Lewis acid activated oxidation of alkanes by barium ferrate

Chi-Ming Ho and Tai-Chu Lau*

Department of Biology and Chemistry, City University of Hong Kong, Tat Chee Avenue,
Hong Kong, P. R. China. E-mail: bhtclau@cityu.edu.hk; Fax: +852 2788 7406

Received (in Montpellier, France) 10th April 2000, Accepted 30th May 2000
Published on the Web 6th July 2000

In the presence of a few equivalents of a metal chloride, barium ferrate ($BaFeO_4$) has been found to oxidize cyclohexane at room temperature in acetic acid–dichloromethane to give a mixture of chlorocyclohexane, cyclohexanol and cyclohexanone. The rates of oxidation for the various metal chlorides follow the order $AlCl_3 > FeCl_3 > MgCl_2 > LiCl > ZnCl_2$. The best yield was obtained with $MgCl_2$, which represents a balance between reactivity and stability. Oxidation of other organic substrates has also been carried out using the $BaFeO_4$–LiCl system. Notably the system is able to oxidize propane and ethane to give a mixture of chloroalkanes and carbonyl products. The deuterium isotope effect for the oxidation of cyclohexane was found to be 2.1, 1.8, and 3.0 for chlorocyclohexane, cyclohexanol and cyclohexanone, respectively. The active intermediate is proposed to be a Lewis acid–ferrate adduct formed by coordination of an oxo ligand of the ferrate to the metal ion. It is suggested that the reactivity of this adduct is reminiscent of a radical species that oxidizes alkanes *via* a hydrogen atom abstraction pathway.

Potassium ferrate and other group 1 and group 2 metal ferrates have been known for a long time and are the most stable and well-defined terminal oxo species of iron. A detailed account of their preparations and properties has appeared recently.[1] The ferrate ion is thermodynamically a stronger oxidant than MnO_4^- and $Cr_2O_7^{2-}$; the E^0 for the Fe(VI)/Fe(III) couple have been estimated to be 2.20 and 0.72 V *vs.* NHE in acidic and basic solutions, eqns. (1) and (2) respectively[2]

$$FeO_4^{2-} + 8\,H^+ + 3\,e \rightarrow Fe^{3+} + 4\,H_2O \qquad (1)$$

$$FeO_4^{2-} + 4\,H_2O + 3\,e \rightarrow Fe(OH)_3 + 5\,OH^- \qquad (2)$$

It should thus be feasible to use metal ferrates to oxidize alkanes under mild conditions. The selective and efficient oxidation of alkanes continues to be a challenge to chemists.[3,4] Our interest in the ferrate ion also arises from its relevance to the enzymes cytochrome P-450 and methane monooxygenase, where it is generally believed that an iron-oxo species is the active intermediate.[5–7] Although there are a number of reports on the use of the ferrate ion for the oxidation of alcohols and amines,[8–11] there is only one study on the oxidation of alkanes, where it was found that cycloalkanes could be oxidized to the corresponding alcohols and ketones in 30–40% yield at 75 °C by K_2FeO_4 adsorbed on K10 montmorillonite clay.[1] Our strategy to activate the ferrate ion is to make use of Brønsted and Lewis acids, since recently we found that the reactivity of metal–oxo species such as RuO_4^{2-}, MnO_4^- and $Cr_2O_7^{2-}$ towards the oxidation of alkanes can be greatly enhanced by the addition of just a few equivalents of Brønsted or Lewis acids.[12–14]

Experimental

$BaFeO_4$ was prepared by a literature procedure.[9] The purity of the complex was determined to be 98 ± 1% using a titration method.[15] The sample was also found to be relatively free of carbonate and water from C,H analysis (C, 0.16%; H, 0.26%). Solvents were of reagent grade and further purified according to standard methods.[16] Anhydrous LiCl, Li(OTf), Li(OAc), $MgCl_2$, $ZnCl_2$, $FeCl_3$ and $AlCl_3$ were purchased from Aldrich and were used as received.

Gas chromatographic analyses were performed on a Hewlett–Packard 5890 series II gas chromatograph with an Ultra 2 (25 m × 0.2 mm i.d.) or FFAP (25 m × 0.2 mm i.d.) column. GC-MS measurements were carried out on an HP 5890 gas chromatograph interfaced to an HP 5970 mass selective detector.

Oxidation of alkanes by barium ferrate

In a typical experiment, 0.4 mmol of the metal chloride was added to a mixture containing 1.5 ml of dichloromethane, 0.5 ml of acetic acid, 0.5 ml of cyclohexane and 5 μl of chlorobenzene (internal standard). The mixture was stirred until all the Lewis acid dissolved, and 20 mg (0.078 mmol) of $BaFeO_4$ was than added. Aliquots were drawn at various time intervals and analyzed by GC and GC-MS. Elemental analyses were done on a Elementar Vario EL Analyser.

Results

$BaFeO_4$ is a stable, non-hygroscopic solid that can be readily prepared by the reaction of hypochlorite with iron(III) nitrate. It is insoluble in common organic solvents including CH_2Cl_2 or CH_3CN. Initial experiments indicated that when $BaFeO_4$ was added to a solution of cyclohexane in these solvents at room temperature, the purple solid remained undissolved and unchanged even after 24 h, and no products could be detected. It was then found that $BaFeO_4$ is slightly soluble in acetic acid to give a red solution, which appears to be stable for at least a few hours at room temperature since no color change was observed and no precipitate of $Fe(OH)_3$ occurred. However, when cyclohexane was added to this red solution no reaction occurred after more than 5 h at room temperature. As indicated in eqn. (1), the oxidizing power of ferrate, like most metal-oxo species, is enhanced by protons. Experiments were thus carried out by using a stronger acid, CF_3COOH. Upon addition of CF_3COOH (0.5 ml) to a suspension of $BaFeO_4$ in cyclohexane (0.5 ml) and CH_2Cl_2 (1.5 ml), the purple solid dissolved to give a brown solution within a few

DOI: 10.1039/b002907f

minutes, with the concomitant production of chlorocyclohexane (6%) and cyclohexanone (3%). In an attempt to improve the yield we turned to Lewis acids, since we have found earlier that Lewis acids such as metal ions and boron halides work just as well as protons in activating anionic metal-oxo species of ruthenium, chromium and manganese. Hence experiments using metal salts to replace CF_3COOH were carried out. Indeed addition of 5 equiv. of LiCl to $BaFeO_4$ and cyclohexane in CH_2Cl_2–CH_3COOH produced chlorocyclohexane, cyclohexanol and cyclohexanone within 2 h at room temperature with an overall yield of 42% (Table 1). A brown precipitate was also formed, which presumably was $Fe(OH)_3$. The formation of $Fe(OH)_3$ from the reduction of ferrate is well documented in the literature.[1,11] The acetic acid functions as a solvent for the metal chloride and the barium ferrate. In the absence of CH_3COOH, both the Lewis acid and the $BaFeO_4$ remained undissolved in CH_2Cl_2 and no products were found. Besides LiCl, activating effects were also observed with other metal chlorides. The rates of formation of products for the metal chlorides follow the order: $AlCl_3 > FeCl_3 > MgCl_2 > LiCl > ZnCl_2$. However, the best yield (54%) was obtained with $MgCl_2$. Surprisingly $BF_3 \cdot Et_2O$ gave very low yields although it is a rather strong Lewis acid. Moreover, the metal acetates and triflates were much less effective than the chlorides. However, PPh_4Cl is also ineffective, indicating that the activating effect is not due to Cl^- alone.

The Lewis acid–$BaFeO_4$ system was also able to oxidize other alkanes; again chloroalkanes, alcohols and ketones were produced and the results are tabulated in Table 2. LiCl was used rather than $MgCl_2$ because less chlorine-containing products were found. Notably the system was able to oxidize ethane and propane at room temperature to give predominantly chloroalkanes. Oxidation of propane occurred at both primary and secondary C–H bonds to produce 1-chloropropane, 2-chloropropane and propanone. In the oxidation of ethane, chloroethane and ethanal were detected; acetic acid was probably also produced, but we were not able to confirm this since the solvent used contained acetic acid. Although both alcohol and ketone products were produced in the oxidation of cyclohexanol, no alcohols were detected for

Table 1 Effect of Lewis acids on cyclohexane (CyH) oxidation by barium ferrate in CH_3COOH–CH_2Cl_2[a]

		Products (% yield)[b]		
Lewis acid	Time	Chlorocyclohexane	Cyclohexanol	Cyclohexanone
TFA[c]	20 min	[d]	6	3
$AlCl_3$	5 min	24	[d]	5
$Al(OTf)_3$	10 min	[d]	5	2
$FeCl_3$	5 min	14	2	1
$BF_3 \cdot Et_2O$	1.5 h	[d]	3	1
$ZnCl_2$	6 h	5	5	5
$MgCl_2$[e]	2 h	31	10	13
$Mg(OTf)_2$	2 h	[d]	3	4
LiCl	2 h	21	8	13
Li(OTf)	24 h	[d]	[d]	[d]
Li(OAc)	24 h		[d]	[d]
PPh_4Cl	24 h	<1	3	1

[a] Reaction conditions: temperature, 23 °C; cyclohexane, 0.5 ml; CH_2Cl_2, 1.5 ml; CH_3COOH, 0.5 ml; Lewis acid, 0.4 mmol; barium ferrate, 0.078 mmol. Mole ratio of Lewis acid to barium ferrate is approximately 5 : 1. [b] Mol of product/mol of barium ferrate × 100. [c] CF_3COOH (0.5 ml) was used instead of CH_3COOH. [d] Not detected. [e] CH_2Cl_2, 1 ml; acetic acid, 1 ml.

Table 2 Oxidation of organic substrates by $BaFeO_4$–LiCl in CH_2Cl_2–CH_3COOH[a]

Substrate (amount)	t/min	Product (% yield)[b]
Cyclohexane (0.5 ml)	120	Chlorocyclohexane (21), Cyclohexanol (8), Cyclohexanone (13)
Cyclohexane (0.5 ml)[c]	120	Chlorocyclohexane (29), Cyclohexanol (3), Cyclohexanone (5)
Cyclohexane (0.5 ml)[d]	120	Chlorocyclohexane (51), Cyclohexanol (7), Cyclohexanone (14)
n-Hexane (0.5 ml)	120	1-Chlorohexane (2), 2-Chlorohexane (7), 3-Chlorohexane (7), Hexan-2-one (6), Hexan-3-one (6)
Propane (2 atm)[e]	120	1-Chloropropane (10), 2-Chloropropane (19), Propanone (6)
Ethane (2 atm)[e]	120	Chloroethane (28), Ethanal (4)
Methylbenzene (0.5 ml)	90	Benzyl chloride (2), Benzaldehyde (6)
Ethylbenzene (0.5 ml)	90	1-Chloro-1-phenylethane (6), Acetophenone (13), 1-Phenylethanol (7)
Adamantane (50 mg)	120	1-Chloroadamantane (22), 2-Chloroadamantane (9), 1-Adamantanol (17), 2-Adamantanol (2), 2-Adamantanone (5)
Cyclohexene (0.5 ml)	90	*trans*-1,2-Dichlorocyclohexane (7), 3-Chlorocyclohexene (3), Cyclohexen-3-ol (10)
Styrene (0.5 ml)		Benzaldehyde (12)
trans-Stilbene (100 mg)		Benzaldehyde (17)
Cyclohexanol (40 μl)	90	Cyclohexanone (36)
Triphenylphosphine (80 mg)	20	Triphenylphosphine oxide (100)

[a] Reaction conditions: temperature, 23 °C; $BaFeO_4$, 0.078 mmol; CH_2Cl_2, 1.5 ml; CH_3COOH, 0.5 ml; LiCl, 0.4 mmol. Reaction were done in air, unless otherwise specified. [b] Based on amount of barium ferrate used. [c] Reaction under argon. [d] In the presence of 0.03 mmol Cl_2. [e] CH_2Cl_2 (2 ml).

the less easily oxidized substrates *n*-hexane, propane and ethane. It was found that cyclohexanol was readily oxidized to cyclohexanone, with a rate that is at least a few times faster than the oxidation of cyclohexane. Thus, for the less easily oxidized substrates, the initial production of alcohol should be much slower than its further oxidation to carbonyl compounds. No oxidation product arising from attack of the benzene ring was found when aromatic hydrocarbons were used as substrates, suggesting that involvement of hydroxyl radicals is unlikely. The oxidation of PPh_3 proceeded with the formation of PPh_3O in 100% yield, indicating that the system is capable of transferring one oxygen atom per mole of $BaFeO_4$. Oxidation of alkenes by the $BaFeO_4$–LiCl system gave allylic oxidation products and C=C bond cleavage products rather than epoxides. In the case of cyclohexene, *trans*-1,2-dichlorocyclohexane was also produced.

Discussion

Mechanism

The activation of a metal-oxo species through protonation of an oxo ligand is well established. Thus the effect of CF_3COOH on $BaFeO_4$ can be represented by eqn. (3):

$$[FeO_4]^{2-} + CF_3CO_2H \rightarrow [(O)_3Fe(OH)]^- + CF_3CO_2^- \quad (3)$$

The protonated ferrate species is thermodynamically a much stronger oxidant and is capable of oxidizing alkanes at room temperature. The low yields are probably due to the instability of $[(O)_3Fe(OH)]^-$. The activating effects of metal chlorides are most likely similar to that of protons, that is, an oxo ligand of the ferrate coordinates to the metal ion (which probably also has one or more chloro ligands):

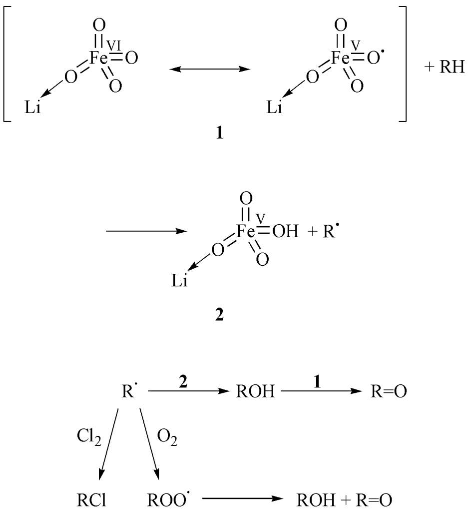
$$FeO_4^{2-} + M^{n+} \rightarrow [(O)_3Fe{=}O{\rightarrow}M]^{(n-2)+} \quad (4)$$

The withdrawal of electron density from the oxo ligand by the metal ion acting as a Lewis acid makes the oxo-iron species more oxidizing. The rates of formation of products for the hard Lewis acids increase in the order $LiCl < MgCl_2 < FeCl_3 < AlCl_3$, in accordance with the increase in acidity from Li^+ to Al^{3+}. However, for $FeCl_3$ and $AlCl_3$, although the rates are fast, the yields are relatively low; presumably the ferrate adducts of these Lewis acids were very unstable and readily decomposed to iron(III) oxide. The best yields were obtained from $MgCl_2$, presumably because this represents a balance between reactivity and stability. The rates and yields with $ZnCl_2$ were low, probably because of the low affinity of the soft Zn^{2+} for the hard oxo ligand. The acetate and triflate salts were less effective than the chloride salts; this is not unexpected since the oxygen atoms on these anions should compete effectively with the oxo ligand of the ferrate for the metal ions. Attempts to isolate or detect the proposed Lewis acid–ferrate adducts have so far been unsuccessful due to their instability. However, adducts of less-oxidizing metal-oxo species with Lewis acids are known in the literature. For example, the complexes $[(ReMe_4O)_2Mg(THF)_4]$ and $[\{Re(CH_2SiMe_3)_4O\}_2Mg(THF)_2]$ have been prepared and their X-ray crystal structures showed the bonding of oxo ligands to Mg^{2+}.[17]

A proposed mechanism for the oxidation of alkanes by the Lewis acid–ferrate adduct is shown in Scheme 1. The first step of the mechanism is the formation of alkyl radicals, this is followed by their conversion to chlorinated and oxygenated products.

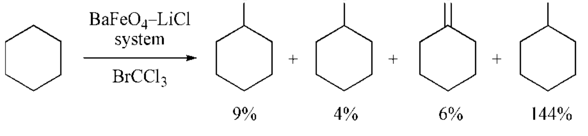
Formation of alkyl radicals

Kinetic isotope effects ($KIE = k_H/k_D$) were determined from the competitive oxidation of cyclohexane and d_{12}-cyclohexane by the $BaFeO_4$–LiCl system. The KIE values for chlorocyclohexane, cyclohexanol and cyclohexanone are 2.1, 1.8 and 3.0, respectively. For the $BaFeO_4$–TFA system, the KIE values for cyclohexanol and cyclohexanone are 2.6 and 3.6, respectively. Similar KIE values were also obtained for cytochrome P-450 catalyzed hydrocarbon hydroxylation when the rates of product formation from deuterated and undeuterated substrates were compared, although the intrinsic isotopic effects are actually much higher.[5] Small KIE values were also observed for the oxidation of cyclohexane by strongly oxidizing metal-oxo species; for example, the KIE for the formation of cyclohexanone by $(Bu_4N)_2(Cr_2O_7)$–BF_3,[13] $KMnO_4$–BF_3[13] and CrO_2Cl_2[18] are 3.9, 4.1 and 2.5, respectively. The rather small values of deuterium isotope effects for the $BaFeO_4$–LiCl system are also in the range expected for abstraction of hydrogen by radical reagents.[19] This suggests that the reactivity of the Lewis acid–ferrate adduct is reminiscent of a radical species that oxidizes alkanes *via* a hydrogen atom abstraction pathway (Scheme 1).



Scheme 1

Oxidation of adamantane occurs predominantly at the tertiary positions, the C^2/C^3 product ratio is 0.41, which is consistent with a radical intermediate. The presence of an alkyl radical intermediate is also supported by the use of a radical trap. Oxidation of cyclohexane by the $BaFeO_4$–LiCl system (under the conditions shown in Table 2) in the presence of 0.1 ml $BrCCl_3$ produced a substantial amount of bromocyclohexane in the dark [eqn. (5)]. The large yield of bromocyclohexane suggests that a radical chain reaction involving trichloromethyl radicals occurs.



(5)

Formation of chlorinated products

It was found that the yield of chlorocyclohexane as well as cyclohexanol and cyclohexanone remained roughly the same when 2 mol equiv. of PPh_4Cl was added as an external chloride source to the $BaFeO_4$–LiCl–cyclohexane reaction. This suggests that the cyclohexyl radical was not further oxidized to a carbocation, which would react rapidly with Cl^- to produce chloroalkane. It is also known that alkyl radicals react rapidly with oxidizing metal halides such as $CuCl_2$ and $FeCl_3$ to produce cyclohexyl halide with the concomitant reduction of the metal center.[20] In the present system, this pathway can also be excluded since the metal chlorides used are not redox active. Since FeO_4^{2-} is a very strong oxidant, it is reasonable to expect that it will oxidize chloride to chlorine, which would then react rapidly with cyclohexyl radicals to produce chlorocyclohexane. The presence of chlorine in the system was tested by addition of 0.1 mmol of cyclohexene to the $BaFeO_4$–LiCl–cyclohexane reaction. The products consisted of 3% 3-chlorocyclohexene, 9% *trans*-2-chlorocyclohexanol, 6% *trans*-1,2-dichlorocyclohexane and 11% chlorocyclohexane + cyclohexanol + cyclohexanone. The presence of chlorine in the reaction was indicated by the formation of *trans*-1,2-dichlorocyclohexane through the addition reaction of cyclohexene with chlorine. If chlorocyclohexane was produced form chlorine, then addition of external chlorine to the system should increase its yield. Indeed, when chlorine was added, the yield of chlorocyclohexane was greatly increased, but the yields and distribution of the oxygenated products were not affected (Table 2). This suggests that the increase in yield of chlorocyclohexane is mainly due to a chain reaction involving the chlorine radical, which results in the production of more cyclohexyl radicals:

$$R\cdot + Cl_2 \rightarrow RCl + Cl\cdot \tag{6}$$

$$Cl\cdot + RH \rightarrow HCl + R\cdot \tag{7}$$

Formation of oxygenated products

An oxygen rebound mechanism is proposed for the formation of alcohol from alkyl radicals and species **2**, *i.e.*

$$Fe^{V}\text{–}OH + R\cdot \rightarrow Fe^{IV} + ROH \tag{8}$$

The alcohol is further oxidized to ketone by **1** (and possibly **2** as well):

$$ROH + \mathbf{1} \rightarrow R{=}O + Fe^{IV} \tag{9}$$

It is not clear how the Fe^{IV} species is ultimately reduced to iron(III) oxide; it could participate in the oxidation of the alkyl radical or undergo disproportionation. Another possible pathway for the formation of oxygenated products from alkyl radicals involves the participation of dioxygen. It is well known that alkyl radicals react with dioxygen at diffusion controlled rates to produce alkylperoxyl radicals, which then give rise to ketone and alcohol in a 1 : 1 ratio.[19] In order to examine the effects of dioxygen a number of experiments on the oxidation of cyclohexane were carried out under argon, and an increase in the yield of chlorocyclohexane together with a significant decrease in the yields of oxygenated products were observed compared with the reactions done in air. This suggests that both species **2** and O_2 are responsible for producing oxygenated products from $R\cdot$. The increase in yield of chlorocyclohexane under argon is a result of the removal of dioxygen, which is a competitor of chlorine for cyclohexyl radicals in the reaction.

In conclusion, metal chlorides readily activate barium ferrate towards the oxidation of alkanes at room temperature. Similar activating effects have been observed with anionic oxo species of ruthenium, chromium and manganese. It is believed that the metal ions function as Lewis acids, and it is remarkable that in most cases just a few equivalents of rather mild Lewis acids is enough to produce a large activating effect.

Acknowledgements

We thank the Hong Kong Research Grants Council (CERG 9040197) and the City University of Hong Kong (DAG 7100042) for financial support.

References

1 L. Delaude and P. Laszlo, *J. Org. Chem.*, 1996, **61**, 6360.
2 R. H. Wood, *J. Am. Chem. Soc.*, 1958, **80**, 2038.
3 *Selective Hydrocarbon Activation, Principles and Progress*, eds. J. A. Davis, P. L. Watson, J. F. Liebman, A. Greenberg, VCH, New York, 1990.
4 M. Hudlicky, *Oxidation in Organic Chemistry*, ACS monograph, American Chemical Society, Washington, DC, 1990.
5 *Cytochrome P-450: Structure, Mechanism and Biochemistry*, ed. P. R. Ortiz de Montelllano, Plenum Press, New York, 1986.
6 R. H. Holm, P. Kennepohl and E. I. Solomon, *Chem. Rev.*, 1996, **96**, 2607.
7 B. J. Wallar and J. D. Lipscomb, *Chem. Rev.*, 1996, **96**, 2625.
8 R. J. Audette, J. W. Quail and P. J. Smith, *Tetrahedron Lett.*, 1971, 279.
9 R. J. Audette and J. W. Quail, *Inorg. Chem.*, 1972, **11**, 1904.
10 R. J. Audette, J. W. Quail and P. J. Smith, *J. Chem. Soc., Chem. Commun.*, 1972, 38.
11 Y. Tsuda and S. Nakajima, *Chem. Lett.*, 1978, 1397.
12 T. C. Lau and C. K. Mak, *J. Chem. Soc., Chem. Commum.*, 1993, 766.
13 T. C. Lau, Z. B. Wu, Z. L. Bai and C. K. Mak, *J. Chem. Soc., Dalton Trans.*, 1995, 695.
14 T. C. Lau and C. K. Mak, *J. Chem. Soc., Chem. Commum.*, 1995, 695.
15 J. M. Schreyer, G. W. Thompson and L. T. Ockermain, *Anal. Chem.*, 1950, **22**, 1426.
16 D. D. Perrin and W. L. F. Armarego, *Purification of Laboratory Chemicals*, 3rd edn., Pergamon Press, Oxford, 1988.
17 P. Stravropoulos, P. G. Edwards, G. Wilkinson, M. Motevalli, K. M. A. Malik and M. B. Hursthouse, *J. Chem. Soc., Chem. Commun.*, 1981, 1039.
18 G. K. Cook and J. M. Mayer, *J. Am. Chem. Soc.*, 1994, **116**, 1855.
19 G. A. Russel, *J. Am. Chem. Soc.*, 1957, **79**, 3871.
20 J. K. Kochi, *Acc. Chem. Res.*, 1974, **7**, 351.