Is it Possible to Achieve Highly Selective Oxidations in Supercritical Water? Aerobic Oxidation of Methylaromatic Compounds

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Abstract: We have demonstrated that different methylaromatic compounds [1,4-dimethylbenzene (*p*-xylene), 1,3-dimethylbenzene (*m*-xylene), 1,2-dimethylbenzene (*o*-xylene), 1,3,5-trimethylbenzene (mesitylene) and 1,2,4-trimethylbenzene (pseudocumene)] can be aerobically oxidized in supercritical water (scH_2O) using manganese(II) bromide as catalyst to give corresponding carboxylic acids in the continuous mode over a sustained period of time in good yield. No partially oxidized intermediates (i.e., toluic acid and benzaldehydes) were detected for the dimethylbenzenes and mesitylene reactions. By fine tuning pressure and temperature, scH_2O becomes a solvent with physical properties suitable for single-phase oxidation since both organic substrate and oxygen are soluble in scH_2O. There is a strong structural similarity of metal/bromide coordination compounds in the active oxidation solvents (acetic acid and scH_2O) which does not exist in the much less active H_2O at lower temperatures. This may account for the successful catalysis of the reactions reported herein. Aromatic acids produced by the loss of one methyl group occurred in all of these reactions, i.e., 3–6% benzoic acid formed during the oxidation of the dimethylbenzenes. Part of this loss is thought to be due to thermal decarboxylation. The thermal decarboxylation process is monitored *via* Raman spectroscopy.

Keywords: arenes; benzylic oxidation; C–H activation; green chemistry; manganese; supercritical water (scH_2O)

Introduction

Processes involving the partial selective oxidation of organic molecules (typically hydrocarbons) are diverse and widely used in almost all of the important commodity and fine chemical industries including the manufacture of pharmaceuticals, agrochemicals and polymer intermediates. In an ideal partial oxidation process, only the organic substrate, air and (a small amount of) catalyst should be introduced into the reactor. If the catalyst can be recovered in some way (and subsequently recycled *via* either a continuous mode of operation, or *via* a fresh charge of substrate), then the work-up of the reaction mixture should only involve product recovery and recycling of un-reacted substrate.

At least 73 different aromatic carboxylic compounds have been prepared from methylaromatic compounds by aerobic oxidation in yields from 75–95% using around 30 different combinations of soluble metal/bromide catalysts.[1] Acetic acid is the usual solvent used although many others have been evaluated. There is a strong synergistic relationship of the metals with the bromide anion. The most common catalyst is a mixture of cobalt (II) and manganese (II) acetates with hydrobromic acid at a metals/bromide ratio of 1:1 mol/mol. This area has been extensively reviewed.[1] Nearly all aerobic oxidations, such as the methylaromatic compounds considered here (Scheme 1), generate water as a by-product. For this reason, water is a good candidate for use as an environmentally more friendly solvent for oxidation. Indeed, the use of water as solvent has already been explored by different research groups.[2] Water, however, is a poor solvent for autoxidation, that is, for oxidations in which molecular oxygen is activated *via* a free radical chain process.

Water is generally known to both activate[3] and deactivate[1,4] metal-catalysed autoxidations. However, with the amounts of water produced during most reactions (>5 wt %) a large decrease in the rate of reaction occurs. A direct comparison of water, i.e., non-

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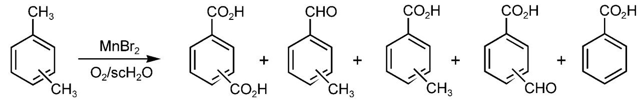


1a: *p*-
b: *m*-
c: *o*-

5a: *p*-
b: *m*-
c: *o*-

2a: *p*-
b: *m*-
c: *o*-

3a: *p*-
b: *m*-
c: *o*-

4a: *p*-
b: *m*-
c: *o*-

6

Scheme 1. Aerobic oxidation of xylene isomers in scH$_2$O.

scH$_2$O, as a solvent to acetic acid is available.[5] For instance, in the Co/Mn/Br catalyzed autoxidation of *p*-xylene to terephthalic acid, to obtain the same purity of terephthalic acid, 33 times more catalyst was required. Furthermore, the use of water as a solvent required a two-stage continuous oxidation at 185 °C and 220 °C rather than a single stage continuous oxidation at 226 °C in acetic acid. Even then, there was no net decrease in the carbon loss due to carbon dioxide and monoxide formation and significant amounts of ring-brominated product were produced. Therefore non-scH$_2$O water as a solvent is distinctly inferior and much less active than acetic acid/water as solvent.

Supercritical water, scH$_2$O, has generated considerable interest as a reaction medium. As water approaches its critical point ($T_c = 374\,^\circ C$, $P_c = 22.1$ MPa, $\rho_c = 0.317\ g \cdot cm^{-3}$), the physical and chemical properties of the fluid change.[6] As the temperature is increased, at a fixed pressure of 20.0 MPa, the dielectric constant (ε) decreases from 80 (25 °C) to 36 (200 °C), 21 (300 °C), 14 (350 °C) and 2 (400 °C).[7] An additional change also occurs as a result of the breakdown in hydrogen bonding as the temperature is increased: the ionic product (K_w) increases with enhanced local concentrations of H_3O^+ and OH^- ions, which reach a maximum value at around 250 °C and 25 MPa and then decreases sharply at the critical point.[8] This property has allowed acid- and base-catalysed reactions to be performed at this near-critical temperature without adding further acid or base.[9] scH$_2$O has been widely investigated as a medium for the *total* oxidation of organic species, normally at temperatures of 550–750 °C. The advantage is that scH$_2$O acts as a solvent for both O$_2$ and organic substrates providing, in this way, a single phase oxidation which overcomes mass-transfer limitations.[10]

We have recently reported the continuous partial selective oxidation of *p*-xylene to terephthalic acid using manganese(II) bromide as the catalytic system.[11] Our process generates almost no 4-carboxybenzylaldehyde, thought to be due to the faster reaction rates experienced at higher temperature. 4-Carboxybenzaldehyde is one of the most problematic impurities in current commercial processes for manufacturing terephthalic acid since it co-precipitates with the terephthalic acid and requires an additional plant to hydrogenate it back to *p*-toluic acid, which is then separated by fractional crystallisation in water. In our process, the first example of continuous selective oxidation in scH$_2$O, terephthalic acid does not precipitate within the reactor, a factor in reducing the amount of 4-carboxybenzylaldehyde in the recovered product. Furthermore, the use of high temperature water offers significant cost advantages through enhanced energy recovery due to a higher process temperature and the elimination of organic solvents. The process totally eliminates the use of organic solvents and has promise as a potentially cleaner and economical route to manufacture of terephthalic acid, as has been recently shown by Dunn and Savage.[12] In their work the economic and environmental assessment of the scH$_2$O process is compared with the use of acetic acid as solvent.

We chose to study the methylaromatic compounds *o*-, *m*- and *p*-xylene and 1,2,4-trimethylbenzene (pseudocumene), because the oxidation of these compounds gives products that are all large volume, industrial processes with a large variety of uses.[13] *p*-Xylene, *m*-xylene and pseudocumene (**1a**, **1b**, and **11**) are usually oxidised in the liquid phase while *o*-xylene (**1c**) is oxidised in the gas phase using vanadium/oxide heterogeneous catalysis. The oxidations of *p*-xylene and *m*-xylene are continuous processes while pseudocumene oxidation requires a batch approach.[1] The oxidation of 1,3,5-trimethylbenzene (**7**) was chosen to compare with 1,2,4-trimethylbenzene (**11**) because it is known that methyl groups *ortho* to each other can deactivate the oxidation system.[14] Here, we show that despite significant reactivity differences of these reagents, high yields of the expected aromatic acids can be easily achieved in continuous aerobic selective partial oxidation in scH$_2$O.

Results

All the experiments were conducted using a continuous flow reactor, the configuration of which is shown in Figure 1. The continuous oxidation was carried out in scH$_2$O with MnBr$_2$ as the catalyst, using the optimised conditions previously reported by us for the oxidation of *p*-xylene to terephthalic acid.[11] The expected products and reactivity of methylaromatic compounds in the metal/bromide-catalysed autoxidation of methylaromatics in acetic acid have been detailed in the literature.[1] To a first approximation, the methyl groups are consecutively oxidised to their aromatic acids. The intermediates in highest yield are benzylic alcohols, acetates and benzaldehydes.[1,15]

In supercritical water oxidation, the same products have been seen for *p*-xylene.[11,16] Similarly, in subcritical batch aerobic oxidation of toluene, *m*-xylene, *o*-xylene, mesitylene (1,3,5-trimethylbenzene), and ethylbenzene, the same intermediates and products have been reported which have also been seen in metal/bromide-catalysed autoxidation in acetic acid.[1,17] The response to increases in catalyst concentration, temperature, and reaction time reported in scH$_2$O are similar to

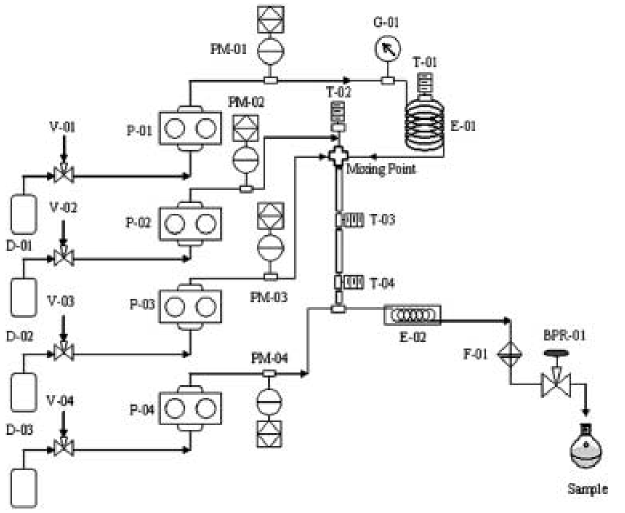


Figure 1. Configuration of the continuous aerobic oxidation reactor. D-01: water feed vessel, D-02: catalyst feed vessel, D-03: organic feed vessel, D-04: NaOH feed vessel, P-01: feed water pump (HPLC pump, Gilson model 305 Q-rate = 10 mL/min, $\Delta P_{max} = 60$ MPa), P-02: feed catalyst pump (HPLC pump, Gilson model 305 Q-rate = 10 mL/min, $\Delta P_{max} = 343$ bar), P-03: feed organic pump (HPLC pump, Gilson model 306 Q-rate = 5 mL/min, $\Delta P_{max} = 343$ bar), P-04: feed NaOH pump (HPLC pump, Gilson model 303 Q-rate = 5 mL/min, $\Delta P_{max} = 343$ bar), G-01: pressure gauge (Bourdon gauge), PM-01: pressure transducer and pressure monitor (water line), PM-02: manometric module, Gilson 802C, maximum pressure: 60 MPa (catalyst line), PM-03: pressure transducer and pressure monitor (organic line), PM-03: manometric module, Gilson 802C, maximum pressure: 60 MPa (NaOH line), T-01, T-02, T-03 and T-04: thermocouples, E-01: pre-heater (1/4 in AISI 316 tubing with 300 cm heat supplied by a cartridge heater (Watlow, 240 V, 200 W, FIREROD0247 NFRG2E164A) and a band heater (Watlow, 240 V, 200 W, NTB25X6UA3, 47 02 DM)), R-01: tubular reactor (1/4 in Hastelloy C276 pipe, 30 cm), E-02: cooler, F-01: 0.5 μm filter, BPR-01: Tescom, 26-1722-24-043 back pressure regulator max. P: 6000 psi.

those in acetic acid-based reactions.[1,11] For instance, increasing the catalyst concentration will result in increasing activity with the expected changes in *p*-toluic and terephthalic acid yields. These results suggest a similar operating mechanism in both acetic acid and scH_2O.Therefore it seems reasonable to assume that the reactivity of the methylaromatic compounds reported here will be similar to that in acetic acid, which has been detailed in the literature.[1]

The results obtained in the oxidation of the xylene isomers in scH_2O are summarised in Table 1. In the case of *p*-xylene (**1a**, Entry 3, Table 1) the previously optimised conditions gave around a 90% yield of terephthalic acid (**5a**).[11] Under the same experimental conditions (residence time, catalyst, and oxygen concentration, see Experimental Section) the oxidation of *m*-xylene (**1b**) and *o*-xylene (**1c**) led to reasonable yields of the corresponding dicarboxylic acid, 66% of isophthalic acid (**5b**, Entry 2, Table 1) and 46% of *o*-phthalic acid (**5c**, Entry 1, Table 1). In all cases a high selectivity was achieved for the production of the corresponding dicarboxylic acid derivative **5a–c** and with levels of the incompletely oxidised intermediate products such as tolualdehydes **2a–c**, toluic acids **3a–c** and carboxybenzaldehydes **4a–c** below the detection limits. The only side product detected was benzoic acid **6**, which is likely to be produced by decarboxylation of either the dicarboxylic acid or other intermediates (**3a–c** or **4a–c**), since such a process is favoured by longer residence times at high temperature.[16] The lower yield observed for *m*-xylene oxidation (66%) as compared to *p*-xylene (90%) may be expected since there is no resonance stabilisation of the radical intermediate at the *meta* position as seen at the *ortho* and *para* position.[17]

The mesitylene oxidation (**7**) was also successful with a 78% yield and a selectivity of 93%, see Table 2. The only by-product detected was phthalic acid (**5c**, 7%). In this case, benzoic acid was not detected. The reaction of 1,2,4-trimethylbenzene (**10**, pseudocumene) gave trimellitic acid (**11**) with 80% selectivity and 50% yield, see Table 2 and Figure 2. The major by-product was isophthalic acid (**5b**), the concentration of which was around twice that of *o*-phthalic (**5c**) and terephthalic acids (**5a**), 10, 5 and 5% respectively.

Thus even given the significant differences in reactivities of these substrates,[1] we have demonstrated that all of the methyl groups on a given methylaromatic compound can be selectively oxidised to the carboxylic acid in good yield.

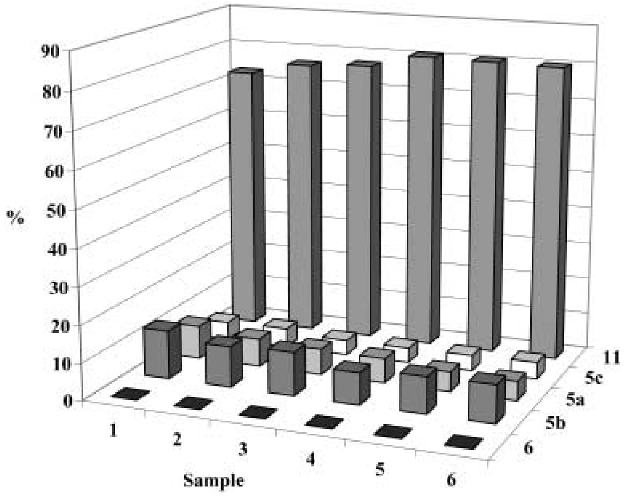


Figure 2. Plot showing the variation in selectivity of the aerobic oxidation reaction of 1,2,4-trimethylbenzene (**10**) in scH_2O observed during a continuous run. Samples were collected at 15 min intervals after the reactor had reached stable conditions. Catalyst strength: Mn 1719 ppm/Br 5000 ppm, flow rates: H_2O_2 (P-01): 8.19 mL/min, catalyst (P-02): 4.023 mL/min, organic (P-03): 0.061 mL/min, NaOH (P-04): 3.4 mL/min, temperature 380 °C and pressure 230 bar.

Table 1. Yields and selectivities observed in the continuous Mn/Br-catalysed oxidation of xylene isomers in scH_2O.

Entry	Compound	Yield [%][a] of dicarboxylic acid	Selectivity [%]				
			5 [%][b] (dicarboxylic acid)	**4** [%][c] (carboxylic benzaldehyde)	**3** [%][d]0 (toluic acid)	**2** [%][e] (tolualdehyde)	**6** [%][f] (benzoic acid)
1	*o*-xylene	46	94	<0.1	<0.1	<0.1	6
2	*m*-xylene	66	97	<0.1	<0.1	<0.1	3
3	*p*-xylene	90	95	<0.1	<0.1	<0.1	5
4	*o*-xylene[g]	–	–	1	30	–	69
5	*m*-xylene[g]	16	80	–	20	–	

[a] All calculated by HPLC. Yield [%] refers to the corresponding diacids phthalic, isophthalic and terephthalic acid respectively.

[b] Selectivity for *ortho-*, *meta-*, *para*-dicarboxylic acid calculated as % **5** = (molar conc. **5**/molar conc. **5** + molar conc. **2** + molar conc. **3** + molar conc. **4** + molar conc. **6**) × 100.

[c] Selectivity for *ortho-*, *meta-*, *para*-carboxylic benzylaldehyde as % **4** = (molar conc. **4**/molar conc. **2** + molar conc. **3** + molar conc. **4** + molar conc. **5** + molar conc. **6**) × 100.

[d] Selectivity for *ortho-*, *meta-*, *para*-toluenecarboxylic acid calculated as % **3** = (molar conc. **3**/molar conc. **5** + molar conc. **2** + molar conc.**3** + molar conc. **4** + molar conc. **6**) × 100.

[e] Selectivity for *ortho-*, *meta-*, *para*-tolualdehyde calculated as % **2** = (molar conc. **2**/molar conc. **5** + molar conc. **2** + molar conc.**3** + molar conc. **4**+ molar conc. **6**) × 100.

[f] Selectivity for carboxylic acid calculated as % **6** = (molar conc. **6**/molar conc. **5** + molar conc. **2** + molar conc. **3** + molar conc. **4** + molar conc. **6**) × 100.

[g] Results from Holliday et al.[17]

Table 2. Yields and selectivities observed in the continuous Mn/Br-catalysed oxidation of trimethylbenzene isomers in scH_2O.

Entry	Compound	Yield [%][a] of tricarboxylic acid	Selectivity [%]					
			8 [%][b] (trimesic acid)	**11** [%][c] (trimellitic acid)	**5a** [%][d] (terephthalic acid)	**5b** [%][e] (isophthalic acid)	**5c** [%][f] (phthalic acid)	**6** [%][g] (benzoic acid)
1	mesitylene	78	93	–	–	–	7	<0.1
2	pseudocumene	50	–	80	5	10	5	<0.1

[a] All calculated by HPLC. Yield [%] refers to trimesic and trimelitic acid respectively.

[b] Selectivity for trimesic acid calculated as % **8** = (molar conc. **8**/molar conc. **8** + molar conc. **5b** + molar conc. **6**) × 100.

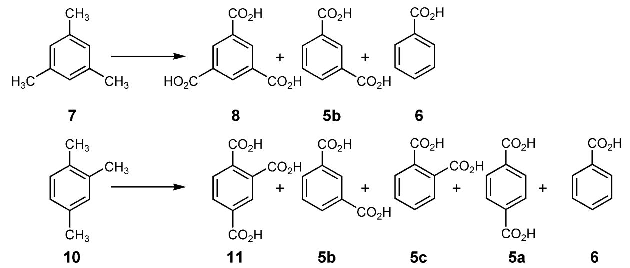
[c] Selectivity for trimelitic acid calculated as % **11** = (molar conc. **11**/molar conc. **11** + molar conc. **5a** + molar conc. **5b** + molar conc. **5c** + molar conc. **6**) × 100.

[d] Selectivity for terephthalic acid calculated as % **5a** = (molar conc. **5a**/molar conc. **11** + molar conc. **5a** + molar conc. **5b** + molar conc. **5c** + molar conc. **6**) × 100.

[e] Selectivity for isophthalic acid calculated as % **5b** = (molar conc. **5b**/molar conc. **11** + molar conc. **5a** + molar conc. **5b** + molar conc. **5c**+ molar conc. **6**) × 100.

[f] Selectivity for phthalic acid calculated as % **5c** = (molar conc. **5c**/molar conc. **11** + molar conc. **5a** + molar conc. **5b** + molar conc. **5c** + molar conc. **6**) × 100.

[g] Selectivity for carboxylic acid calculated as % **6** = (molar conc. **6**/molar conc. **5** + molar conc. **2** + molar conc. **3** + molar conc. **4** + molar conc. **6**) × 100.



Scheme 2. Aerobic oxidation of trimethylbenzene in scH_2O.

Discussion

Comparison of Results to the Literature

The 90% yield of terephthalic acid obtained in this work *via* continuous oxidation can be compared to the maximum yield of 57% obtained by Dunn, Urquhart and Savage *via* batch oxidation in supercritical water[16] and to 64% terephthalic acid yield from batch oxidation in sub-critical water by Holliday et al.,[17] see Table 1. Comparable yields of terephthalic acid can be obtained in hot water and in acetic acid ($>90\%$) but the latter requires a two-stage continuous process with a Co/Mn/Br catalyst. The first stage has an 80 min residence time and the second stage a 60 min residence time.[18] This is in contrast to that reported here with much lower catalyst concentration and very fast reactions (residence time <20 seconds), with comparable yields, and no detectable 4-carboxybenzaldehyde in the product.

Our results for the continuous oxidation of *m*-xylene and *o*-xylene both gave better selectivity and yields than those obtained by Holliday et al. using batch reactors and sub-critical temperatures in water.[17] The continuous oxidation of *o*-xylene gave an up to 46% yield of phthalic acid, while none was detected in the batch process, see Table 1. The long reaction time used in the batch experiments (~ 60 min for batch experiment against <20 seconds in a continuous reactor) favours the reaction of *o*-xylene to phthalic acid followed by the thermal decarboxylation at the *o*-position. Indeed, the major product produced by a batch process was benzoic acid, instead of phthalic acid (Entry 4, Table 1). Holliday obtained a 16% of isophthalic acid from *m*-xylene (Entry 5, Table 1) in contrast to this work with 66%.[17] The results by Holliday are likely to have been incomplete oxidations as evidenced by the presence of 4% *m*-toluic acid and the poor mass accountability (20%).

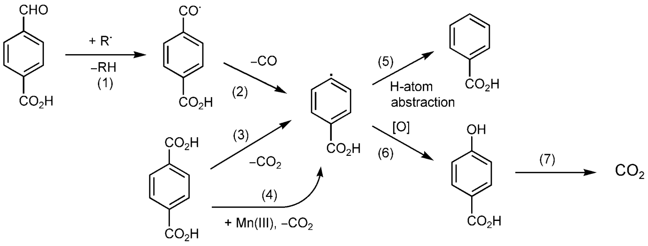
Except for *p*-xylene, the yields reported here for the other methylaromatic compounds are less than those in a conventional metal/bromide acetic acid solvent process.[1] Thus reported yields in acetic acid for *o*-xylene to *o*-phthalic acid of 82–83% compare to 46% reported here. Similarly for *m*-xylene to isophthalic acid the reported yields are between 61–93% compared to 66% in this work. For mesitylene to trimesic acid the reported yields are 71–88% compared to 78% found here. The yields of trimellitic acid from pseudocumene are 90–92% as compared to 50% found here However, these are perhaps unfair comparisons since the reported yields in acetic acid are ***optimised*** yields obtained by variation of the type of catalyst, the catalyst concentration, the reagent concentration, the temperature and the pressure. These variables are adjusted for each reaction to take into account the initial reactivity of the methylaromatic compound and the decreasing reactivity of the reagent as the successive methyl groups have been reacted to carboxylic acid groups. This has been discussed in detail.[1,19] In this work, the optimised conditions for the oxidation of *p*-xylene to terephthalic acid in scH_2O were used for all the other methylaromatic compounds. For instance, if pseudocumene was oxidised in an identical manner to the commercial process where p-xylene is converted to terephthalic acid using a Co/Mn/Br catalyst in acetic acid, much lower yields would be obtained. Therefore, a further optimisation of the reaction conditions for each particular substrate should lead to the corresponding aromatic acid with higher yields and selectivities in scH_2O.

Methyl Carbon Atom Loss: Evidence for Decarboxylation in scH_2O

It has been reported that the terephthalic acid yields in supercritical water decrease with longer reaction times due to 'decarboxylation'.[11,16] Here we find that *o-, m- and p-* xylene produce benzoic acid, mesitylene produces isophthalic acid and pseudocumene produces three different diacids. The appearance of the by-product containing one less methyl carbon atom is ***not*** unique to supercritical water oxidation but occurs in all metal- and metal/bromide-catalysed homogeneous autoxidations.[1]

Benzoic acid formed during dimethylbenzene oxidation can be formed in at least three different ways: 1) *via* decarbonylation (rxn 2, Scheme 3), 2) *via* thermal decarboxylation (rxn 3, Scheme 3) and 3) metal-catalysed decarboxylation (rxn 4, Scheme 3). For the metal-catalysed pathway, Mn(III,IV) and Co(III) are known to be present in autoxidation[1] and it is well-established that they can decarboxylate aromatic acids.[20,21] The decarbonylation mechanism is important in aliphatic aldehydes but not aromatic ones.[22] To confirm this we aerobically oxidised heptaldehyde and *p*-tolualdehyde with a Co/Mn/Br catalyst in acetic acid at 95 °C and atmospheric pressure. The vent carbon monoxide level during the oxidation of heptaldehyde was 5 vol % but only 0.03% for *p*-tolualdehyde.[39] Thus most of the observed benzoic acid is from a decarboxylation mechanism and not from decarbonylation.

Previously we provided experimental evidence for the metal-catalysed decarboxylation mechanism using a Co/Mn/Br catalyst in acetic acid.[14] During the aerobic oxidation of 1,2,4-trimethylbenzene, the addition of benzoic acid which has a ^{13}C-labelled carboxyl group, resulted in the observation of $^{13}CO_2$ in the vent gases. The temperature of the reaction was 95 °C where thermal decarboxylation should be very minimal. The increase in carbon dioxide formation is directly proportional to the concentration of benzoic acid added. The



Scheme 3. Pathways leading to loss of methyl carbon atoms during autoxidation.

addition of *o*-phthalic acid, rather than benzoic acid, results in more than twice the rate of formation of carbon dioxide, on a mole basis, consistent with *ortho* acid groups being more effectively decarboxylated than *para* or *meta*-ones.[14]

The oxidation of the phenyl radical to phenol and its subsequent complete combustion to carbon oxides, rxn 6, Scheme 3, is probably important since, except for *p*-xylene, the mass accountabilities are < 100.

The significantly lower yield of *o*-xylene to *o*-phthalic acid as compared to the *para*- and *meta*-isomers and its higher benzoic acid yield suggest that the decarboxylation of the *ortho* diacid grouping is more important than for the carboxylic acids in the *para*- and *meta*-positions. This is particularly striking during the oxidation of 1,2,4-trimethylbenzene where the yield of terephthalic acid is twice as high as that of *o*-phthalic acid. The results are consistent with the reported thermal decarboxylation of terephthalic and isophthalic acid which gave yields of 10–15% at 350 °C for 30 min but for trimellitic anhydride, complete degradation was observed.[23] Similarly it has been shown that *o*-phthalic acid thermally decarboxylates much easier than terephthalic and isophthalic acids. However, the relative importance of thermal and metal-catalysed decarboxylation, remains uncertain.

A better understanding of methyl carbon atom loss processes in the autoxidation process in scH_2O is important in order to minimise the extent of this side reaction. The stability of the aromatic carboxylic acid in ncH_2O and scH_2O has received previous attention.[23] We have started to study such reactions using high temperature, high pressure Raman spectroscopy. We have demonstrated that phthalic acid **2a** decarboxylates to benzoic acid **6** (Figure 3) at high temperature and in the presence of NaOH. The decarboxylation takes place above 300 °C and can be monitored using Raman spectroscopy.[24] The first clear observation is that a new band, which corresponds to the *mono*-substituted benzene ring, appears at 1000 cm^{-1}, whilst the existing peak corresponding to the disubstituted benzene ring at ~ 1049 cm^{-1}, decreases significantly in intensity. Further investigations based on this technique are in progress in order to obtain kinetic data relevant to this process, which will allow us to minimise its effect.

Mechanism of Autoxidation in Acetic acid and scH_2O

The mechanism of metal- and metal/bromide-catalysed aerobic oxidation of hydrocarbons is thought to be a metal-modified free radical chain mechanism. Benzylic radicals ($PhCH_2^{\bullet}$) are initially formed which react rapidly with dioxygen to form benzyl peroxy radicals ($PhCH_2OO^{\bullet}$). Benzyl hydroperoxides ($PhCH_2OOH$) are formed *via* a subsequent propagation step with a second molecule of the hydrocarbon. Recent publications have discussed the structure of the Co/Mn/Br catalyst in acetic acid and acetic acid/water mixtures,[25] the nature of the metal bromide species that activates the methylaromatic compound *via* benzylic radical formation,[26] and the activity and selectivity enhancements due to the presence of cobalt [27] and zirconium in the catalysts.[28] The key question here is why H_2O lowers the activity of $MnBr_2$ in CH_3CO_2H, but does not appear to affect its activity under supercritical conditions. We suggest that this is largely the consequence of the low dielectric constants of CH_3CO_2H and scH_2O, both of which promote ion-pairing. The role of Br^- is discussed in detail in separate paper in this issue.[38]

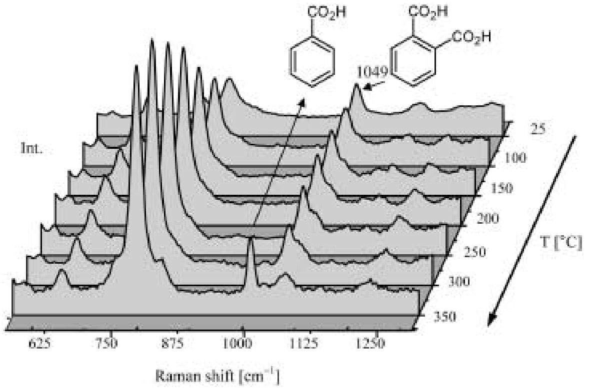


Figure 3. Raman spectra of an aqueous solution of 0.08 M phthalic acid and 0.04 M sodium hydroxide, heated isobarically at 300 bar stepwise from 25 to 350 °C. An NIR GaAlAs diode laser (785 nm) was used to excite the samples with ~ 300 mW output power using an Almega (Thermo Nicolet) Raman spectrometer, exploiting 180° (back-scattered) geometry of scattering.

Current evidence suggests that $MnBr_2$ (Mn/Br 1:2 mol/mol) produces higher yields of terephthalic acid in batch autoxidation in sub- and supercritical water than $CoBr_2$.[9,10] In acetic acid, the Mn/Br catalyst is much less active than the Co/Br catalyst. Also, increasing the bromide/metal ratio above 1/1 mol/mol reduces the activity.[1,28] This may suggest that $MnBr_2$ might be a *negative catalyst* in scH_2O rather than a positive one, i.e., decreases the rate of reaction rather than enhances it. With a system that produces high yields in 20 sec rather than 0.5–2.0 h, the necessity for catalyst control of the desirable reaction pathways may be much more important. For example, we have found in the highly active non-catalysed autoxidation of benzaldehyde in acetic acid that addition of Co and Co/Mn/Br catalysts ***reduces*** the rate of autoxidation and increases the selectivity as well.[39]

In metal/bromide autoxidations in acetic acid, the bromine is known to enhance selectivity by reducing the steady state concentration of the higher oxidation states of the metals *via* the reaction in Eq. (1).

Table 3. Selected properties of H_2O, scH_2O, acetic acid, and acetic acid/water mixtures and self dissociation constant (log K) of Co(II), Mn(II), Ni(II) bromide mixtures.

Solvent	T [°C]	ε	Self-dissociation [log K]	Geometry of Co(II) coordination compound	Metal bromide stability constants, K_1, K_2		
					Co(II)	Mn(II)	Ni(II)
H_2O	25	78	– 14	oct.	0.74, 0.40	1.9, 0.55	0.76, -
HOAc	25	6.2[34]	– 13.9[34]	oct., tet.	851, 295	large[15]	not available
10% HOAc-H_2O	25	17[34]	– 8.1[34]	oct.	small[25]	Small[15]	not available
scH_2O	427	4	– 20	oct., tet	not available	not available	large[31]
acetone	25	1.35	–	oct. tet	$K_1 \times K_2 = 1.6 \times 10^9$[34]	not available	not available

$$[\mathrm{M(III)}] + \mathrm{Br}^- \Longrightarrow [\mathrm{M(II)}] + \mathrm{Br(0)} \qquad \mathrm{M = Co, Mn} \tag{1}$$

Since these are homogeneously catalysed autoxidations, the speed of this critical reaction is obviously related to the metal/bromide coordination compounds that exist in solution. We would suggest that the metal/bromide coordination compounds in the active acetic acid, water/acetic acid, and scH_2O solvents are similar to each other but dissimilar to the much less active non-scH_2O water. The structural change of Ni(II)/Br mixtures from non-scH_2O water to scH_2O has been studied by Fulton and co-workers.[31] Using XAFS, near IR, and molecular dynamics, they followed the change from the simple $[\mathrm{Ni(II)(H_2O)_6}]^{2+}_{\mathrm{oct.}}$ at room temperature, to an equilibrium mixture of octahedral and tetrahedral species at 325 °C to a series of tetrahedral, ion-paired, Ni–Br species at 425 °C, Eq. (2).

$$\underset{\text{Ni/Br in non-scH}_2\text{O}}{[\mathrm{Ni(II)(H_2O)_6}]_{\mathrm{(oct.)}}} \longrightarrow \underset{\substack{\text{Ni/Br mixture in scH}_2\text{O}\\ \text{where \{\} indicates ion pairing}}}{\begin{array}{l}\{[\mathrm{Ni(II)(H_2O)_3Br]Br}\}_{\mathrm{(tet.)}} + [\mathrm{Ni(II)(H_2O)Br}]_{2\ \mathrm{dimer}}\\ \{\mathrm{Na[Ni(II)(H_2O)Br_3]}\}_{\mathrm{(tet.)}} + \{\mathrm{Na_2[Ni(II)Br_4]}\}\end{array}} \tag{2}$$

The structures of these compounds, as drawn by Fulton, are reproduced in Figure 4 and compared to those in anhydrous and hydrous acetic acid. Tomiyasu and coworkers have described a similar transition for $CoCl_2$ using UV-visible spectroscopy in water from 50 to 400 °C.[35]

In the following discussion we shall assume that the coordination chemistry for Ni(II) is the same as that for Co(II) and Mn(II). This assumption is reasonable since these are first row transition metals in the same oxidation state. Some of the similarities are shown in various different solvents in Table 3.

At concentrations normally used in autoxidations (0.01 – 0.02 M), these three metals are essentially octahedral ions of the structure $[\mathrm{M(II)(H_2O)_6}]^{2+}$ in non-scH_2O. Very little bromide is coordinated to the metals as evidenced by their small measured stability constants, see Table 3. Little ion-pairing occurs between the hydrated metals and the bromide.[32]

By contrast, in acetic acid most of the bromide is coordinated to the metals (see Table 2) and virtually all of the non-coordinated bromide exists as ion pairs.[25] Furthermore, the solution contains both octahedral and tetrahedral Co/Br species.[25,33] Additionally, there is evidence for both mononuclear and polynuclear metal

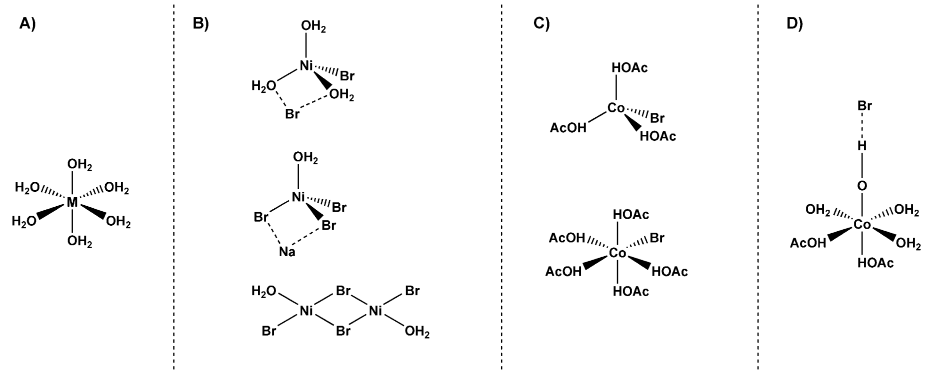


Figure 4. Predominant metal/bromide coordination compounds in active (scH_2O, HOAc, H_2O/HOAc) and inactive (non scwater) solvents **A)**: in water M = Ni(II), Mn(II), Co(II). **B)**: In supercritical water. **C)**: In anhydrous acetic acid. D): In hydrous acetic acid. Adapted from Refs.[31,35]

species.[25] The change in coordination from water to anhydrous acetic acid can be approximately illustrated as shown in Eq. (3).

$$[Co(II)(H_2O)_6]^{2+}_{(oct.)} \longrightarrow [Co(II)(HOAc)_5Br]^{+}_{(oct.)} + [Co(II)(HOAc)_3Br]^{+}_{(tet.)}$$

Co/Br mixture in non-scH_2O
Inactive Oxidation solvent

Co/Br mixture in anhydous acetic acid
Very Active Oxidation solvent

(3)

Addition of water to acetic acid, results in most of the bromide being no longer bound to the Co and Mn and the tetrahedral species no longer exists. But, because the dielectric constant only increases modestly (6.2 to 17 when changing from 0 to 10 wt % water), virtually all of the anions are ion-paired with the metals.[34] The oxidation activity of the coordination compounds in this mixture is attributed to the presence of the easily accessible ion-paired bromide to the metallic oxidants [typically Co(III), Mn(III), Mn(IV)] and the labile, weak acetic acid ligand in the coordination sphere.[25] The decreased activity in the presence of the water is due to many factors including the replacement of the acetic acid ligands by water and the reduced reaction rate *via* the metal-catalysed pathways. The transition of cobalt/bromide mixtures from anhydrous acetic acid to hydrous acetic acid can be written as in Eq. (4).

$$[Co(II)(HOAc)_5Br]^{+}_{(oct.)} + [Co(II)(HOAc)_3Br]^{+}_{(tet.)} \longrightarrow \{[Co(II)(HOAc)_n(H_2O)_{n-6}]Br\}^{+}$$

Co/Br mixture in anhydous acetic acid
Active Oxidation solvent

Co/Br mixture in hydrous acetic acid
where {} indicates ion pairing
Less Active Oxidation solvent

(4)

Thus, in the transitions from non-scH_2O to ***either*** acetic acid or scH_2O, the types of coordination compounds formed are similar. Transitions occur from simple hexaaquometal(II) ions to dimeric species and to mixtures of octahedral and tetrahedral species, the bromide being either predominately directly bonded or ion-paired to the metal. The probable reason for these structural changes is understandable from the changes in dielectric constant in Table 3. Non-scH_2O water has a value of 78 at room temperature which changes to 4 in scH_2O. The dielectric constant in scH_2O is similar to that of acetic acid. Acetone, another solvent of low dielectric constant, also has an equilibrium between octahedral and tetrahedral bromo compounds with large equilibrium constants.[26b]

Conclusion

By tuning the pressure and temperature, scH_2O becomes a solvent with physical properties which provide for single-phase oxidation (organic and oxygen are soluble in scH_2O) and an environment which activates the catalytic system. However, to quote V. Krukonis *"There is no point in doing something in a supercritical fluid just because it's neat. Using the fluids must have some real advantage"*.[36] We believe that continuous selective partial oxidation is an excellent example of the use of supercritical fluids to develop processes which are not only environmentally more friendly but also have additional technical and commercial advantages.

Certainly, a remarkable advantage is the absence of detectable intermediate alcohols and aldehydes during the continuous scH_2O aerobic oxidation of *p*-, *m*- and *o*-xylenes. This is important because it is not enough to produce a product in high yield but also it must be of sufficient purity. This is commercially important, because the purification process can be as expensive as the reaction producing the product. In terephthalic acid manufacture, for example, the oxidation reaction produces solid terephthalic acid with a purity of about 99.6%. However, the purification of these solids to remove 4-carboxybenzaldehyde requires a separate plant to be built which hydrogenates the 4-carboxybenzaldehyde back to *p*-toluic acid which can be subsequently separated by fractional crystallization from water. In the case of *o*-xylene oxidation, the selectivity achieved here in scH_2O is better than that reported for the equivalent transformation of *o*-xylene in the gas phase (78% based on o-xylene). The by-products in the gas phase process include *o*-toluic acid, phthalide, benzoic acid, and maleic acid. Two-stage distillations are required to reach a purity of at least 99.8%. In our case, simple selective crystallisation may give the product with the required degree of purity and would eliminate the need for expensive distillations. The replacement of the current gas-phase process by the liquid phase scH_2O oxidation would result in about a 10-fold increase in the space-time yield.[3]

Additionally, the elimination of the acetic acid solvent by use of scH_2O would constitute a major process improvement because the metal-catalysed and free radical-catalysed decomposition of acetic acid is one of the major variable costs in the current manufacture of terephthalic acid. Indeed, one of the major uses of acetic acid is to replace the acetic acid losses in terephthalic acid and dimethyl terephthalate manufacture.[16]

Summarising, the use of supercritical water for this process provides: i) chemical advantages (high selectivity by fine-tuning of solvent properties), ii) the environmental advantage of avoiding the use of acetic acid as solvent and its subsequent waste by unwanted oxidation and production of CO_2, iii) economic advantages resulting from more selective and faster reaction, plus operation at high temperature allowing for more efficient energy recovery from the exothermic reactions. and iv) process intensification: from a large reactor volume with a *ca.* 40 minute residence time, to a reactor with a residence time of the order of a few minutes. The new technology offers a cleaner route to aromatic carboxylic acids.

Experimental Section

Oxidation of Polymethylbenzenes

All the experiments were conducted using a continuous flow reactor, the configuration of which is shown in Figure 1. The continuous oxidation was carried out in scH_2O with $MnBr_2$ as the catalyst, using the optimised conditions previously reported by us for the oxidation of *p*-xylene to terephthalic acid.[11] The exotherm was minimised by using relatively dilute solutions (< 5% organic w/w). An aqueous solution of H_2O_2 of 2% vol. was used as the source of O_2, supplied *via* a reciprocating pump which allows for accurate control of the O_2:substrate ratio; a critical parameter which is difficult to achieve reproducibly using a gas booster pump to compress oxygen gas. The H_2O_2 solution was pumped into the system and decomposed in the pre-heater under the required experimental conditions to generate a homogeneous mixture of O_2 and scH_2O. Tester and co-workers have shown that there are no significant differences in the oxidation of MeOH in scH_2O using either O_2 or decomposed H_2O_2 as the oxidant.[37] The residence time in the pre-heater was long enough to guarantee total decomposition of H_2O_2 to H_2O and O_2. The residence time calculated from the total reactor volume divided by the volumetric flow rate, was around 5 seconds. The total volume was taken as the sum of the volume of the tubular reactor, constructed from a 30-cm length of ¼ in o.d. Hastelloy C276 pipework. The volumetric flow rate was based on the physical properties of H_2O at the reaction conditions ($T = 380\,^\circ C$ and 230 bar), as published in International Steam Tables and by the US National Institute. The volumetric flow rate was 12.184 mL/min, being: flow rates: H_2O_2 (P-01): 8.19 mL/min, catalyst (P-02): 4.023 mL/min, organic (P-03): 0.061 mL/min, NaOH (P-04): 3.4 mL/min. All the chemicals were purchased from Aldrich Ltd and used without further purification.

***CAUTION*: *This type of oxidation is potentially extremely hazardous, and must be approached with care and a thorough safety assessment must be made*.** Before each run, the apparatus was hydrostatically pressure tested when cold, and was then heated with a flow of pure water (5 – 10 mL/min). Once the operating temperature had been reached, the pumps for organic, H_2O_2 solution, catalyst solution $MnBr_2$ (catalyst strength: Mn 1719 ppm/Br 5000 ppm) and NaOH (0.5 mol/L) were started. Typically, an experiment was run for 1 – 4 h. The products were collected for sequential periods of 15 min and analysed. The yield was calculated by HPLC as a percentage of the stoichiometric amount expected from the measured amount of organic pumped into the apparatus. The selectivity was calculated by HPLC as a percentage based on the molar amount of a particular component in the solution recovered from the reactor. A Waters Xterra reverse phase C18 column, maintained at 40 °C, was used (flow rate 0.7 mL/min, run time 14 min; UV detection at 230 nm). Gradient elutions with solvents CH_3CN (16.7%) and CH_3CO_2Na/CH_3CO_2H buffer (83.3 to 60% and back to 83.3%) or the isocratic method CH_3CN (10%) and CH_3CO_2Na/CH_3CO_2H buffer (90%) were used. The stock buffer solution was prepared by dissolving 15 g CH_3CO_2Na anhydrous in 250 mL de-ionised water, before adding CH_3CO_2H (50% v/v, 100 mL). The pH was adjusted to 3.9 ± 0.01 with 5% CH_3CO_2H, before diluting to 500 mL. The dilute buffer was prepared by diluting 30 mL of the stock buffer solution to 500 mL with de-ionised water.[11]

Thermal Stability Study by Raman Spectroscopy

The cell and the experimental set-up has already been described in detail in a recent publication.[24] The cell was constructed from the heat-resistant EI437B Ni-Cr alloy and was designed to withstand pressures of up to 1500 bar at temperatures up to 550 °C. The cell was installed in the sample compartment of the Almega (Thermo Nicolet) Raman spectrometer, exploiting 180° (back-scattered) geometry of scattering to collect spectra. An NIR GaAlAs diode laser (785 nm) was used to excite the samples with ~300 mW output power. The motorised stage of the Almega made it possible to align the cell by using the xyz coordinates *via* the Omnic 6.0a Software (© 1992 – 2000 ThermoNicolet Corp.). A 100 μm slit provided sufficient resolution. Each spectrum was collected over four scans of which each had an exposure time of five seconds. For each experiment, spectra were collected at regular time intervals, typically during a 30 min period.

Reagent grade samples of the acids were purchased from Aldrich and were used as received. NaOH was obtained in pellet form. After ensuring the alignment was complete and that the highest S/N ratio was achieved for the spectrum of the solution under question, the cell was heated up to the desired temperature and pressure. After thermal equilibrium was reached, a set of spectra was obtained with a GaAlAs diode laser (excitation at 785 nm), with an output of *ca.* 300 mW. Each trace was the sum of 16 scans, at 10 seconds each, using a 100 μm slit. At the end of each experiment, the heater was switched off and the cell was allowed to cool down to ambient temperature. The system was then depressurised and the contents were collected for high performance liquid chromatography (HPLC) analysis.

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References and Notes

[1] W. Partenheimer, *Catal. Today* **1995**, *23*, 69.
[2] R. Noyori, M. Aoki, K. Sato, *Chem. Commun.* **2003**, 1977.
[3] a) I. V. Zakharov, Y. V. Geletii, V. A. Adamyan, *Kinet. i Katali* **1991**, *32*, 31; b) M. P. Czytko, G. K. Pub, *Ind. Eng. Chem. Prod. Res. Dev.* **1961**, *20*, 481; c) L. J. Csanyi, K. Jaky, *J. Catal.* **1993**, *41*, 721.
[4] a) V. N. Aleksandrov, *Kinetika i Kataliz* **1979**, *20*, 357; b) K. Ukegawa, Y. Kamiya, *Bull. Chem. Soc. Jap.* **1976**, *49*, 1632.
[5] N. R. Nowicki, J. D. Lowry, *US Patent 4,892,970*, **1990**.
[6] N. Akiya, P. E. Savage, *Chem. Rev.* **2002**, *102*, 2725.
[7] M. Uematsu, E. U. Franck, *J. Phys. Chem. Ref. Data* **1980**, *9*, 1291.
[8] a) R. C. Weast, *CRC Handbook of Chemistry and Physics: a ready reference book of chemical and physical*

data, CRC Press, Ohio, **1972**; b) W. L. Marshall, E. U. Franck, *J. Phys. Chem. Ref. Data* **1981**, *10*, 295.

[9] a) M. J. Antal, M. Carlsson, X. Xu, D. G. M. Anderson, *Ind. Eng. Chem. Res.* **1998**, *37*, 3820; b) H. P. Lesutis, R. Gläser, C. L. Liotta, C. A. Eckert, *Chem. Commun.* **1999**, 2063; c) Y. Ikushima, K. Hatakeda, O. Sato, T. Yokoyama, M. Arai, *J. Am. Chem. Soc.* **2000**, *122*, 1908; d) N. Akiya, P. E. Savage, *Ind. Eng. Chem. Res.* **2001**, *40*, 1822; e) H. Oka, S. Yamago, J. Yoshida, O. Kajimoto, *Angew. Chem. Int. Ed.* **2002**, *41*, 623; f) L. M. Dudd, E. Venardou, E. Garcia-Verdugo, P. Licence, A. J. Blake, C. Wilson, M. Poliakoff, *Green Chemistry* **2003**, *5*, 187.

[10] H. Schmieder, J. Abeln *Chem. Eng. Technol* **1999**, *22*, 903.

[11] P. A. Hamley, T. Ilkenhans, J. M. Webster, E. Garcia-Verdugo, E. Vernardou, M. J. Clarke, R. Auerbach, W. B. Thomas, K. Whiston, M. Poliakoff, *Green Chemistry* **2002**, *4*, 235.

[12] J. B. Dunn, P. E. Savage, *Green Chemistry* **2003**, *5*, 649.

[13] a) P. Wiseman, *An Introduction to Industrial Organic Chemistry*, Applied Science Publishers, London, 2nd edn., **1979**; b) K. Weissermel, H.-J. Arpe, *Industrial Organic Chemistry*, Verlag Chemie, Weinheim, New York, 3rd edn., **1997**.

[14] W. Partenheimer, J. A. Kaduk, *Deactivation Mechanisms in Liquid Phase Oxidation Caused by Carboxylic Acids*, In: *Proceedings from the 4th International Symposium on Activation of Dioxygen and Homogeneous Catalytic Oxidation*, 613–621 Elsevier, New York.

[15] More detail in the intermediate products is given in: W. Partenheimer, *J. Mol. Catal.* **2003**, *206*, 105.

[16] J. B. Dunn, D. I. Urquhart, P. E. Savage, *Adv. Synth. Catal*. **2002**, *344*, 385.

[17] a) R. L. Holliday, B. Y. M. Jong, J. W. Kolis, *J. Supercritical Fluids*, **1998**, *12*, 255; b) Y.-L. Kim, J.-D. Kim, J. S. Lim, Y.-W. Lee, S.-C. Yi, *Ind. Eng. Chem. Res.* **2002**, *41*, 5576.

[18] N. R. Nowicki, J. D. Lowry, *US Patent 4,892,970*, **1990**.

[19] W. Partenheimer, *A Method for the Preparation of Pyromellitic Acid from Durene via Liquid Phase Oxidation.* In: D. W. Blackburn (Ed.), *Catalysis of Organic Reactions*, Marcel Dekker, New York, **1994**, chap. 14.

[20] S. S. Lande, J. K. Kochi, *J. Am. Chem. Soc*. **1968**, *90*, 5196.

[21] J. M. Anderson, J. K. Kochi. *J. Am. Chem. Soc*. **1970**, *92*, 2450.

[22] R. A. Sheldon, J. K. Kochi, *Metal Catalyzed Oxidations of Organic Compounds*, Academic Press, New York, **1981**, p 360.

[23] J. B. Dunn, M. L. Burns, S. E. Hunter, P. E. Savage, *J. Supercritical Fluids* **2003**, *27*, 263.

[24] Y. E. Gorbaty, G. V. Bondarenko, E. Venardou, E. Garcia-Verdugo, M. Sokolova, J. Ke, M. Poliakoff, *Appl. Spectros.* **2003**, *57*, 1300.

[25] W. Partenheimer, *J. Mol. Catal.* **2001**, *174*, 29.

[26] a) P. D. Metelski, V. A. Adamian, J. H. Espenson, *Inorg. Chem.* **2000**, *39*, 2434; b) X.-D. Jiao, J. H. Espenson, *Inorg. Chem.*, **2000**, *39*,1549.

[27] W. Partenheimer, V. V. Grushin, *Adv. Synth. Catal*. **2001**, *343*, 102.

[28] a) W. Partenheimer, *J. Mol. Catal.* **2003**, *206*, 105; b) . W. Partenheimer *J. Mol. Catal.* **2003**, *206* , 131.

[29] J. B. Dunn, P. E. Savage, *Ind. Eng. Chem. Res.* **2002**, *41*, 4460.

[30] G. H. Jones, *J. Chem. Soc. Chem. Commun*. **1979**, 536.

[31] M. M. Hoffmann, J. G. Darah, B. J. Palmer, J. L. Fulton, *J. Phys. Chem. A* **1999**, *103*, 8471.

[32] J. Burgess, *Metals Ions in Solution*, Wiley, Chichester, New York, **1978**.

[33] a) K. Sawada, M. Tanaka, *J. Inorg. Nucl. Chem*. **1997**, *39*, 339; b) K. Sawada, T. Onoda, T. Suzuki, *J. Chem. Soc. Dalton Trans.* **1983**, 1565.

[34] A. I. Popov, in: *The Chemistry of Nonaqueous Solvents III*, (Ed.: J. J. Lagowski), Academic Press, New York, **1970**, p 2145.

[35] H. Tamiyasu, S. Y. Kim, T. Tsukahara, M. Harada, *Progress in Nuclear Energy* **2000**, *37*, 411.

[36] J. A. Darr, M. Poliakoff, *Chem. Rev*. 1999; 99, 495.

[37] B. D. Phenix, J. L. DiNaro, J. W. Tester, J. B. Howard, K. A. Smith, *Ind. Eng. Chem. Res.* **2002**, *41*, 624.

[38] See the paper by W. Partenheimer in this issue: W. Partenheimer, *Adv. Synth. Catal.* **2004**, *346*, 297–306.

[39] Observations from the laboratory of W. Partenheimer.