of HA, is the lowest potential at which HA can be reduced to dihydrogen and A^- with given pressure and activities.

What is the value of *E* under voltammetric conditions? Catalysts are usually evaluated using a certain concentration of HA, no added A^-, and no dihydrogen. These conditions are not well defined but the principles of voltammetry[7] can be used to make reasonable approximations. In general, when the current has reached approximately half its maximum value, the surface concentration of HA will have dropped to approximately one-half its bulk solution concentration and the surface concentration of A^- will have increased to about

---

(2) (a) Gao, W.; Liu, J.; Ma, C.; Weng, L.; Jin, K.; Chen, C.; Akermark, B.; Sun, L. *Inorg. Chim. Acta* **2006**, *359*, 1071−1080. (b) Song, L.-C.; Jea, C.; Yan, J.; Wang, H.-T.; Liu, X.-F.; Hu, Q,-M. *Organometallics* **2006**, *25*, 1544−1547. (c) Jiang, S.; Liu, J.; Sun, L. *Inorg. Chem. Commun.* **2006**, *9*, 290−292. (d) Capon, J.-F.; Gloaguen, F.; Schollhammer, P.; Talarmin, J. *Coord. Chem. Rev.* **2005**, *249*, 1664−1676. (e) Tard, C.; Liu, X.; Ibrahim, S. K. Bruschi, M.; De Giola, L.; Daview, S. C.; Yang, X.; Wang, L.-S.; Sawers, G.; Pickett, C. J. *Nature* **2004**, *433*, 610−613. (f) Song, L.-C.; Yang, Z.-Y.; Bian, H.-Z.; Liu, Y.; Wang, H.-T.; Liu, X.-F.; Hu, Q.-M. *Organometallics* **2005**, *24*, 6126−6135. (g) Tye, J. W.; Lee, J.; Wang, H.-W., Mejia-Rodriguez, R.; Reibenspies, J. H.; Hall, M. B.; Darensbourg, M. Y. *Inorg. Chem.* **2005**, *44*, 5550−5552. (h) Borg, S. J.; Behrsing, T.; Best, S. P.; Razavet, M.; Liu, X.; Pickett, C. J. *J. Am. Chem. Soc.* **2004**, *126*, 16988−16999. (i) Ott, S.; Kritikos, M.; ?kermark, B.; Sun, L.; Lomoth, R. *Angew. Chem., Int. Ed.* **2004**, *43*, 1006−1009. (j) Liu, T.; Wang, M.; Shi, Z.; Cui, H.; Dong, W.; Chen, J.; ?kermark, B.; Sun, L. *Chem. Eur. J.* **2004**, *10*, 4474−4479. (k) Mejia-Rodriguez, R.; Chong, D.; Reibenspies, J. H.; Soriaga, M. P.; Darensborg, M. Y. *J. Am. Chem. Soc.* **2004**, *126*, 12004−12014. (l) Capon, J.-F., Gloaguen, F.; Schollhammer, P.; Talarmin, J. *J. Electroanal. Chem.* **2003**, *566*, 241−247. (m) Chong, D.; Georgakaki, I. P.; Mejia-Rodriguez, R.; Sanabria-Chinchilla, J.; Soriaga, M. P.; Darensbourg, M. Y. *Dalton Trans.* **2003**, 4158−4163. (n) Gloaguen, F.; Lawrence, J. D.; Rauchfuss, T. B.; Bénard, M.; Rohmer, M.-M. *Inorg. Chem.* **2002**, *41*, 6573−6582. (o) Gloaguen, F.; Lawrence, J. D.; Rauchfuss, T. B. *J. Am. Chem. Soc.* **2001**, *123*, 9476−9477.

(3) (a) Vincent, K. A.; Armstrong, F. A. *Inorg. Chem.* **2005**, *44*, 798−809. (b) Vincent, K. A.; Cracknell, J. A.; Parkin, A.; Armstrong, F. A. *Dalton Trans.* **2005**, 3397−3403. (c) Lamle, S. E.; Vincent, K. A.; Halliwell, L. M.; Albracht, S. P. J.; Armstrong, F. A. *Dalton Trans.* **2003**, 4152−4157. (d) Léger, C.; Jones, A. K.; Roseboom, W.; Albracht, S. P. J.; Armstrong, F. A. *Biochemistry*, **2002**, *41*, 15736−15746.

(4) Macías-Ruvalcaba, N. A.; Evans, D. H. *J. Phys. Chem. B* **2005**, *109*, 14642−14647.

(5) Treimer, S. E.; Evans, D. H. *J. Electroanal. Chem.* **1998**, *449*, 39−48.

(6) (a) Daniele, S.; Ugo, P.; Mazzocchin, G.-A.; Bontempelli, G. *Anal. Chim. Acta* **1985**, *173*, 141−148. (b) For a more recent estimate of the potential, −0.05 V, based on early potentiometric measurements, see ref 6c (Supporting Information). (c) Ellis, W. W.; Raebiger, J. W.; Curtis, C. J.; Bruno, J. W.; DuBois, D. L. *J. Am. Chem. Soc.* **2004**, *126*, 2738−2743.

(7) Bard, A. J.; Faulkner, L. R. *Electrochemical Methods. Fundamentals and Applications*, 2nd ed.; Wiley: New York, 2001.

Table 1. Acid Dissociation Constants and Standard Potentials for the $HA/A^-,H_2$ Half Reaction for a Few Acids in Acetonitrile and *N*,*N*-Dimethylformamide[a]

	acetonitrile		*N*,*N*-dimethylformamide	
acid	pK_a	E°_{HA}	pK_a	E°_{HA}
perchloric acid	2.1	−0.26	dissociated	−0.77
trifluoromethanesulfonic acid	2.6	−0.29	dissociated	−0.77
p-toluenesulfonic acid	8.7	−0.65	2.6	−0.92
methanesulfonic acid	10.0	−0.73	3.0	−0.95
trichloroacetic acid	10.6	−0.77	3.5	−0.98
anilinium	10.7	−0.77	3.7	−0.99
pyridinium	12.3	−0.87	3.3	−0.96
dichloroacetic acid	13.2	−0.92	7.2	−1.20
chloroacetic acid	15.3	−1.05	10.0	−1.36
2,6-dinitrophenol	16.5	−1.12	6.1	−1.13
p-nitrobenzoic acid	18.5	−1.23	10.6	−1.40
triethylammonium	18.7	−1.25	9.2	−1.31
benzoic acid	20.7	−1.36	12.2	−1.49
p-nitrophenol	20.7	−1.36	12.2	−1.49
acetic acid	22.3	−1.46	13.2	−1.55
phenol	27.2	−1.75	>18	<−1.83
4-*tert*-butylphenol	27.5	−1.77	unavailable	

[a] E°_{HA} vs Fc^+/Fc calculated at 298 K from eq 4 with $E^\circ_{H+} = -0.14$ V (acetonitrile; calculated from ref 6a) and −0.77 V (DMF) (from ref 6a) and pK_a values from ref 10b. Fc^+/Fc refers to the potential of the ferrocenium/ferrocene couple as reference in the indicated solvent. Conversion to other reference electrodes can be achieved by the following: aqueous SCE is −0.40 V vs Fc^+/Fc in acetonitrile or DMF.[11] Aqueous NHE is −0.24 V vs aqueous SCE.

half the bulk concentration of HA, i.e., the surface concentrations of HA and A^- are equal, and thus, their activities (which are approximately equal to the molar concentrations) will cancel in eq 5. Also, the surface concentration of H_2 will increase and approach about half the surface concentration of the other product, A^-. However, due to the low solubility of H_2 (1.1, 2.8, 4.1, and 3.4 mM in dimethyl sulfoxide, chloroform, acetone, and tetrahydrofuran, for example;[8] 1.8 mM in DMF[9]), the solution near the electrode will be saturated or supersaturated with H_2 when catalytic reduction of a few millimolar HA occurs. This suggests that $P^\circ = 1$ atm is a good approximation. Thus, it is concluded that the least negative possible potential in the catalytic reduction of HA when the current is half its maximum value, $E_{half,HA}$, will be very close to the standard potential, E°_{HA}.

Equation 4 will allow us to compute E°_{HA} for any acid whose pK_a is known in a solvent for which E°_{H+} is known. Acid dissociation constants have been obtained for many acids in several different dipolar aprotic solvents.[10] As many evaluations of the catalytic efficiency of hydrogenase mimics have been conducted in acetonitrile, values of E°_{HA} have been calculated for a few acids in that solvent, and they are listed in Table 1. Results for DMF are also included.

These values of E°_{HA} can be used to evaluate the efficiency of a given catalyst. For example, acetic acid has been used as the proton source to evaluate catalytic activity in acetonitrile.[2b,f,g,j,k,m] Table 1 indicates that the least negative potential that will allow reduction of acetic acid to dihydrogen is −1.58 V vs Fc^+/Fc. It is thermodynamically impossible to achieve catalysis at a less negative potential.

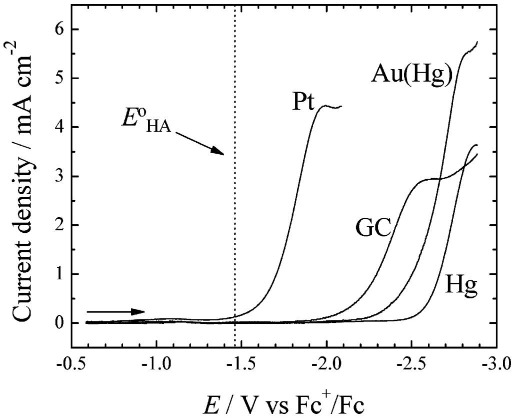


Figure 1. Linear sweep voltammograms for 10 mM acetic acid in acetonitrile with 0.10 M tetrabutylammonium hexafluorophosphate at 1.0 V/s. Electrodes as indicated (GC is glassy carbon). The electrode areas differed so the current density has been plotted. Reversible potential for reduction of acetic acid shown for comparison.

Suppression of the Direct Reduction of the Acid at the Electrode. In all cases reported to date, catalysis of dihydrogen generation by the iron-only hydrogenase mimics occurs with a large overpotential for all acids studied. This being the case, it is thermodynamically possible for the acid to be reduced directly at the electrode surface according to reaction 1 to give additional current and contribute to the amount of dihydrogen produced. Depending upon the nature of the electrode, this direct reduction can be very important, as illustrated in Figure 1. Voltammograms for reduction of 10 mM acetic acid in acetonitrile are shown for platinum, glassy carbon, amalgamated gold, and mercury electrodes. Platinum has the smallest overvoltage for the hydrogen evolution reaction in aqueous media, and it is clear that in acetonitrile as well, the reduction of acetic acid proceeds quite close to its reversible potential of −1.46 V vs ferrocene. The overpotential for reduction of acetic acid at glassy carbon is ~0.6 V larger than at platinum, making the direct reduction of little significance except for catalysis occurring at potentials more negative than about −2.2 V. The overpotential with the gold amalgam electrode is 0.2 V larger than with glassy carbon, and another 0.15 V increase is seen on going to a hanging mercury drop electrode.

For acids stronger than acetic, the reduction processes at each electrode are shifted toward less negative potentials roughly in accord with the shift in the standard potentials (Table 1).

The glassy carbon electrode is widely used for studies of hydrogenase mimics, and some, but not all, authors have explicitly considered the contributions of direct reduction[2e,h,i,k−n] by publishing voltammograms of acid alone or reporting the percent contribution of direct reduction to the total current. In some cases, additional current is seen with catalysts at very negative potentials and here it would be wise to consider

---

(8) Young, C. L., Ed. *I.U.P.A.C. Solubility Data Series*; Pergamon: New York, 1981; Vols. 5/6.

(9) Linke, W. F., Ed. *Solubilities. Inorganic and Metal-Organic Compounds*, 4th ed.; D. VanNostrand Co.: New York, 1958; Vol 1, p 1089.

(10) (a) Izutsu, K. *Acid−Base Dissociation Constants in Dipolar Aprotic Solvents*; Blackwell Scientific Publishers: Oxford, UK, 1990. (b) Izutsu, K. *Electrochemistry in Nonaqueous Solvents*: Wiley-VCH: Weinheim, 2002; Chapter 3. (c) Bordwell, F. G. *Acc. Chem. Res.* **1988**, *21*, 456−463.

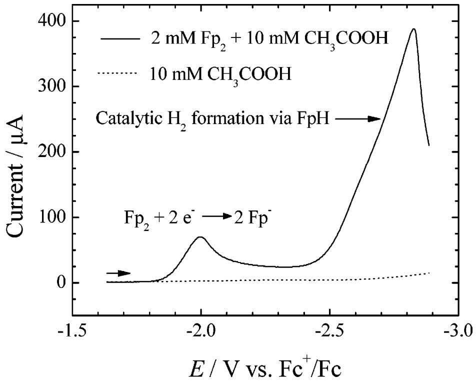


Figure 2. Linear sweep voltammograms for 2 mM $[(\eta^5\text{-}C_5H_5)Fe(CO)_2]_2$ (Fp$_2$) plus 10 mM acetic acid in acetonitrile with 0.10 M tetrabutylammonium hexafluorophosphate (solid curve) and 10 mM acetic acid alone (dashed curve). Hanging mercury drop electrode (area = 0.0096 cm^2). Scan rate: 1.0 V/s.

whether some or most of the additional current may arise from direct reduction rather than catalysis.

An alternative suggested by the present work is the use of mercury electrodes which will result in negligible direct reduction even past −2.5 V. Gold amalgam electrodes have a somewhat lower overvoltage for reduction of acid but are still superior to glassy carbon. A hanging mercury drop electrode was used for the reduction of triethylammonium ion to dihydrogen in DMF catalyzed by Fe(TTP)Cl (TTP = *meso*-tetraphenylporphyrin)[12] where scans were conducted to about −2.3 V without discernible direct reduction. The reaction occurs with ∼0.7 V overpotential based on the reversible potential for reduction of triethylammonium ion (Table 1).

Another example of the use of a hanging mercury drop electrode to suppress the direct reduction is shown in Figure 2. Here $[(\eta^5\text{-}C_5H_5)Fe(CO)_2]_2$ (Fp$_2$) is used as a pro-catalyst for the reduction of acetic acid in acetonitrile. Fp$_2$ is reduced to the anion Fp$^-$ at about −2.0 V,[13] a reaction that is followed by rapid protonation of Fp$^-$ to form the hydride, FpH,[14,15] which, as shown in separate experiments, is reduced near −2.5 V. This reduction initiates a catalytic cycle that efficiently produces dihydrogen, as inferred from the sharp increase in current upon addition of acetic acid.

The results in Figure 2 will be used to illustrate both of the principal points of this paper. First, the peak current for the catalyzed reduction of acetic acid occurs at about −2.8 V, whereas the standard potential for acetic acid is −1.46 V (Table 1; Figure 1). Thus, catalysis via FpH occurs with 1.3 V overpotential, pointing to the relative ineffectiveness of FpH as a catalyst. However, we have found that FpH also catalyzes the reduction of 4-*tert*-butylphenol, whose standard potential is −1.77 V (Table 1), also giving a catalytic peak near −2.8 V which indicates an overpotential of 1.0 V. This reduced overpotential comes at the cost of rate: the catalytic current is less than diffusion controlled due qualitatively to a lower rate of protonation of Fp$^-$ by the phenol. A mechanistic study of the action of the Fp$_2$/Fp$^-$/FpH system is underway.

The second point of the paper is that the direct reduction of the acid can be suppressed by the use of mercury or gold amalgam electrodes. This is illustrated by the voltammogram obtained for 10 mM acetic acid in the absence of Fp$_2$ (dashed, Figure 2) where only minor currents due to direct reduction are seen even at −2.9 V. This points to the value of these mercury-containing electrodes for studies of catalytic efficiency. Examination of Figure 1 shows that considerable direct reduction of acetic acid would have been seen had a glassy carbon electrode been used.

In summary, we emphasize two points that should be useful in the evaluation of catalysts for hydrogen generation. First, we provide a simple expression for computing the standard potential for reduction of any acid whose pK_a is known in the solvent being used. Second, we suggest that mercury or gold amalgam electrodes can be used to suppress the direct reduction of the acid on the electrode even at quite negative potentials.

Acknowledgment. The support of the National Science Foundation through the Collaborative Research in Chemistry program, Grant No. CHE 0527003, is gratefully acknowledged.

IC060984E

---

(11) Shalev, H.; Evans, D. H. *J. Am. Chem. Soc.* **1989**, *111*, 2667−2674.

(12) Bhugun, I.; Lexa, D.; Savéant, J.-M. *J. Am. Chem. Soc.* **1996**, *118*, 3982−3983.

(13) Fukuzumi, S.; Ohkubo, K.; Fujitsuka, M.; Ito, O.; Teichmann, M. C.; Maisonhaute, E.; Amatore, C. *Inorg. Chem.* **2001**, *40*, 1213−1219.

(14) Moore, E. J.; Sullivan, J. M.; Norton, J. R. *J. Am. Chem. Soc.* **1986**, *108*, 2257−2263.

(15) Shackleton, T. A.; Mackie, S. C.; Fergusson, S. B.; Johnston, L. J.; Baird, M. C. *Organometallics.* **1990**, *9*, 2248−2253.