Baeyer–Villiger Oxidation of Cyclic Ketones with Hydrogen Peroxide Catalyzed by Cationic Complexes of Platinum(II): Selectivity Properties and Mechanistic Studies

Marino Del Todesco Frisone, Francesco Pinna, and Giorgio Strukul*

Department of Chemistry, University of Venice, Dorsoduro 2137, 30123 Venezia, Italy

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The catalytic oxidation of simple cyclic ketones with hydrogen peroxide to give the corresponding lactones is reported. The reaction is catalyzed by complexes of Pt(II) of the type $[(P\text{-}P)Pt(CF_3)(solv)]^+$ (P-P = diphosphine) that may be deactivated by the hydroxy acids formed by hydrolysis of the lactones. The selectivity of the catalyst is studied in the oxidation of substrates like camphor, 2-cyclohexene-1-one, menthone, carvone, and indanones. Cyclobutanone is used to determine the mechanism of the reaction from initial rate studies. The reaction scheme proposed, which accounts for the observed effects of the various reactants, involves the coordination of the ketone on the vacant coordination site of the complex followed by nucleophilic attack of free hydrogen peroxide on the carbonyl carbon. The involvement of a quasi-peroxymetallacyclic intermediate is suggested which rearranges to give the lactone and the starting complex. A comparison with the mechanistic behavior of organic peroxy acids is given.

Introduction

Lactones are important industrial intermediates that find application in the production of polymers, pharmaceuticals, and herbicides and as solvents.[1] An easy and convenient route to this class of compounds involves the Baeyer–Villiger oxidation of the corresponding ketones employing organic peroxy acids as oxidants.[2] Although some of the latter are commercially available on a large scale, their industrial use is impractical, requiring necessarily the separation of the reaction products from the reduced organic acid that, in turn, has to be recovered and sent to a separate production unit in order to regenerate the peroxy acid. For these reasons the use of low cost oxidants associated with a catalyst has long been the subject of both industrial and academic research.

Hydrogen peroxide has a number of advantages compared to other easily accessible oxidants, namely, low cost, high active oxygen content, minimal byproduct formation, and innocuous nature of the reduction product (water).[3] However, its use as primary oxidant for the Baeyer–Villiger oxidation reaction has been only occasionally reported in the literature associated with As,[4] Mo,[5] or Se[6] compounds as catalysts, only the latter being of some synthetic utility. Of these, both As and Se derivatives behave like organic peroxy acids, while the Mo peroxo complexes $[MoO(O_2)_2(pic)]^-$ or $[MoO(O_2)(dipic)]$ (pic = picolinato; dipic = pyridine-2,6-dicarboxylato) were found to catalyze the lactonization of a variety of simple cyclic ketones at 60 °C but using 90% H_2O_2 as primary oxidant.[5] Both the high concentration of the oxidant employed (above the detonation limit) and the moderate to low yields in lactones obtained seem to reflect the difficulty to make interact two putative electrophiles i.e. the carbonyl carbon and one of the peroxygens (Scheme I).

On the other hand, it has long been known that the platinum dioxygen complex $(PPh_3)_2Pt(O_2)$ interacts easily with ketones to give stable peroxymetallacyclic insertion products,[7] where the driving force of the reaction has been recognized to be the nucleophilic attack of the peroxo moiety on the electrophilic carbonyl carbon[7] (Scheme I). Unfortunately, in this case the insertion complex is very stable and no oxidation products evolve. However, by avoiding complete cyclization and/or providing a good leaving group, it is possible to force Pt-based systems to work catalytically. In fact, in a preliminary account[8] we have observed that $(P\text{-}P)Pt(CF_3)X$ complexes (P-P = various diphosphines; X = solvent, –OH) that are known to form the corresponding PtOOH complexes by interaction with H_2O_2[9] can catalyze the conversion of a variety of simple ketones to the corresponding lactones. We now

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[1] *Kirk-Othmer Encyclopedia of Chemical Technology*, 3rd ed.; Wiley-Interscience: New York, 1980–84: Vol. 1, p 129; Vol. 13, p 97; Vol. 14, p 313; Vol. 14, p 780; Vol. 15, p 770; Vol. 18, p 361; Vol. 23, p 585.

[2] See for example: (a) Sutherland, I. O. In *Comprehensive Organic Chemistry*; Sutherland, I. O., Ed.; Pergamon: Oxford, U.K., 1979; Vol. 2, Chapter 98, p 894. (b) House, H. O. *Modern Synthetic Reactions*, 2nd ed.; Benjamin: Menlo Park, CA, 1972; Chapter 6, p 321. (c) Plesnicar, B. In *Oxidation in Organic Chemistry*; Trahanovsky, W. S., Ed.; Academic: New York, 1978; Part C, p 211. (d) Hassall, C. H. *Org. React.* **1957**, *9*, 53 and references cited therein.

[3] (a) Sheldon, R. A. *Studies on Surface Science and Catalysis*; Centi, G., Trifirò, F., Eds.; Elsevier: Amsterdam, 1991; Vol. 55, p 1. (b) Sheldon, R. A. *CHEMTECH* **1991**, 556. (c) Strukul, G. In *Catalytic Oxidations with Hydrogen Peroxide as Oxidant*; Strukul, G., Ed.; Kluwer: Dordrecht, The Netherlands, 1992, in press.

[4] Jacobson, S. E.; Mares, F.; Zambri, P. M. *J. Am. Chem. Soc.* **1979**, *101*, 6938.

[5] Jacobson, S. E.; Tang, R.; Mares, F. *J. Chem. Soc., Chem. Commun.* **1978**, 888.

[6] (a) Grieco, P. A.; Yokoyama, Y.; Gilman, S.; Ohfune, F. *J. Chem. Soc., Chem. Commun.* 1977, 870. (b) Syper, L. *Synthesis* **1989**, 167 and references therein cited.

[7] (a) Ugo, R.; Conti, F.; Cenini, S.; Mason, R. *J. Chem. Soc., Chem. Commun.* 1968, 1948. (b) Hayward, P. J.; Blake, D. M.; Wilkinson, G.; Nyman, C. J. *J. Am. Chem. Soc.* **1970**, *92*, 5873. (c) Ugo, R.; Zanderighi, G. M.; Fusi, A.; Carreri, D. *J. Am. Chem. Soc.* **1980**, *102*, 3745 and references therein.

[8] Del Todesco Frisone, M.; Giovanetti, R.; Pinna, F.; Strukul, G. In *Studies on Surface Science and Catalysis*; Simandi, L., Ed.; Elsevier: Amsterdam, 1991; Vol. 66, p 405.

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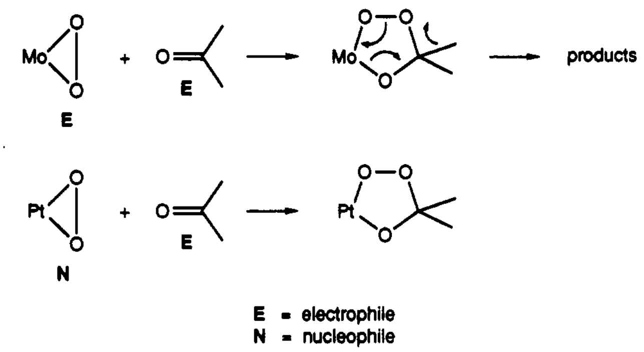
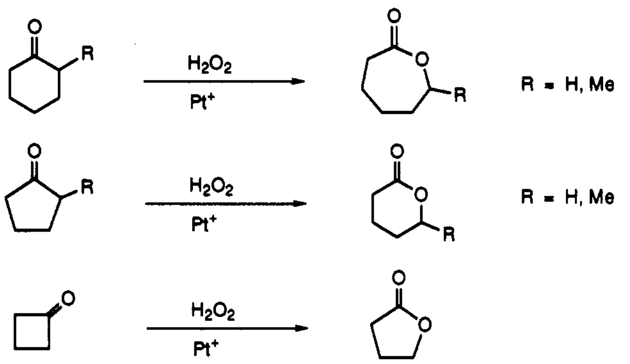
Scheme I



Scheme II

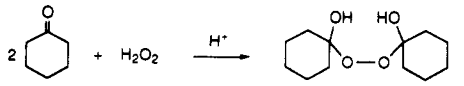


wish to report a more thorough description of this system with the aim of widening its scope toward both simple and more complex substrates and clarifying the mechanistic details of the reaction in the oxidation of a prototype compound (cyclobutanone).

Results and Discussion

Oxidation of Simple Ketones. Typically,[8] the (dppe)-$Pt(CF_3)(CH_2Cl_2)^+$ complex (dppe = 1,2-bis-diphenylphosphinoethane) catalyzes the lactonization of simple cyclic ketones (Scheme II) with good efficiency at room temperature and using commercial 35% solutions of hydrogen peroxide. Basic features of these reactions are as follows: (i) the carbon migratory aptitude follows the order already known for the classical Baeyer–Villiger oxidation with peracids;[2] (ii) the use of hydroxo complexes like (dppe)-$Pt(CF_3)(OH)$ reduces the activity of about 10 times; (iii) hydrolysis of lactones to the corresponding hydroxy acids is in some cases induced by the acidity of the commercial H_2O_2 solutions (~pH 3) and/or the protons liberated in the reaction medium by eq 1; (iv) the same acidity can catalyze the addition of H_2O_2 to the ketone (this is particularly evident with cyclohexanone) giving 1,1′-dihydroxy diperoxide (eq 2).

$$Pt^+ + H_2O \rightleftharpoons Pt\text{-}OH + H^+ \qquad (1)$$

 (2)

Feature iv above constitutes a major limitation in achieving higher yields of ϵ-caprolactone that is industrially the most valuable product. To see whether the problems related to the acidity of the system could be solved, the reaction was tested using H_2O_2 solutions buffered at different pH. Even at pH 5 byproduct formation is completely suppressed and the catalytic activity decreases with increasing pH, but as can be seen from Figure 1, compared to the unbuffered solution, the catalytic activity drops off almost 2 orders of magnitude.

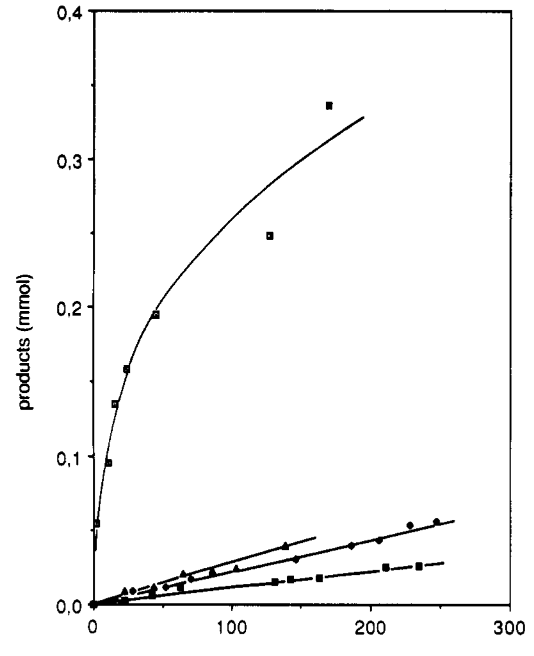


Figure 1. Lactonization of cyclohexanone catalyzed by $[(dppe)Pt(CF_3)(CH_2Cl_2)]^+$ at different pH values: Open squares (ordinate scale × 10), original pH (~pH 3, no buffer); open triangles, pH 5; filled diamonds, pH 6; filled squares, pH 7. Amounts: Pt, 0.045 mmol; cyclohexanone, 45 mmol; H_2O_2, 22.5 mmol.

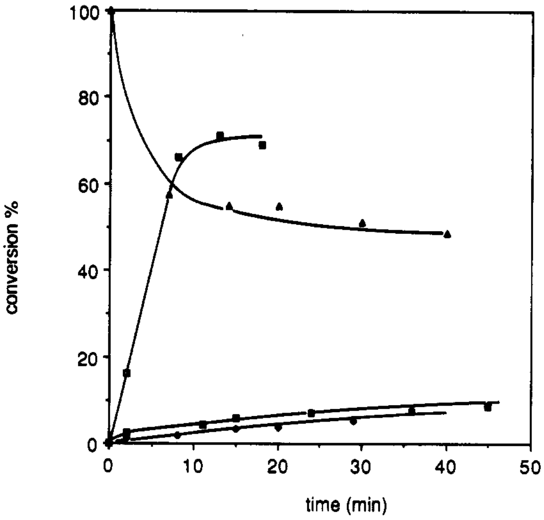


Figure 2. Comparison among the lactonizations of cyclobutanone (open squares), cyclopentanone (filled diamonds), and cyclohexanone (filled squares). Conversion refers to hydrogen peroxide. The parallel consumption of hydrogen peroxide is shown (open triangles) for the oxidation of cyclopentanone. Amounts: Pt, 0.045 mmol; ketone, 45 mmol; H_2O_2, 22.5 mmol.

The oxidations of cyclobutanone, cyclopentanone, and cyclohexanone have been compared under the same experimental conditions. As can be seen from Figure 2 the initial rate and the yield in lactone increase very much on going from 6- and 5- to 4-membered ring ketones and seem to reflect the ease with which ring expansion occurs. Under identical experimental conditions but in the absence

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(9) (a) Zanardo, A.; Pinna, F.; Michelin, R. A.; Strukul, G. *Inorg. Chem.* **1988**, *27*, 1966. (b) Zanardo, A.; Michelin, R. A.; Pinna, F.; Strukul, G. *Inorg. Chem.* **1989**, *28*, 1648.

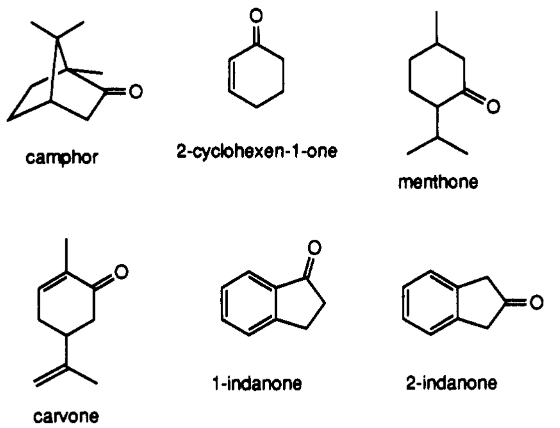
of catalyst, conversion was <1% in the case of cyclobutanone and <0.1% for cyclopentanone and cyclohexanone. Effects of this type using peracids as oxidants are not novel in the literature[5,10] and have been interpreted similarly, although a direct comparison among the three ketones can be found only in the case of polystyrenearsonates.[4] What is evident from Figure 2 is also that the amount of H_2O_2 that is actually utilized to give products depends largely on the ketone. In fact only in the case of cyclobutanone the conversion with respect to the H_2O_2 introduced is close (70%) to the complete consumption, while with cyclopentanone and cyclohexanone it is around 10%. This tendency to convert into products only a (minor) part of the H_2O_2 employed is common to other systems.[4]

However, the amount of H_2O_2 still present in the reaction mixture never balances the reaction products formed, as is shown for example in Figure 2 for the case of cyclopentanone where the parallel reaction profile for the H_2O_2 consumption is also reported. In other words there is both a partial decomposition of the oxidant and a deactivation of the catalyst.

As to the former, the possible contribution of the catalyst and the substrate trace impurities to a radical type decomposition via Haber–Weiss chemistry was checked in the case of cyclopentanone, by running blank reactions. Under the experimental conditions of Figure 2, it was found that the catalyst contributes to a maximum 5% decomposition, while the substrate to less than 3%. Consequently, the largest part of the decomposition comes from a different pathway, and a possible explanation will be given later (vide infra).

A possible origin of catalyst deactivation might be associated to the observed formation of small amounts of hydroxy acids by hydrolysis of the corresponding lactones (see feature iii above). In fact, it is known that these Pt(II) complexes give easily condensation reactions with several types of weak acids.[11] To check this possibility, the catalyst $[(dppe)Pt(CF_3)(CH_2Cl_2)]ClO_4$ was precipitated from solution with Et_2O in the oxidation of cyclohexanone after product formation had virtually stopped. The solid obtained was analyzed spectroscopically. IR analysis in Nujol mull showed a medium-intensity band at 1720 cm^{-1} typical of free cyclohexanone and two strong bands at 1540 and 1185 cm^{-1} that could be attributed to the asymmetric and symmetric stretchings of the carboxylate anion.[12] Similar bands were observed also in dimeric Pt(II) complexes of the type $P_2(CF_3)Pt(\mu\text{-}OCO)Pt(CF_3)P_2$.[13] The same solid analyzed in CH_2Cl_2 solution showed the same carboxylate bands at 1540 and 1185 cm^{-1}, the absence of the band at 1720 cm^{-1}, and the presence of a weak-intensity band at 3610 cm^{-1} typical of the complex $(dppe)Pt(CF_3)(OH)$.[14] These data suggest that the solid may be constituted, at least in part, of an hexanoate complex arising from reaction of the starting complex with hydroxyhexanoic acid according to eqs 1 and 3.

Chart I



$$(dppe)Pt(CF_3)(OH) + HOOC(CH_2)_5OH \rightarrow (dppe)Pt(CF_3)(OOC(CH_2)_5OH) + H_2O \qquad (3)$$

The ${}^{19}F$ NMR spectrum (CD_2Cl_2) confirmed the results of the IR analysis. In fact it showed the existence of a mixture of products where two main signals can be recognized, one centered at δ −27.66 ppm (${}^3J_{FP_{cis}}$ 9.7 Hz, ${}^3J_{FP_{trans}}$ 57.0 Hz) and one at δ −28.25 ppm (${}^3J_{FP_{cis}}$ 9.5 Hz, ${}^3J_{FP_{trans}}$ 56.4 Hz), both with evidence of Pt couplings even if these can hardly be calculated because of extensive overlapping. Of these, the latter can again be assigned to $(dppe)Pt(CF_3)(OH)$,[11] while the former may be attributed to the hexanoate derivative.

To confirm these data, a hexanoate complex was prepared according to eq 3 using hexanoic acid, and the spectroscopic characterization showed IR bands at 1540 and 1185 cm^{-1}, while the ${}^{19}F$ NMR spectrum showed a doublet of doublets at δ −27.88 ppm (${}^3J_{FP_{cis}}$ 9.9 Hz, ${}^3J_{FP_{trans}}$ 57.9 Hz, ${}^2J_{FPt}$ 568 Hz).

At the end of all the above reactions the catalyst can be precipitated with Et_2O, filtered out, dissolved in CH_2Cl_2, and treated with an aqueous solution of $HClO_4$ to yield again $[(dppe)Pt(CF_3)(CH_2Cl_2)]ClO_4$.

Reactivity of More Complex Ketones. We have extended the study of the catalytic properties of the Pt^+/H_2O_2 system to the case of the ketones indicated in Chart I. As can be seen, some of these have also different functionalities and allow the study of the selectivity properties of the catalyst. For a comparison standard reactions using *m*-chloroperbenzoic acid (MCPBA) have also been carried out. In all these reactions the catalyst used was $[(dppe)Pt(CF_3)(CH_2Cl_2)]ClO_4$ and, where possible, reactions were carried out in the absence of solvent. A summary of the results obtained is reported in Table I.

a. Camphor. As can be seen from Table I, camphor produces two different lactones: one corresponding to the migration of the tertiary carbon atom (2) and the other from the migration of the primary carbon atom (1). The reaction profile is similar to that already observed for simple ketones (see for example Figure 1) with a rather fast product formation at the beginning followed by a slow down for longer times. As can be seen from Table I, the selectivity of the system favors product 1 in agreement with with what is observed with MCPBA but at variance to what should be expected from the carbon migratory aptitude known for this reaction.[2] The generally accepted

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[10] (a) Friess, S. L. *J. Am. Chem. Soc.* **1949**, *71*, 2571. (b) Friess, S. L.; Frankenburg, P. E. *J. Am. Chem. Soc.* **1952**, *74*, 2679. (c) Mateos, J. L.; Menchaca, H. *J. Org. Chem.* **1964**, *29*, 2026. (d) Starcher, P. S.; Phillips, B. *J. Am. Chem. Soc.* **1958**, *80*, 4079. (e) Johnson, P. Y.; Yee, J. *J. Org. Chem.* **1972**, *37*, 1058. (f) Bogdanowicz, J.; Ambelang, T.; Trost, B. M. *Tetrahedron Lett.* **1973**, 923.

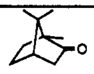
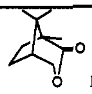
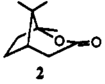
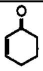
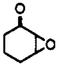
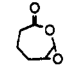
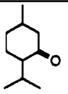
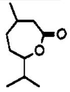
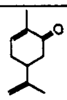
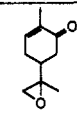
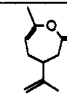
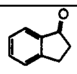
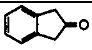
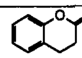
[11] Michelin, R. A.; Napoli, M.; Ros, R. *J. Organomet. Chem.* **1979**, *175*, 239.

[12] Silverstein, R. M.; Bassler, C. G.; Morrill, T. C. *Spectrometric Identification of Organic Compounds*, 3rd ed.; Wiley: New York, 1974; p 101.

[13] Torresan, I.; Michelin, R. A.; Marsella, A.; Zanardo, A.; Pinna, F.; Strukul, G. *Organometallics* **1991**, *10*, 623.

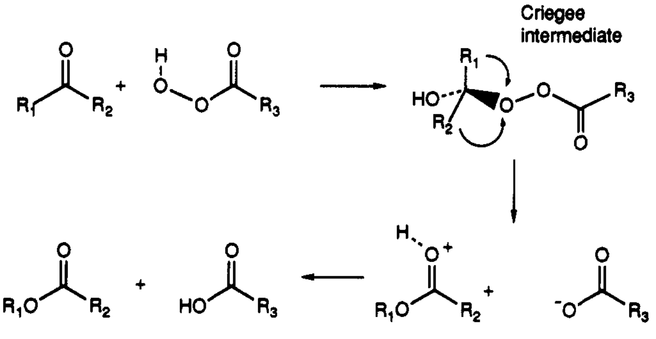
[14] Michelin, R. A.; Ros, R.; Strukul, G. *Inorg. Chim. Acta* **1979**, *37*, L491.

Table I. Oxidation of Complex Ketones with (A) 35% H_2O_2 Catalyzed by $[(dppe)Pt(CF_3)(CH_2Cl_2)]ClO_4$[a] and (B) *m*-Chloroperbenzoic Acid (MCPBA)[b]

oxidant	substrate	T (°C)	time (h)	products (mmol)		
				 1	 2	
A[c,d]		25	24	0.96	0.16	
A[d,e]		25	24	0.60	0.08	
B		25	24	0.70	0.37	
				 3	 4	 5
A		25	5	0.63	0.33	0.03
A[f]		25	5	0.15	0.04	0.04
B		25	3.5		0.75 + 3.0[g]	
				 6	others unidentified	
A		25	24	0.41	0.02	
A		50	6	0.34	0.03	
				 7	 8	
A		25	24	0.08	0.02	
A		80	25	0.90	0.04	
B		25	24	0.10	0.15	
				no reaction products		
A		50	24			
				 9	others unidentified	
A[h]		50	6	0.24	0.2	
B		25	24	6.5		

[a] Reactions conditions: Pt, 0.012 mmol; substrate, 12.5 mmol; 35% H_2O_2, 6.25 mmol; N_2, 1 atm. [b] Reaction conditions: substrate, 12.5 mmol; 90% MCPBA, 12.5 mmol; dichloroethane, 10 mL; N_2, 1 atm. [c] Pt, 0.118 mmol. [d] Dichloroethane, 2 mL. [e] 65% H_2O_2. [f] 2-Cyclohexen-1-one, 2.95 mmol; H_2O_2, 3.15 mmol; dichloroethane, 0.5 mL. [g] Hydroxy acid. [h] Pt, 0.025 mmol; tetrahydrofuran, 1 mL.

Scheme III



mechanism for the Baeyer–Villiger reaction with peracids[2] is shown in Scheme III. With camphor the above reported inversion of selectivity has been attributed[15] to steric hindrance effects caused by the two apical methyl groups in the Criegee intermediate which disfavor the migration of the tertiary C, and this situation reflects in the final products. Apparently, with Pt catalyst the selectivity toward lactone 1 is more pronounced indicating the formation of a similar reaction intermediate where the Pt complex is directly bound to the ketone. Under this hypothesis, steric effects should be even more marked because of the larger steric hindrance of the Pt complex compared to organic peroxy acids.

Although moderate, the amounts of products formed are comparable to those obtained with MCPBA under similar conditions (Table I). Other features of this reaction are the relative inertness to the use of more concentrated H_2O_2 solutions (Table I), the drop of the activity with H_2O_2 buffered at pH 7, and a H_2O_2 consumption similar to that shown in Figure 2 for cyclopentanone, although in this case the efficiency is better, ~30% of the H_2O_2 consumed going into products.

b. 2-Cyclohexen-1-one. In the classical Baeyer–Villiger oxidation of α,β-unsaturated ketones, lactonization normally precedes epoxidation, while in excess of peracid the epoxylactone is always the preferred product.[16] The data obtained with MCPBA (Table I) confirm this view, lactone 4 being produced and then partly hydrolyzed in the reaction medium. The selectivity of the system is totally changed with the Pt^+/H_2O_2 oxidation system (Table I), epoxide 3 being the main product. This catalyst produces also minimal amounts of the fully oxidized epoxylactone 5 and no hydrolysis products. The selectivity to 3 can be increased in the presence of a solvent (CH_2Cl_2) and using an excess of H_2O_2. However, the productivity is significantly lowered.

It has to be pointed out that no epoxidation of simple cyclic olefins was previously observed with this catalytic system.[16] It is likely that the conjugated carbonyl group induces a partial electron deficiency on the carbon–carbon double bond making the latter more susceptible to nucleophilic attack by PtOOH as is known from the mechanism of the epoxidation reaction.[9a]

c. Menthone. Historically, menthone was the first substrate studied by Baeyer and Villiger in 1899 using HSO_5^- as oxidant, which led to the discovery of the reaction.[17] As expected, the use of the Pt^+/H_2O_2 catalytic system led to the formation of the lactone corresponding to the migration of the secondary carbon i.e. lactone 6 together with some minor amounts of other unidentified oxidation products. Increasing the temperature to 50 °C leads to an increase in activity; however, the yields in product are strongly limited by the low efficiency with respect to hydrogen peroxide, whose decomposition increases with temperature.

d. Carvone. This substrate exhibits three different functionalities which may be oxidized and allows the study of the selectivity of the catalyst with respect to epoxidation vs lactonization. If only the lactonization arising from secondary C migration is considered, the number of possible mono-, di-, and trioxidation products sums up to 7. Indeed, only two oxidation products were observed both with the Pt^+/H_2O_2 system and with MCPBA, i.e. epoxide 7 and lactone 8. Similarly to 2-cyclohexen-1-one lactonization is preferred with the peracid, although in this case, under the standard conditions of Table I, the selectivity is much less pronounced. With Pt the selectivity is again inverted with the terminal epoxide as the preferred product. The productivity and the selectivity of the system are significantly increased when the reaction temperature is raised to 80 °C.

e. Indanones. As can be seen from Table I, 1-indanone did not give any products, while 2-indanone gave a mixture

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(15) (a) Sauers, R. R. *J. Am. Chem. Soc.* **1959**, *81*, 925. (b) Sauers, R. R.; Aheran, G. P. *J. Am. Chem. Soc.* **1961**, *83*, 1759. (c) Meinwald, J.; Frauenglass, E. *J. Am. Chem. Soc.* **1960**, *82*, 5253. (d) Murray, M. F.; Johnson, B. A.; Pederson, R. R.; Ott, A. C. *J. Am. Chem. Soc.* **1956**, *78*, 981. (e) Sauers, R. R.; Beisler, J. A. *J. Org. Chem.* **1964**, *29*, 210.

(16) Strukul, G.; Michelin, R. A. *J. Chem. Soc., Chem. Commun.* **1984**, 1538.

(17) Baeyer, A. V.; Villiger, V. *Ber. Dtsch. Ges.* **1899**, *32*, 3625.

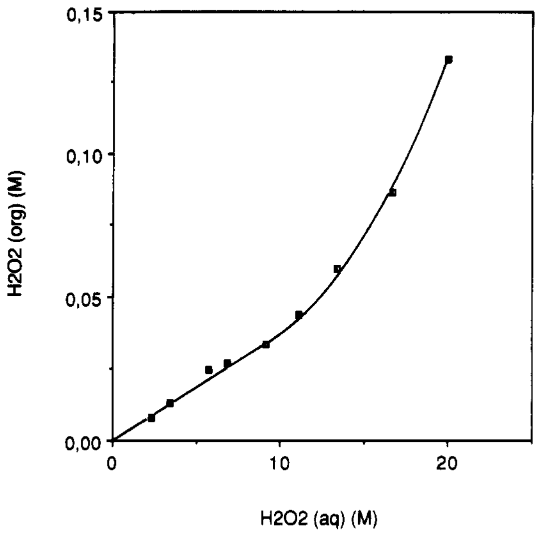

Figure 3. Partition curve of H_2O_2 between water and the organic phase (CH_2Cl_2 + cyclopentanone).

of seven different products of reaction, only one of which could be identified as lactone **9**; this constitutes about half of the total mixture. MCPBA seems to be much more productive and selective since lactone **9** is the only product that can be observed.

Mechanism of the Catalytic Reaction. The mechanism of the Baeyer–Villiger oxidation reaction with H_2O_2 was studied by initial rate analysis in the case of cyclobutanone using $[(dppe)Pt(CF_3)(CH_2Cl_2)]ClO_4$ as catalyst. Although the generalization of the results here reported to larger, more complex substrates must be taken cautiously, the choice of cyclobutanone for a mechanistic study was dictated by the observation that among the cyclic ketones so far reported it is the one that gives by far the cleanest and fastest reaction. In fact, we observed neither formation of the hydroxy acid from butyrolactone which leads to deactivation of the catalyst nor formation of peroxides that to some extent seems to be present in most simple cyclic ketones.[8] Moreover, as can be seen from Figure 2, H_2O_2 consumption seems to acceptably correspond to the amount of lactone formed. All reactions have been carried out in CH_2Cl_2 which allows one to observe also the effect of the substrate and minimizes possible side reactions as was observed with other simple ketones. The choice of a water-immiscible solvent was suggested from previous experience in the oxidation of olefins,[9,18] where the use of THF, dioxane, and alcohols (all good donors) that can compete with the substrate for coordination on the metal resulted in a severe lowering of the catalytic activity.

Since under the experimental conditions used the reaction medium is two-phase (CH_2Cl_2/H_2O), we first calculated the actual concentration of H_2O_2 in the organic phase according to a modification of the spectrophotometric method described by Wolfe.[19] The partition curve of H_2O_2 is shown in Figure 3. As can be seen, linearity holds for concentrations of H_2O_2 in water up to 10 M, while for higher concentrations the partition coefficient $[H_2O_2]_{aq}/[H_2O_2]_{org}$ decreases from ~300 to ~175.

A typical reaction profile for the oxidation of cyclobutanone obtained by gas chromatographic analysis is shown in Figure 4. This indicates that there is a sufficiently wide range where the initial rate can be conveniently analyzed. A summary of the kinetic data obtained by varying the concentrations of the reactants is reported in Table II.

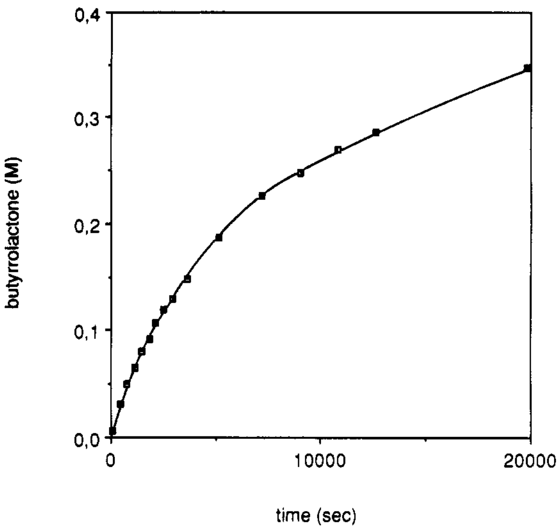

Figure 4. Typical reaction profile for the lactonization of cyclobutanone: [Pt], 2.85 × 10^{-3} M; [cyclobutanone], 0.55 M; $[H_2O_2]$, 0.038 M; solvent CH_2Cl_2.

Table II. Oxidation of Cyclobutanone: Summary of the Kinetic Data[a]

[ketone] (M)	$10^3[Pt]_T$ (M)	$10^3[H_2O_2]$ (M)	10^4 × rate (M s^{-1})
0	11.3	38	0
0.27	11.3	38	0.96
0.55	11.3	38	1.8
0.85	11.3	38	2.8
0.55	0	17 (38)	0.06 (0.04)
0.55	1.42	17 (38)	0.14 (0.45)
0.55	2.85	17 (38)	0.23 (0.77)
0.55	5.05	17	0.28
0.55	5.71	17 (38)	0.34 (1.1)
0.55	8.63	(38)	(1.7)
0.55	11.4	17	0.56
0.55	2.85	0	0
0.55	2.85	4.9	0.08
0.55	2.85	7.2	0.10
0.55	2.85	(8.2)	(0.11)
0.55	2.85	8.2	0.12
0.55	2.85	10	0.13
0.55	2.85	(16)	(0.25)
0.55	2.85	15	0.23
0.55	2.85	15	0.24
0.55	2.85	(27)	(0.50)
0.55	2.85	30	0.54
0.55	2.85	35	0.66
0.55	2.85	36	0.70
0.55	2.85	(38)	(0.77)
0.55	2.85	(80)	(1.7)

[a] Reactions were carried out in CH_2Cl_2 at 25 °C under an N_2 atmosphere, using H_2O_2 buffered at pH 2. The concentration of H_2O_2 refers to the organic phase and was determined according to Figure 3. Number in parentheses refer to experiments carried out with unbuffered H_2O_2.

a. Effect of the Catalyst. The kinetic study was initially carried out without buffering the H_2O_2 solution. Under these conditions the effect of total Pt concentration is shown in Figure 5A. As is clear there is a first-order dependence with an intercept on the *y* axis. This indicates that there is a modest contribution from a non-Pt-catalyzed oxidation of cyclobutanone. Baeyer–Villiger oxidation of cyclobutanone with H_2O_2 is known to occur under basic conditions;[20] however, as said above, hydrogen peroxide is moderately acidic and the contribution observed at $[Pt]_T$

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[18] Strukul, G.; Michelin, R. A. *J. Am. Chem. Soc.* **1985**, *107*, 7563.

[19] Wolfe, W. C. *Anal. Chem.* **1962**, *34*, 1328.

[20] (a) Bogdanowicz, M.; Ambelang, T.; Trost, B. M. *Tetrahedron Lett.* **1973**, 932. (b) Grieco, P. A. *J. Org. Chem.* **1972**, *37*, 2363.

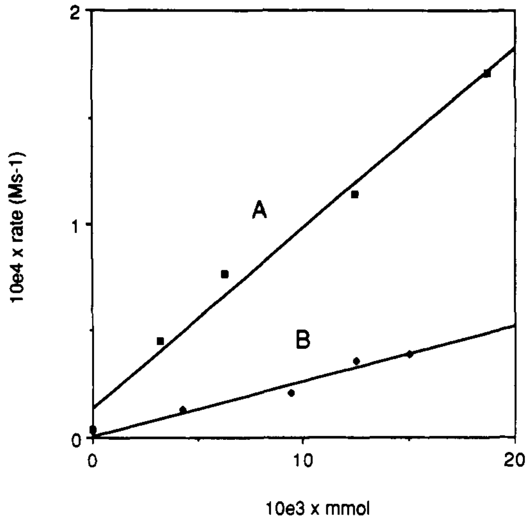


Figure 5. Lactonization of cyclobutanone: Effect of increasing amounts of Pt complex (A) and perchloric acid (B) on the reaction rate.

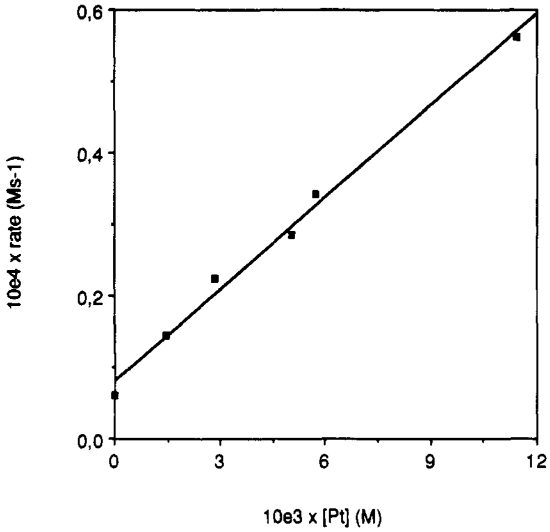


Figure 6. Lactonization of cyclobutanone: Effect of the concentration of Pt catalyst on the rate of lactonization of cyclobutanone. Experiments were buffered at pH 2.

= 0 might be due to the acidity of the oxidant. Moreover, when Pt concentration is increased, the acidity of the system increases because of eq 1. We have therefore checked the effect of increasing amounts of H^+ (added as $HClO_4$) in the absence of catalyst on the reaction rate, and the results are shown in Figure 5B. Since the Pt complex dissolves in the organic phase, while $HClO_4$ is added to the aqueous phase, in order to have comparable data, Figure 5 is plotted in millimoles. As can be seen, even H^+ displays a first-order effect on the reaction rate. These observations suggest that the effect of platinum concentration can be correctly analyzed only at constant pH. This was determined at pH 2 by buffering the aqueous phase with hydrosulfate buffer and is shown in Figure 6. Again a first-order dependence is observed.

b. Effects of Cyclobutanone and Hydrogen Peroxide. Since cyclobutanone is not involved in eq 1, its effect on the reaction rate was determined without buffering the aqueous phase. For concentrations higher than ~0.9 M the reaction rate starts becoming controlled by the diffusion of H_2O_2 from the aqueous to the organic phase and meaningless data are obtained. The effect of cyclobutanone concentration is shown in Figure 7.

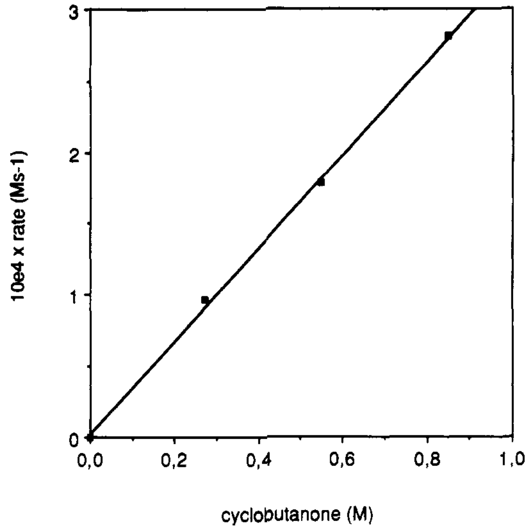


Figure 7. Lactonization of cyclobutanone: Effect of the concentration of cyclobutanone on the reaction rate.

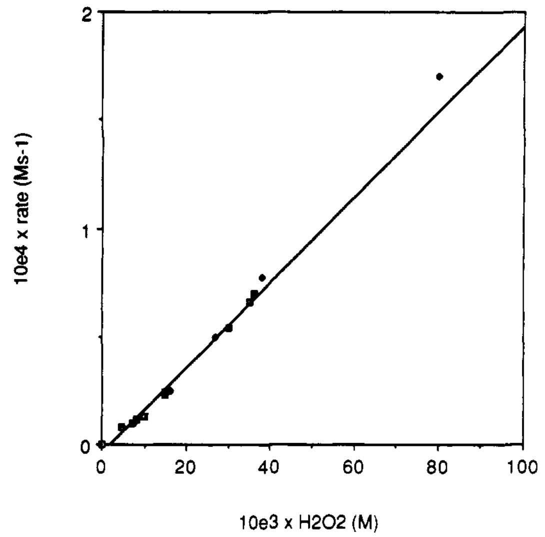


Figure 8. Lactonization of cyclobutanone: Effect of the concentration of hydrogen peroxide on the reaction rate. Key: Open squares, experiments buffered at pH 2; filled diamonds, unbuffered experiments.

Conversely, the effect of H_2O_2 concentration was determined both with and without buffer in the aqueous phase, but the differences on the corresponding reaction rates were found to be modest. Data are plotted together in Figure 8 and seem to fit a linear relationship.

c. Evolution of the Catalyst. In order to gain some information on the evolution of the$[(dppe)Pt(CF_3)(CH_2Cl_2)]ClO_4$ complex in the early stage of the process, a step-by-step spectroscopic analysis of the reaction mixture was carried out following the sequence with which the reactants are introduced into the catalytic system.

The complex (0.050 g) was dissolved in CD_2Cl_2 (0.6 mL), and the corresponding ^{19}F NMR spectrum was run at room temperature. This showed the usual[9a] broad, fluxional spectrum with a signal centered at δ −30.24 ppm ($^3J_{FP_{cis}}$ ~8 Hz, $^3J_{FP_{trans}}$ 54 Hz, $^2J_{FPt}$ 510 Hz). The addition of cyclobutanone (0.1 mL) led to the complete formation of a new species (δ −29.61 ppm; $^3J_{FP_{cis}}$ 8.4 Hz, $^3J_{FP_{trans}}$ 56.2 Hz, $^2J_{FPt}$ 519 Hz), still moderately fluxional, indicating that the ketone is a labile ligand.

A parallel IR analysis on a similar solution (CH_2Cl_2) showed two bands in the C=O region, one centered at 1785 cm^{-1} typical of free cyclobutanone and another at

1710 cm^{-1} that may well be assigned to coordinated cyclobutanone.

Addition of 35% H_2O_2 (0.01 mL) to the ^{19}F NMR experiment resulted in a drastic lowering of the quality of the spectrum due to the presence of at least two major species (all fluxional) that can be hardly analyzed because of extensive overlapping of signals.

d. Oxygen-Transfer Step. The observed first-order dependence on the three reactants seems to suggest a different reaction mechanism with respect to what was observed in the epoxidation of olefins[9a] and in the hydroxylation of aromatics,[21] where a second-order dependence with respect to the catalyst suggested a rate-determining step of the type shown in eq 4. Conversely,

$$\mathrm{Pt(substrate)^+ + PtOOH \rightarrow oxid.\ product + Pt^+ + PtOH} \quad (4)$$

in the present case the rate-determining step can be viewed as either the oxygen transfer from free H_2O_2 to a complex of the type Pt(ketone)$^+$ formed in the reaction mixture (eq 5) or the interaction (eq 6) between the free ketone

$$\mathrm{Pt(ketone)^+ + H_2O_2 \rightarrow lactone + PtOH + H^+} \quad (5)$$

$$\mathrm{PtOOH + ketone \rightarrow lactone + PtOH} \quad (6)$$

$$\mathrm{PtOH + H_2O_2 \rightarrow PtOOH} \quad (7)$$

and the PtOOH species formed from eq 7. Equation 6 bears strong similarities to the reaction between the platinum dioxygen complex and ketones reported in Scheme I and studied kinetically by Ugo and co-workers.[7c] Moreover, the electrophilic character of the peroxygen is certainly more marked in PtOOH rather than in free H_2O_2.[9a] However, this view contrasts with the observation that PtOH complexes are less effective catalysts compared to Pt$^+$ complexes.

In order to find experimental support for the oxygen-transfer step, the three possibilities indicated in eqs 4–6 have been reproduced in stoichiometric experiments.

In the first experiment, cyclobutanone (0.40 mmol) was dissolved in dry CH_2Cl_2 (2.1 mL), and to the solution was added (dppe)Pt(CF_3)(OOH) (0.0473 mmol). The formation of γ-butyrolactone was monitored with GLC, and the corresponding results are shown in Figure 9A.

In the second experiment, [(dppe)Pt(CF_3)(CH_2Cl_2)]ClO_4 (0.0473 mmol) was dissolved in dry CH_2Cl_2 (2.1 mL) and an excess of cyclobutanone was added (0.40 mmol) to ensure displacement of the coordinated solvent, followed by (dppe)Pt(CF_3)(OOH) (0.0473 mmol). The corresponding reaction profile is shown in Figure 9B.

In the third experiment, [(dppe)Pt(CF_3)(CH_2Cl_2)]ClO_4 (0.0473 mmol) was dissolved in dry CH_2Cl_2 (2.1 mL) and an excess of cyclobutanone was added (0.40 mmol) followed by free H_2O_2 (0.0473 mmol). The corresponding reaction profile is shown in Figure 9C.

A comparison of the experimental results indicates that eq 5 is viable, while eq 6 is very unlikely to occur in the catalytic system. Furthermore, an oxygen transfer of the type indicated in eq 4 is also possible, but this seems to be excluded by the kinetic analysis of the catalytic system, at least under the experimental conditions used.

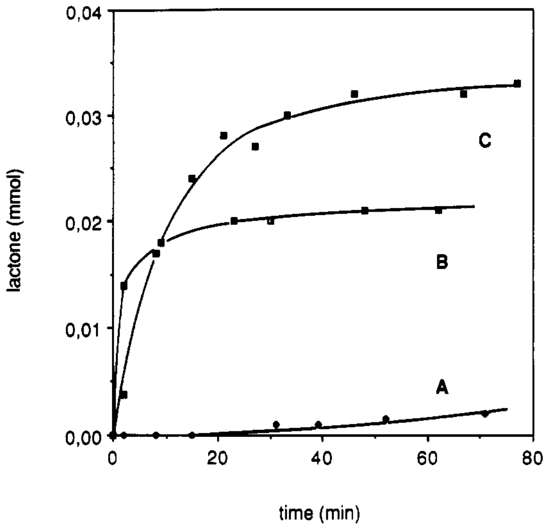
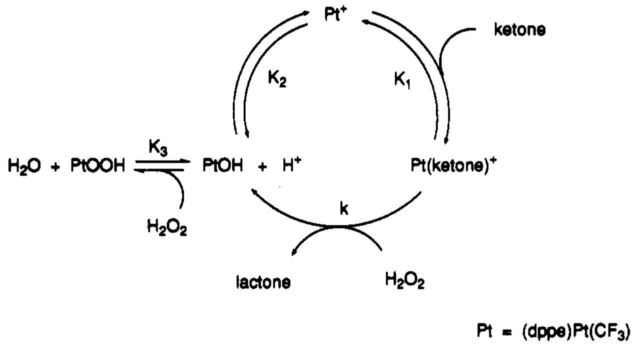


Figure 9. Lactonization of cyclobutanone: Oxygen-transfer step. Key: Curve A, cyclobutanone + PtOOH; curve B, Pt(cyclobutanone)$^+$ + PtOOH; curve C, Pt(cyclobutanone)$^+$ + H_2O_2.

Scheme IV



These experiments seem to suggest also that the role of Pt is more to increase the electrophilicity of the ketone rather than increasing the nucleophilicity of the oxidant, at variance with Scheme I.

e. Kinetic Scheme and Rate Law. Summing up the above experimental observations, the reaction mechanism indicated in Scheme IV is suggested. Evidence for Pt(ketone)$^+$ formation and the oxygen-transfer step have been reported above, while the protonation equilibrium and the PtOOH formation have been repeatedly observed in the chemistry of these complexes.[9,18,22] If all the equilibria are assumed as fast and the oxygen-transfer step is considered rate determining, the corresponding rate expression can be written as

$$\text{rate} = \frac{d[\text{lactone}]}{dt} = \frac{kK_1[\text{Pt}_\text{T}][\text{ketone}][\text{H}_2\text{O}_2]}{1 + K_1[\text{ketone}] + [\text{H}_2\text{O}_2]/K_2[\text{H}^+] + K_3[\text{H}_2\text{O}_2]/K_2[\text{H}^+]}$$

This is in agreement with the experimental observations if it is assumed that in the ranges of concentration explored both [ketone] and [H_2O_2] are never sufficiently high to display saturation effects in the corresponding plots (Figures 7 and 8).

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(21) Marsella, A.; Agapakis, S.; Pinna, F.; Strukul, G. *Organometallics*, submitted for publication.

(22) (a) Strukul, G.; Ros, R.; Michelin, R. A. *Inorg. Chem.* **1982**, *21*, 495. (b) Strukul, G.; Sinigalia, R.; Zanardo, A.; Pinna, F.; Michelin, R. A. *Inorg. Chem.* **1989**, *28*, 554. (c) Strukul, G.; Zanardo, A.; Pinna, F. *Studies on Surface Science and Catalysis*; Centi, G., Trifirò, F., Eds.; Elsevier: Amsterdam, 1990; Vol. 55, p 81.

Table III. Effect of pH on the Oxidation of Cyclobutanone with and without [(dppe)Pt(CF$_3$)(CH$_2$Cl$_2$)]ClO$_4$ Catalyst[a]

	10^4 × rate (M s^{-1})			10^4 × rate (M s^{-1})	
pH	with Pt	without Pt	pH	with Pt	without Pt
1.0	0.28	0.137	1.8	0.15	0.045
1.5	0.20	0.071	2.2	0.12	0.037

[a] Experimental conditions: [Pt], 2.85×10^{-3} M; [cyclobutanone], 0.55 M; [H$_2$O$_2$]$_{org}$, 15.7×10^{-3} M; hydrosulfate buffer; solvent CH$_2$Cl$_2$; N$_2$, 1 atm; T = 25 °C.

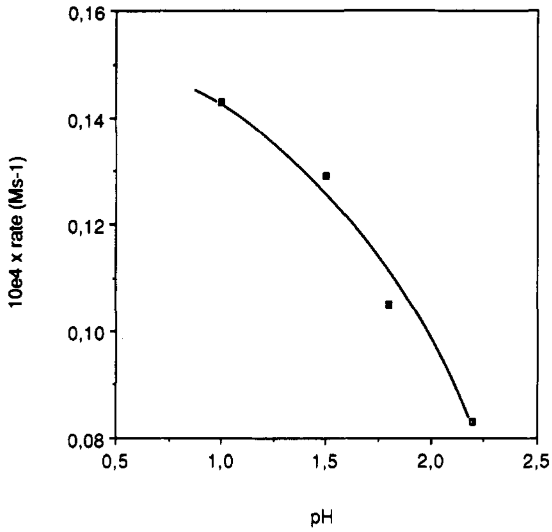
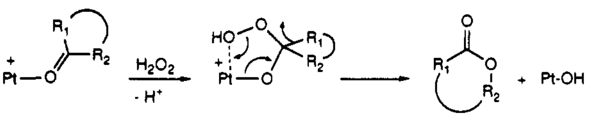


Figure 10. Lactonization of cyclobutanone: Net effect of pH on the reaction rate of the catalyzed reaction.

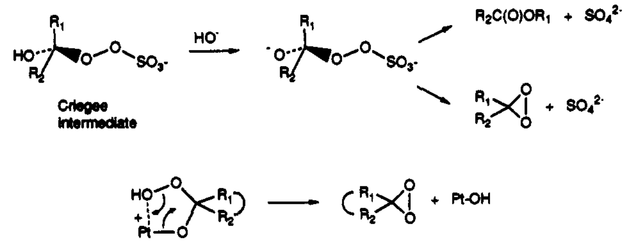
f. Effect of pH. The rate equation predicts a direct dependence on H$^+$ concentration. This effect is in qualitative agreement with the experimental observations reported above for cyclohexanone but was further checked for cyclobutanone by buffering the catalysis solution in a pH interval between 1.0 and 2.2 with hydrosulfate. Because of the ability of H$^+$ itself to catalyze the reaction, experiments were carried out both with and without the Pt catalyst. The results are reported in Table III, and the net effect of the Pt catalyst at different pH values is shown in Figure 10 providing further experimental support to the mechanism proposed.

Comparison with Peracids. In the kinetic scheme proposed, no details are given on the mechanism with which oxygen transfer occurs. In view of the isolation of peroxymetallacyclic products in the reaction between ketones and the Pt dioxygen complex[7] shown in Scheme I and the involvement of quasi-peroxymetallacyclic species in the epoxidation or the ketonization of olefins with this class of complexes,[9,18,22b,c] it seems reasonable to suggest a similar intermediate in the present oxidation reaction (eq 8). The positive charge and the presence of the electron-withdrawing –CF$_3$ ligand on Pt increase the electrophilic character of the carbonyl carbon that is made prone to attack even by a poor nucleophile like free H$_2$O$_2$ to produce the quasi-peroxymetallacyclic intermediate. The easy migration of the terminal OH group on Pt (a pathway that is not accessible to the peroxymetallacyclic compound in Scheme I) provides the necessary driving force leading to C=O formation and the breaking of the C–C bond and subsequent ester formation.



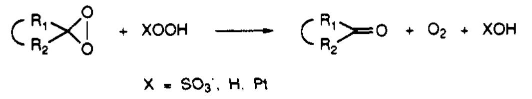
(8)

Scheme V



According to this interpretation, the oxygen transfer results from a two-stage process, similarly to the oxidation with peracids (Scheme III). In the latter case, the critical stage for product formation has been recognized to be the leaving of the organic acid anion,[2] a process that is facilitated by electron-withdrawing substituents in the peroxy acid. This is why free H$_2$O$_2$ (leaving group HO$^-$) is normally a very poor oxidant for this reaction. According to the mechanism proposed, there seems to be no role of Pt in increasing the nucleophilicity of H$_2$O$_2$ as was found in other oxidation reactions.[9,18,22b,c] In addition to some general advantages in the use of this class of complexes as catalysts for oxidation reactions with hydrogen peroxide (compatibility with aqueous media, minimal Fenton-type H$_2$O$_2$ decomposition, stability of all the reaction intermediates, etc.), the key importance of Pt in the present Baeyer–Villiger oxidation reaction seems to be the ability to provide an assisted pathway to the leaving of the HO$^-$ group, which makes hydrogen peroxide a viable oxidant for the reaction.

A further analogy with the behavior of peracids is found in H$_2$O$_2$ consumption. Since their original discovery,[17] Baeyer and Villiger recognized the ability of cyclic ketones to decompose the oxidant (in their case monopersulfuric acid) and suggested the possible formation of an elusive dioxirane compound as the intermediate leading to lactone formation. The recent synthesis of dioxiranes[23] from potassium monoperoxysulfate and ketones and a careful labeling study by Edwards et al.[24] have revalued the intermediacy of dioxiranes in Baeyer–Villiger oxidation, at least for those ketones containing groups of moderate migratory aptitude. According to these authors the Criegee intermediate (Scheme V) can either evolve to esters or dioxiranes and these, in turn, can further react with the oxidant leading to its decomposition. This view may well apply to our Pt-catalyzed oxidation (Scheme V), dioxirane formation occurring from the quasi-peroxymetallacyclic intermediate via a simple rearrangement that is identical to that suggested for epoxide formation[9,22b] and parallel to lactone formation (eq 8). Once formed, the dioxirane would easily react with the oxidant according to eq 9. This could well explain the excess hydrogen peroxide that is lost during the catalytic reaction that is due only in minimal amount to impurities in the reactants.



(9)

Conclusion

The results reported in this work represent the first fully characterized, unequivocal example of a Baeyer–

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(23) For recent reviews, see: (a) Murray, R. W. *Chem. Rev.* **1989**, *89*, 1187. (b) Adam, W.; Curci, R.; Edwards, J. O. *Acc. Chem. Res.* **1989**, *22*, 205. (c) Curci, R. In *Advances in Oxygenated Processes*; Baumstark, A. L., Ed.; JAI Press: Greenwich, CT, 1990; Vol. 2, p 1 and references therein.

(24) Edwards, J. O.; Pater, R. H.; Curci, R.; Di Furia, F. *Photochem. Photobiol.* **1979**, *30*, 73.

Villiger oxidation of ketones with H_2O_2 catalyzed by a transition metal complex. The approach of using a d^8 transition metal as catalyst seems to fit better the electronic requirements of the reaction and provides a viable way for the leaving of the reduced oxidant from the reactive intermediate. In some cases interesting selectivities in multifunctional substrates have been observed, different from those obtained with organic peracids. The catalytic system suffers from problems like the possible deactivation of the catalyst by the hydroxy acids formed from lactone hydrolysis and the moderate to low efficiency with which hydrogen peroxide is consumed, but on the other hand, it compares in activity with well-established synthetic methods based on e.g. the stoichiometric use of MCPBA with several additional advantages: the low cost and safety of diluted hydrogen peroxide, the ease with which the catalyst can be recovered and, after minimal workup, reused, and, more generally, all the advantages intrinsic in the use of a catalytic system.

Experimental Section

Apparatus. IR spectra were taken on a Perkin-Elmer 683 spectrophotometer and on a Digilab FTS 40 interferometer either in Nujol mulls using CsI plates or in CH_2Cl_2 solution using CaF_2 windows. ^1H and ^{19}F NMR spectra were recorded on a Varian FT 80A spectrometer operating in FT mode, using as external references TMS and $CFCl_3$, respectively. Negative chemical shifts are upfield from the reference. GLC measurements were taken on a Hewlett-Packard 5790A gas chromatograph equipped with a 3390 automatic integrator. GLC–MS measurements were performed on a Hewlett-Packard 59940A mass selective detector connected to a Hewlett-Packard 5890 gas chromatograph. Identification of products was made with GLC or GLC–MS by comparison either with authentic samples or computer search in a Wiley library of MS spectra held in the instrumentation. UV–vis spectra were recorded on a Perkin-Elmer 554 spectrophotometer. pHmetric measurements were taken on an Amel 234 titrator pHmeter.

Materials. Solvents were dried and purified according to standard methods. Ketone substrates were purified by passing through neutral alumina, prior to use. Hydrogen peroxide (35% from Fluka, 70% from Degussa), *m*-chloroperbenzoic acid 90% (Janssen), diphoe, dppe, PPh_3 (all from Strem), and most of the oxidation reaction products were commercial products and used without purification.

The following compounds were prepared according to literature procedures: (dppe)Pt(CF_3)(OH),[11] [(dppe)Pt(CF_3)(CH_2Cl_2)]-ClO_4,[11] and (dppe)Pt(CF_3)(OOH).[22a]

(dppe)Pt(CF_3)(hexanoate). The complex (dppe)Pt(CF_3)(OH) (0.20 g, 0.29 mmol) was dissolved in degassed, N_2-saturated CH_2Cl_2 (10 mL), and and the solution was placed under nitrogen. Hexanoic acid (0.10 mL) was added, and the solution was stirred at room temperature for 1 h. Solvent was removed in vacuo, and the addition of Et_2O led to the precipitation of a pale yellow solid. This was filtered out, washed with Et_2O, and dried in vacuo. Recrystallization was from CH_2Cl_2/Et_2O (yield 90%). Anal. Calcd (found) for $C_{33}H_{35}F_3O_2P_2Pt$: C, 50.96 (51.23); H, 4.54 (4.27). IR: ν_{COO} 1540 (asym), 1185 (sym) cm^{-1}. ^{19}F NMR (CD_2Cl_2): δ −27.88 ppm (dd); $^3J_{FP_{cis}}$ 9.9 Hz, $^3J_{FP_{trans}}$ 57.9 Hz; $^2J_{FPt}$ 568 Hz.

Determination of H_2O_2 Concentration in CH_2Cl_2. This was performed according to a modified procedure of the method described by Wolfe.[19] To a typical blank reaction mixture (5 mL of CH_2Cl_2 + 0.25 mL of cyclopentanone) was added 1.0 mL of aqueous H_2O_2 of known concentration. The choice of cyclopentanone instead of cyclobutanone is due to the negligible oxidation of the former compared to the latter in the absence of catalyst. After stirring of the mixture for a few minutes, a 4-mL portion of the organic phase was separated. To this was added an equal amount of a 2.1 M solution of $TiCl_4$ in 6 N HCl, and the mixture was stirred for 10 min. A 3.0-mL portion of the orange aqueous phase was separated and diluted to 25 mL with the same 2.1 M solution of $TiCl_4$ in 6 N HCl. Spectrophotometric analysis was carried out at 418 nm (ϵ 741). The Lambert–Beer law was obeyed in the concentration range explored.

Catalytic Reactions. These were carried out in a 25-mL round-bottomed flask equipped with a reflux condenser, a stopcock for vacuum/N_2 operations, and a side arm fitted with a screw-capped silicone septum to allow sampling. Constant temperature (±0.1 °C) was maintained by water circulation through an external jacket connected with a thermostat. Stirring was performed with a Teflon-coated bar driven externally by a magnetic stirrer. Absence of diffusional problems below 2.8×10^{-4} M s^{-1} initial rates was determined by the conversion vs time plots independence of the stirring rate in kinetic experiments randomly selected from Table II. The concentration of the commercial H_2O_2 solution was checked iodometrically prior to use.

The following general procedure was followed: The required amount of catalyst was placed solid in the reactor which was evacuated and filled with N_2. Purified, N_2-saturated ketone was added under N_2 flow, followed by an appropriate amount of internal standard (cyclooctane) and if necessary by the required amounts of solvent. After heating at the required temperature under stirring for a few minutes, the H_2O_2 solution in the appropriate concentration was injected through the septum and time was started. The mixture was vigorously stirred to avoid diffusional problems. In buffered experiments the buffer solution (hydrosulfate or phosphate) was diluted 1/1 with 70% H_2O_2 prior to use.

All reactions were monitored with GLC by direct injection of samples taken periodically from the reaction mixtures with a microsyringe. In kinetic studies initial rate data were determined from conversion vs time plots. Prior quenching of the catalyst with LiCl did not show any differences in randomly selected analyses. Separation of the product was performed on 25-m HP-5 or 25-m HP-1 capillary columns using a flame ionization detector. Quantitative data were obtained from calibration curves of the various reactants and products vs the internal standard (cyclooctane).

The amount of residual H_2O_2 at different times was determined by sampling 10-μL aliquots from the aqueous phase. These were diluted in water and titrated iodometrically.

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