Stages in Oxidations of Organic Compounds by Potassium Permanganate. Part IV.* Oxidation of Malonic Acid and its Analogues.

### By Alan Y. Drummond and William A. Waters.

The oxidation of malonic acid by manganic pyrophosphate under nitrogen is completed at the stage $CH_2(CO_2H)_2 + 3O = 2CO_2 + H \cdot CO_2H$, tartronic acid being an intermediate. An initial complex, probably $[Mn^{III},CH_2(CO_2H)_2,(H_2P_2O_7)_2]^-$, is first formed and this decomposes by a rate-determining reversible reaction to manganous pyrophosphate and an active radical, $\cdot CH(CO_2H)_2$, capable both of inducing vinyl polymerisation and of oxidising alcohols, ethers, and Mn^{2+} ions. In the presence of oxygen the complete oxidation $CH_2(CO_2H)_2 + 4O = 3CO_2 + 2H_2O$ occurs with consumption of one molecule of oxygen gas per molecule of malonic acid. Oxygen is absorbed by the radical $\cdot CH(CO_2H)_2$, carbon dioxide evolution then sets in, and the molecule $HO \cdot O \cdot CH(CO_2H)_2$ does not seem to be formed. Oxalic acid seems to be an intermediate in this complete oxidation: mesoxalic acid, if formed, would immediately be oxidised to oxalic acid under these conditions.

Both ethyl- and benzyl-malonic acid give radicals $\cdot CR(CO_2H)_2$ which are incapable of oxidising alcohols or Mn^{2+}, but these acids do form hydroperoxides, $HO \cdot O \cdot CR(CO_2H)_2$, of low stability. Weakly acid permanganate does not at once oxidise malonic acid. A trace of Mn^{III} must be present to initiate reaction in which the stage $\cdot CH(CO_2H)_2 + (MnO_4)^- \longrightarrow {}^+CH(CO_2H)_2 + (MnO_4)^{2-}$ is important.

The investigation has disclosed a novel example of a radical-induced oxidation, since Mn^{III} does not attack monohydric alcohols in the absence of malonic acid. Oxidising and reducing powers of radicals are discussed in relation to chemical structure.

THE mechanism of the oxidation of malonic acid by aqueous potassium permanganate has not yet been studied in detail, though several workers have examined the organic products of the oxidation. Whilst Perdrix (*Bull. Soc. chim.*, 1900, **23**, 645), Cameron and McEwan (*Proc.*, 1910, **26**, 144), and Cartledge and Nichols (*J. Amer. Chem. Soc.*, 1940, **62**, 3057) agree that the oxidation in acid solution is

$$CH_2(CO_2H)_2 + 3O = 2CO_2 + H \cdot CO_2H + H_2O$$
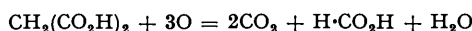
Hatcher and West (*Trans. Roy. Soc. Canada*, 1927, **21**, iii, 269) have argued from kinetic observations of the relative rates of oxidation of formic and malonic acid that the reaction is

$$CH_2(CO_2H)_2 + 2O = CO_2 + 2H \cdot CO_2H$$
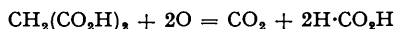
although it is difficult to reconcile this statement with their view that the oxidation proceeds by the stages Malonic acid $\longrightarrow$ Tartronic acid $\longrightarrow$ Glyoxylic acid $\longrightarrow$ CO_2 + Formic acid. This sequence was inferred from the observation that whereas glyoxylic acid is rapidly oxidised by acid permanganate to formic acid and carbon dioxide, mesoxalic acid is completely oxidised to carbon dioxide under similar conditions (Hatcher and Holden, *ibid.*, 1925, **19**, iii, 13; Hatcher, *ibid.*, 1926, **20**, iii, 327). The same sequence also seems to hold for oxidation of malonic acid by hydrogen peroxide (Hatcher and Mueller, *Canad. J. Res.*, 1930, **3**, 291), for its anodic oxidation (Takayama and Miduno, *J. Chem. Soc. Japan*, 1935, **56**, 1460), and for oxidation by periodic acid (Fleury and Courtois, *Compt. rend.*, 1946, **223**, 633; Huebner, Ames, and Bubl, *J. Amer. Chem. Soc.*, 1946, **68**, 1621). Alkaline permanganate however oxidises malonic acid to carbon dioxide and oxalic acid :

$$CH_2(CO_2H)_2 + 3O = CO_2 + H_2C_2O_4 + H_2O$$
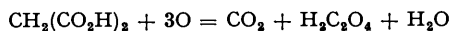
(Drummond and Waters, Part I, *loc. cit.*).

Except for the paper last cited, preceding publications give no indication of the way in which the successive valency changes of the manganese determine the detailed mechanism of the oxidation. The present work attempts to deal with this problem by the selective methods that were broadly reviewed in Part I of this series (*loc. cit.*). Since it was found that manganic salts immediately attack malonic acid, whereas cold acid permanganate reacts with it only after an induction period, the oxidation of malonic acid, and subsequently of ethyl- and benzyl-malonic acid, by manganic pyrophosphate was examined in detail, and thereafter the significance of the manganic ion in initiating reactions involving the permanganate anion was demonstrated. Each aspect of the investigation is reviewed in detail in the following sections.

Oxidation of Malonic Acid with Manganic Pyrophosphate.—The manganic pyrophosphate reagent described in Parts II and III oxidises malonic acid at 25—30° at a rate convenient for kinetic study. In all early experiments, carried out in stoppered vessels, it was found that the initial, fairly rapid reaction slowed down considerably after a few minutes and thereafter proceeded steadily for a long time (Fig. 1, curve *A*). All components of the original reaction mixture, including distilled water, brought about a further brief acceleration of the rate of oxidation when added separately to the slowly oxidising system and it was then realised that dissolved oxygen was accelerating the oxidation. Further

FIG. 1. *Oxidations of malonic acid by manganic pyrophosphate.* pH 1·32; temp. 30·0°.

Initial concns.: Mn^{III} 2·00 × 10^{-2}M; *malonic acid,* 9·26 × 10^{-2}M; *pyrophosphate* 0·169M.
A, Reaction in closed vessel.
B, Reaction under nitrogen.
*C, Reaction under nitrogen*; 4·00 × 10^{-3}M-*manganous sulphate added.*
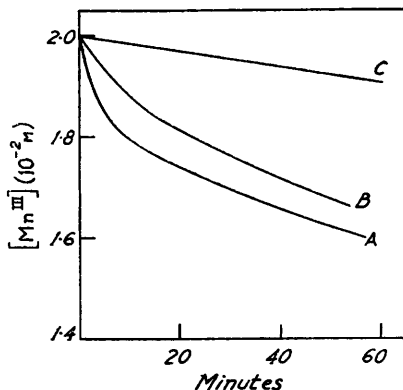
experiments revealed that the oxidations of malonic acid (A, see below) in the absence of oxygen and (B) in aerated solutions proceed by quite different reaction mechanisms.

(A) *Reactions in the absence of oxygen.* These were carried out at 30° in stirred solutions through which purified nitrogen was continually bubbled. The rate of reduction of Mn^{III} was easily followed by iodometric titration. Under these conditions the rate of reaction had still decreased considerably when only 10—20% of the manganic salt had been reduced (see Fig. 1, curve *B*). This was due to the retarding effect of manganous ions, for the initial rate of reduction of manganic salt could be decreased and made linear by adding manganous sulphate equivalent to 25% of the initial Mn^{III} (Fig. 1, curve *C*). With such mixtures it was possible to determine initial reaction orders with respect to the various components of the system. Other bivalent cations, such as magnesium and zinc which also form complexes with pyrophosphate, had no corresponding action.

Possible explanations of the retarding action of manganous ions are (*a*) that the active oxidising agent is in this case Mn^{4+}, produced by the rapid reversible disproportionation

$$2Mn^{3+} \rightleftharpoons Mn^{2+} + Mn^{4+} \quad \cdot \quad \cdot \quad \cdot \quad \cdot \quad \cdot \quad \cdot \quad \cdot \quad \cdot \quad (1)$$

and (*b*) that the initial stage of the oxidation effected by the manganic ion is reversible

$$Mn^{3+} + CH_2(CO_2H)_2 \rightleftharpoons Mn^{2+} + H^+ + \cdot CH(CO_2H)_2 \quad \cdot \quad \cdot \quad \cdot \quad \cdot \quad (2)$$
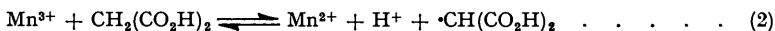
Hypothesis (*a*) is most improbable since it has been shown that an excess of pyrophosphate so strongly displaces the equilibrium (1) in favour of tervalent manganese complexes that there is no evidence of the existence of any quadrivalent manganese complexes in aqueous solutions of moderate acidity (Kolthoff and Watters, *Ind. Eng. Chem. Anal.*, 1943, **15**, 8;

Lingane and Karplus, *ibid.*, 1946, **18**, 191; Watters and Kolthoff, *J. Amer. Chem. Soc.*, 1948, **70**, 2455). Again, manganic pyrophosphate does not attack compounds such as ethanol and *iso*propanol which are thought to be oxidised by Mn^{4+} (Merz, Stafford, and Waters, *J.*, 1951, 638). Moreover equation (1) cannot be reconciled with the observed reaction orders.
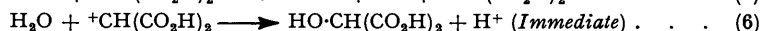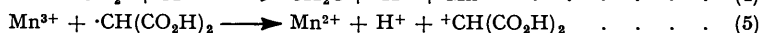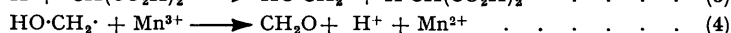
The reversibility of reaction (2) is however in full accord with experimental facts, since (i) the oxidation rate is considerably accelerated by adding vinyl cyanide, which removes the initial radical $\cdot CH(CO_2H)_2$ as an insoluble polymer containing carboxyl groups (Drummond and Waters, *J.*, 1953, 2836), and (ii) the oxidising mixture has been proved to contain a new potent oxidising agent capable of attacking monohydric alcohols and ethers which are unaffected by manganic pyrophosphate. These coupled oxidations immediately increase the rate of reduction of Mn^{III} and must therefore be related to the first stage of the oxidation of the malonic acid, and not to the oxidation of later reaction products such as tartronic acid. Table 1 lists, in increasing order of their accelerating effects, the substances which have been tested. It is significant that the oxidisable com-

TABLE 1. *Organic compounds affecting the rate of reaction of malonic acid with manganic pyrophosphate under nitrogen.*

Substances without effect: Acetic, succinic, and adipic acid; ethyl acetate; chloral hydrate; *tert.*-butanol.

Substances accelerating reduction of Mn^{III}: Methanol (gives formaldehyde) < dioxan < formic acid < formaldehyde < tetrahydropyran < diethyl ether (gives acetaldehyde) < vinyl cyanide (gives polymer) < ethanol (gives acetaldehyde) < propanol (gives propaldehyde) < *iso*propanol (gives acetone) ≪ crotonic acid < maleic and fumaric acid.

pounds in the second section of Table 1 are oxidised by Fenton's reagent by the chain mechanism (Merz and Waters, *J.*, 1949, S 15) and therefore yield strongly reducing radicals of low redox potential, $E_{(-e)}$ (Mackinnon and Waters, *J.*, 1953, 323). The induced oxidations can therefore be ascribed to hydrogen-atom transfers, such as (3), followed by fast oxidations (4) which compete effectively with the much slower oxidation (5) of the malonic acid radical :

$$HO\cdot CH_2\text{-}H + \cdot CH(CO_2H)_2 \longrightarrow HO\cdot CH_2\cdot + H\cdot CH(CO_2H)_2 \quad \ldots \quad (3)$$

$$HO\cdot CH_2\cdot + Mn^{3+} \longrightarrow CH_2O + H^+ + Mn^{2+} \quad \ldots \quad (4)$$

$$Mn^{3+} + \cdot CH(CO_2H)_2 \longrightarrow Mn^{2+} + H^+ + {}^+CH(CO_2H)_2 \quad \ldots \quad (5)$$

$$H_2O + {}^+CH(CO_2H)_2 \longrightarrow HO\cdot CH(CO_2H)_2 + H^+ \; (\textit{Immediate}) \quad \ldots \quad (6)$$

Theoretical implications of these radical-coupled oxidations are examined below (p. 2466).

In prolonged oxidations of malonic acid by manganic pyrophosphate it was found that carbon dioxide was evolved and that formic acid remained in solution after complete reaction. When excess of manganic salt was present a total of 5·95 equivalents of Mn^{III} were eventually destroyed per molecule of malonic acid in the complete absence of oxygen. Consequently the equation $CH_2(CO_2H)_2 + 3O = 2CO_2 + H\cdot CO_2H + H_2O$ adequately represents the stoicheiometry of the oxidations of malonic acid by both cold acid permanganate and manganic cations. Tartronic, mesoxalic, glycollic, glyoxylic, and oxalic acids are all possible intermediates. Of these, mesoxalic and oxalic acids can at once be excluded since manganic pyrophosphate oxidises them completely to carbon dioxide. Glycollic acid too can be excluded since it is oxidised much more slowly than is the malonic acids : had it been formed the oxidation rate would have decreased noticeably when only two equivalents of Mn^{III} had been consumed, and this is not the case. Both tartronic and glyoxylic acids are oxidised rapidly by manganic pyrophosphate and both yield formic acid as well as carbon dioxide. Consequently the complete course of the oxidation by Mn^{III} can be written :
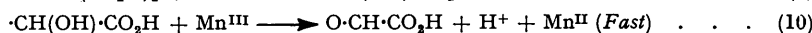
$$CH_2(CO_2H)_2 \longrightarrow HO\cdot CH(CO_2H)_2 \longrightarrow CO_2 + OCH\cdot CO_2H \longrightarrow CO_2 + H\cdot CO_2H$$
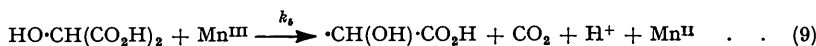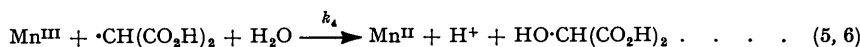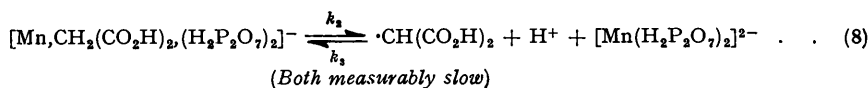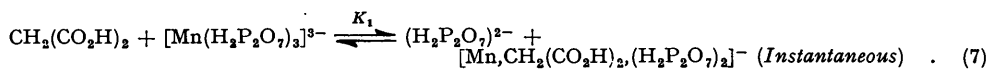
Two equivalents of Mn^{III} are needed in each of these three stages and presumably therefore each of them involves the transient formation of an organic free radical.

The initial rates of oxidation indicate a reaction order with respect to malonic acid of

rather less than one, and this, together with the fact that, at constant pH, the rate of oxidation decreases when the pyrophosphate concentration is increased, indicates that the rate-determining oxidation stage is preceded by the rapid reversible formation of a chelated manganic complex, similar to that which is involved in the oxidation of pinacol (Part III). Malonate complexes of tervalent manganese have been described by Cartledge and Nichols (*loc. cit.*), Taube (*Chem. Reviews*, 1952, **50**, 69; *J. Amer. Chem. Soc.*, 1948, **70**, 3926), and Meyer and Schramm (*Z. anorg. Chem.*, 1921, **123**, 56). These workers have shown that the chelate association is labile and that the complexes soon decompose in acidified solutions.

By neglecting the possible formation of pyrophosphate–dimalonato-complexes when pyrophosphate is present in large excess and also the formation of complexes in which tartronic acid has displaced malonic acid, and further by assuming that the malonate radical $\cdot CH(CO_2H)_2$ reacts with any Mn^{III} ion without prior complex formation, the kinetics of the initial stages of the oxidation can be evaluated from the reaction sequences (7, 8, 5, 6, 9, 10) in which one typical pyrophosphate ion has been selected as representative of those concerned in oxidations throughout the whole of the possible pH range.

$$CH_2(CO_2H)_2 + [Mn(H_2P_2O_7)_3]^{3-} \xrightleftharpoons{K_1} (H_2P_2O_7)^{2-} + [Mn,CH_2(CO_2H)_2,(H_2P_2O_7)_2]^- \ (\textit{Instantaneous}) \quad . \quad (7)$$

$$[Mn,CH_2(CO_2H)_2,(H_2P_2O_7)_2]^- \underset{k_3}{\overset{k_2}{\rightleftharpoons}} \cdot CH(CO_2H)_2 + H^+ + [Mn(H_2P_2O_7)_2]^{2-} \quad . \quad . \quad (8)$$

(Both measurably slow)

$$Mn^{III} + \cdot CH(CO_2H)_2 + H_2O \xrightarrow{k_4} Mn^{II} + H^+ + HO\cdot CH(CO_2H)_2 \quad . \quad . \quad . \quad (5, 6)$$

$$HO\cdot CH(CO_2H)_2 + Mn^{III} \xrightarrow{k_5} \cdot CH(OH)\cdot CO_2H + CO_2 + H^+ + Mn^{II} \quad . \quad . \quad (9)$$

$$\cdot CH(OH)\cdot CO_2H + Mn^{III} \longrightarrow O\cdot CH\cdot CO_2H + H^+ + Mn^{II} \ (\textit{Fast}) \quad . \quad . \quad (10)$$

The oxidation of the tartronic acid will undoubtedly involve formation and breakdown of a chelated complex, but since it is faster than the oxidation of malonic acid the simplified equation (9) adequately represents its significant oxidation stage.

From this scheme, by assuming that the radical $\cdot CH(CO_2H)_2$ attains a stationary concentration, it can be deduced that the initial rate

$$-d[Mn^{III}]/dt = 2K_1 \cdot k_2 k_4 \cdot \frac{[CH_2(CO_2H)_2]}{[(H_2P_2O_7)^{2-}] + K_1[CH_2(CO_2H)_2]} \cdot \frac{[Mn^{III}]^2 + k_5[Mn^{III}]^3 \cdot t}{k_3[Mn^{II}] + k_4[Mn^{III}]}$$

where the term $k_5[Mn^{III}]^3 \cdot t$ represents the small additional oxidation of the tartronic acid present at time t after the start of the reaction, and is a first-order approximation that is valid provided that the concentration of the tartronic acid in the system is low.

Fig. 2 shows that the initial value of $-d[Mn^{III}]/dt$ varies as $a[\text{Malonic acid}]/(b + [\text{Malonic acid}])$ when the concentration of *free* malonic acid $[CH_2(CO_2H)_2]$ has been computed from the initial total concentration of malonic acid by the approximation method described in Part III. Figs. 3 and 4 show that the expected reaction orders with respect to both $[Mn^{III}]$ and $[Mn^{II}]$ hold, provided that $k_3 \gg k_4$, which is necessarily the case since a small amount of manganous salt has a very noticeable effect on the reaction velocity (compare Fig. 1, curves *B* and *C*). It may be noted that since $k_3 \gg k_4$ the reaction order with respect to $[Mn^{III}]$ never falls below *two*, and that it rises beyond this only under the conditions in which the term $k_5[Mn^{III}]^3 \cdot t$ becomes of significance, *i.e.*, at high concentrations of manganic salts.

If glycollic acid, and not tartronic acid, had been an intermediate product, through the occurrence of reaction (11), then at high concentrations of Mn^{III} the reaction order with

$$\cdot CH(CO_2H)_2 + Mn^{III} + H_2O \longrightarrow HO\cdot CH_2\cdot CO_2H + H^+ + CO_2 + Mn^{II} \quad . \quad . \quad (11)$$

respect to $[Mn^{III}]$ should fall below two instead of rising above it, because the rate of oxidation of glycollic acid is negligible.

(B) *Reactions in the presence of oxygen.* Oxygen accelerates the rate of oxidation of malonic acid by manganic pyrophosphate so much that comparative reaction velocities

could be obtained only by lowering the temperature to 25° and decreasing the acidity to the pH range 2·0—2·6. Since many organic free radicals combine with oxygen to give radicals which then form hydroperoxides, RO_2H, of moderate stability, or sometimes hydrogen peroxide, both of which would invalidate iodometric titration of Mn^{III}, colormetric measurements were made which showed that, within their accuracy, oxidising

Fig. 2. *Reaction orders with respect to malonic acid.*
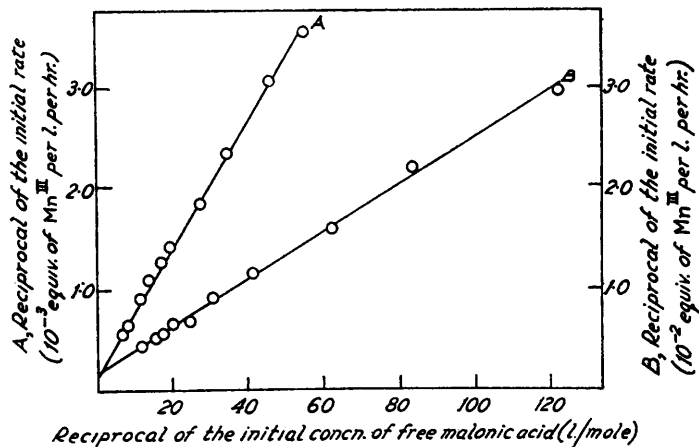
A, *Reaction in nitrogen* : pH 0·92; temp. 30·0°.
Initial concns.: Mn^{III} 1·67 × $10^{-2}M$; *manganous sulphate*, 4·18 × $10^{-3}M$; *pyrophosphate*, 0·146M.
B, *Reaction in oxygen* : pH 2·05; temp. 25·0°.
Initial concns.: Mn^{III} 2·07 × $10^{-2}M$; *pyrophosphate*, 0·178M.

Fig. 3. *Oxidation of malonic acid under nitrogen; reaction order with respect to* $[Mn^{III}]$.
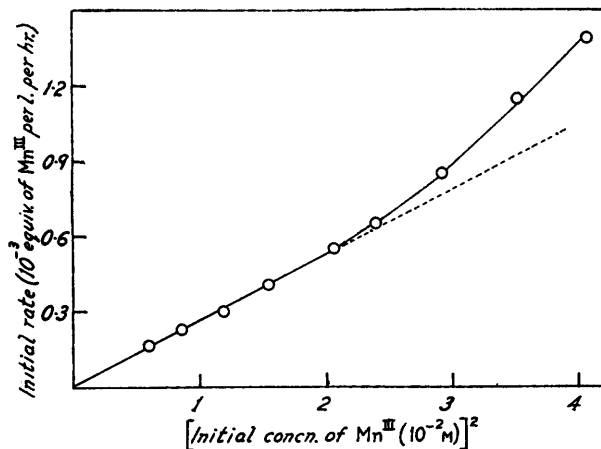
pH 0·80; *temp.* 30·0°. *Initial concns.*: *malonic acid*, 7·31 × $10^{-2}M$; Mn^{II}, 5·83 × $10^{-3}M$; *pyrophosphate* 0·134M.

solutions contain negligible concentrations of hydroperoxides. Tests with titanic sulphate confirmed the absence of hydrogen peroxide after complete reaction.

Solutions of malonic acid which have been completely oxidised by the combined action of manganic pyrophosphate and oxygen do not give positive chromotropic acid tests for the presence of formic acid (Feigl, " Spot Tests," Elsevier Publ. Co., Amsterdam, 1947), so that evidently the oxygen must alter the whole reaction sequence, preventing the formation of tartronic or glyoxylic acid and assisting complete oxidation to carbon dioxide

and water. The oxygen undoubtedly acts at an early stage in the oxidation sequence because its absorption sets in immediately at its maximum rate. Moreover, the rates of the ærobic oxidations, unlike those of anærobic oxidations, are not altered by the addition of manganous cations, but can be reduced slightly by adding high concentrations of methanol or ethanol. Carbon dioxide is evolved from the commencement of the oxidation and consequently it was necessary to use Warburg micro-respirometers containing caustic alkali for measurement of oxygen uptake.

Table 2, in which the oxygen uptake of a reacting mixture is compared with the extent of reduction of Mn^{III}, shows that the two reactions proceed at related speeds; this would be the case if the simplified reaction (2) (p. 2457) was immediately followed by (12) to the complete exclusion of reactions (5), (6), (9), and (10) of the anærobic oxidation sequence, for the ratio $R = $ (Moles of O_2 absorbed)/(Equivs. of Mn^{III} reduced) remains constant though the oxidation rate decreases as the manganic salt is consumed.

$$O_2 + \cdot CH(CO_2H)_2 \longrightarrow \cdot O \cdot O \cdot CH(CO_2H)_2 \quad . \quad . \quad . \quad . \quad . \quad . \quad (12)$$
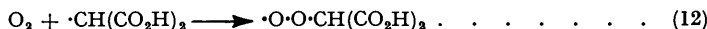
Table 3 gives final values for this ratio R after complete reduction of Mn^{III} for different mixtures all initially containing malonic acid in excess. The bigger the initial excess of malonic acid the more will the measurements relate to the early stages (2) and (12) of the ærobic oxidation : as expected this condition favours high values of R. R also increases when the rate of oxidation is diminished by increasing the pH of the solution.

TABLE 2. *Oxidation of malonic acid by the combined action of manganic pyrophosphate and oxygen.* Temp. $35.08°$; pH 2.35.

Initial concns. (M) : Mn^{III} 1.63×10^{-2}; malonic acid 2.19×10^{-2}; pyrophosphate 1.41×10^{-1}.

Time (min.)	Oxygen absorbed (10^{-3} mole/l.)	Mn^{III} reduced (10^{-3} equiv./l.)	Ratio, R	Time (min.)	Oxygen absorbed (10^{-3} mole/l.)	Mn^{III} reduced (10^{-3} equiv./l.)	Ratio, R
2	0.19	0.65	0.285	40	2.99	11.1	0.269
4	0.43	1.40	0.307	50	3.35	12.4	0.270
6	0.56	2.12	0.264	60	3.61	13.4	0.269
10	0.97	3.59	0.270	70	3.79	14.1	0.269
14	1.37	5.00	0.275	80	3.91	14.6	0.268
20	1.84	6.82	0.270	90	4.03	15.0	0.269
28	2.36	8.79	0.268	∞	4.47	16.3	0.274

TABLE 3. *Completed oxidations of malonic acid by Mn^{III} in the presence of oxygen.* Temp. $35.08°$.

Initial concentrations				Ratio, R	
Mn^{III} (10^{-2}M)	Pyrophosphate (M)	Malonic acid (M)	pH	Final value	Standard deviation (%)
1.20	0.154	0.795	1.15	0.424	1.0
1.20	0.154	0.144	2.41	0.366	1.5
1.20	0.154	0.144	1.15	0.308	0.7
1.79	0.154	0.144	1.15	0.280	0.2
1.20	0.103	0.072	0.96	0.281	0.5
1.20	0.154	0.024	1.15	0.283	0.5

A lower limit of $R = 0.25$ should be reached if complete oxidation of each mole of malonic acid to carbon dioxide and water were attained by the consumption of one mole of oxygen and four equivs. of Mn^{III} :

$$CH_2(CO_2H)_2 + O_2 + 4Mn^{3+} = 3CO_2 + 4H^+ + 4Mn^{2+}$$
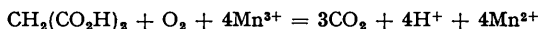
It was difficult however to ensure complete oxidation of malonic acid even by excess of oxygen-saturated solutions of manganic pyrophosphate, for after 12—48 hr. the insoluble red manganic hydrogen phosphate described in Part I is deposited and reaction ceases though only 3.5—3.8 equivalents of Mn^{III} have then been consumed per mole of malonic acid. Since formic acid could not be detected in the resulting solutions it is unlikely that the anærobic sequence had comprised 10—25% of the whole reaction. A more probable explanation is that at low manganic concentrations one of the later stages of the ærobic

oxidation sequence becomes so slow that oxidation to carbon dioxide is not complete. We tentatively suggest that a little oxalic acid may remain, for, as Fig. 5 shows, the reaction

$$H_2C_2O_4 + 2Mn^{3+} = 2CO_2 + 2H^+ + 2Mn^{2+}$$
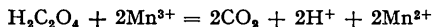
is slow at pH 2·54 in the presence of pyrophosphate, and would cease almost completely when deposition of manganic hydrogen phosphate leaves only a very minute concentration of Mn^{III} in solution. Fig. 5 also shows that the oxidation of mesoxalic acid to oxalic acid is almost instantaneous under similar conditions.
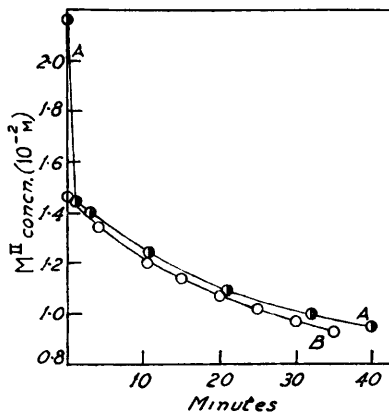
To account for the stages of the aerobic reactions of malonic acid which may follow the

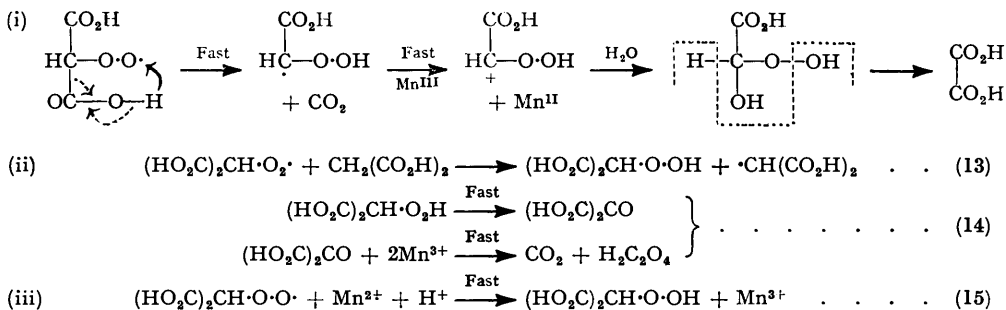Fig. 4. *Oxidation of malonic acid under nitrogen; reaction order with respect to* Mn^{II}.

Fig. 5. *Oxidations of mesoxalic and oxalic acids.*
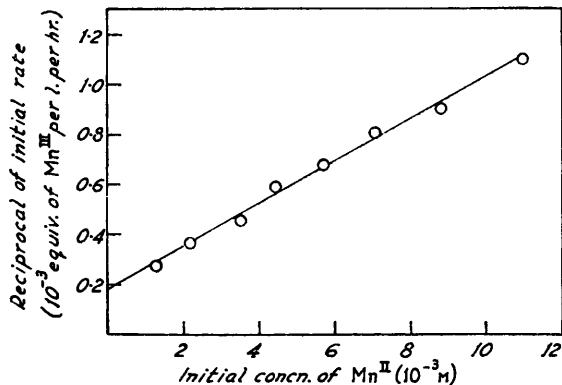
pH 0·79; *temp.* 30·0°. *Initial concns.*: Mn^{III}, 1·76 $\times$ $10^{-2}M$; *malonic acid*, 8·52 $\times$ $10^{-2}M$; *pyrophosphate*, 0·153M.

A, Mesoxalic acid, 3·52 $\times$ $10^{-3}M$, *at* pH 2·51. *B, Oxalic acid,* 3·52 $\times$ $10^{-3}M$, *at* pH 2·54. *Temp.* 30·0°. *Pyrophosphate,* 0·179M.

production of the peroxy-radical by reaction (12) the following three mechanisms are rational :

(i)

(ii) $(HO_2C)_2CH \cdot O_2 \cdot + CH_2(CO_2H)_2 \longrightarrow (HO_2C)_2CH \cdot O \cdot OH + \cdot CH(CO_2H)_2$ . . (13)

$(HO_2C)_2CH \cdot O_2H \xrightarrow{\text{Fast}} (HO_2C)_2CO$
$(HO_2C)_2CO + 2Mn^{3+} \xrightarrow{\text{Fast}} CO_2 + H_2C_2O_4$ $\Big\}$ . . . . (14)

(iii) $(HO_2C)_2CH \cdot O \cdot O \cdot + Mn^{2+} + H^+ \xrightarrow{\text{Fast}} (HO_2C)_2CH \cdot O \cdot OH + Mn^{3+}$ . . . . (15)

immediately followed by (14).

Scheme (i), not involving the formation of a hydroperoxide molecule, postulates an intramolecular hydrogen switch and immediate loss of carbon dioxide to give a carbon radical which by further oxidation by one equivalent of Mn^{III} would give oxalic acid. The rate of oxidation then depends essentially on the rate of formation of the initial radical, $\cdot CH(CO_2H)_2$.

Scheme (ii) requires that the unknown hydroperoxide of malonic acid, formed by the usual metallic ion-catalysed autoxidation chain [reactions (12) and (13) consecutively], breaks down at once to mesoxalic acid which, as has been shown, is immediately oxidised to oxalic acid.

Scheme (iii), an alternative chain sequence for manganic salt-catalysed autoxidation of malonic acid to mesoxalic acid, is the least likely since Mn^{2+} has been shown to have no effect on the reaction rate, but it cannot be quite excluded since the ensuing rapid oxidation of mesoxalic acid to oxalic acid might mask any observable effect of change of Mn^{2+} concentration. This postulate would require reaction (15) to be so fast that it would not be rate-significant provided that a trace of Mn^{2+} was present in the solution. Reaction (15) however does seem to be important in the ærobic oxidations of both ethyl- and benzyl-malonic acids (see p. 2465).

The chain oxidation (ii) also does not appear to be of great significance because, as Table 2 clearly shows, the rate of oxygen uptake is, from the first, strictly proportional to the rate of Mn^{III} reduction and the ratio R does not fall from the high value of 0·5 as would be the case if the initial catalytic formation of the radical $\cdot CH(CO_2H)_2$ (reaction 2) initiated the chain of reactions (12) and (13), giving the hydroperoxide molecule which then consumed two equivalents of Mn^{III} (reaction 14). Again, long reaction chains do not occur because the rate of consumption of oxygen, or of Mn^{III}, does not accord with the theoretical kinetics for such processes {the possibilities are $-d[O_2]/dt = $ approx. $- 2d[Mn^{III}]/dt$ varying as $[Mn^{III}][CH_2(CO_2H)]^2/[Mn^{II}]$ for chain ending by reaction (15), or as $[O_2][CH_2(CO_2H)_2]$ for chain ending by reaction (5) of the anærobic sequence}, but the rate is undoubtedly governed by the forward reactions (7) and (8) between malonic acid and manganic pyrophosphate which require the initial reversible formation of a chelated manganic complex. As in the anærobic oxidation $1/$(Initial rate) varies as $a[$Malonic acid$]/(b + [$Malonic acid$])$ (see Fig. 2). The initial reaction order with respect to Mn^{III} is 1·10 and not exactly unity (see Table 4). This can be expected when the secondary oxidations are faster than the initial step [reactions (7) and (8)].

Scheme (i), not requiring the formation of a hydroperoxide molecule, $HO_2 \cdot CH(CO_2H)_2$, provides, on balance, the most plausible explanation of the available experimental facts, but if very short reaction chains only are involved then scheme (ii) or (iii) may still be cogent. Undoubtedly the uptake of oxygen by the radical $\cdot CH(CO_2H)_2$ [reaction (12)] must be much more rapid than the oxidation of manganous ions by this radical and further reactions of the radical $\cdot O \cdot O \cdot CH(CO_2H)_2$ must be intimately coupled with reactions which consume more Mn^{III}.

The slight retarding actions of methanol and ethanol can easily be accounted for by concluding that the hydrogen-transfer reactions (3) (p. 2458) occur at rates approaching that of (12) and give radicals, $\cdot CH_2 \cdot OH$, etc., which are less reactive to oxygen than is $\cdot CH(CO_2H)_2$.

TABLE 4. *Oxidation of malonic acid in oxygen : reaction order with respect to* Mn^{III}.

A. pH 2·40; temp. 25·0°. Initial [Malonic acid] = 0·0367M. [Pyrophosphate] = 0·189M.
B. pH 2·47; temp. 25·0°. Initial [Malonic acid] 0·0347M. [Pyrophosphate] 0·238M.

Initial rate (10^{-3} equiv. of Mn^{III} per l. per hr.)	$[Mn^{III}]$ (10^{-3}M)	Rate/ $[Mn^{III}]^{1 \cdot 10}$	Initial rate (10^{-3} equiv. of Mn^{III} per l. per hr.)	$[Mn^{III}]$ (10^{-3})	Rate/ $[Mn^{III}]^{1 \cdot 10}$
3·82	6·7	941	2·60	7·3	571
4·65	8·1	930	3·19	8·8	582
6·10	10·8	893	3·84	10·3	589
8·06	13·5	918	4·35	11·7	580
9·90	16·2	917	4·99	13·2	583
10·7	17·5	915	5·71	14·7	592
			6·19	16·1	584
			6·83	17·6	581
			7·38	19·1	571
			8·22	20·5	591
			8·85	22·0	590

Oxidations of Ethylmalonic and Benzylmalonic Acids by Manganic Pyrophosphate.—The oxidations of these two substituted acids were investigated in the hope that the elucidations of their reactions might, by analogy, help to clarify views concerning the secondary stages of both the ærobic and the anærobic oxidation of malonic acid, but since both substituted acids exhibit marked initial differences in behaviour from malonic acid an exhaustive experimental study has not been made.

Oxidations under nitrogen occur more rapidly with the substituted acids (Table 5) and are not affected by the addition of manganous ions. Again, methanol and ethanol have no effect on the rates of oxidation of the substituted acids, and do not produce solutions

TABLE 5. *Comparative rates of oxidation of substituted malonic acids under nitrogen.*

pH 1·40; temp. 35·0°. Initial concns.: Mn^{III} 1·91 × $10^{-2}M$; malonic acids 6·30 × $10^{-2}M$; pyrophosphate 1·62 × $10^{-1}M$.

	Decrease in concentration of Mn^{III} (M × 10^{-3})		
Time (min.)	Malonic acid	Ethylmalonic acid	Benzylmalonic acid
10	1·07	3·15	6·15
20	1·90	6·00	10·3
30	2·50	8·45	13·2
40	3·00	10·3	14·8

containing aldehydes. Evidently the radicals $(HO_2C)_2C(R)\cdot$ are much weaker oxidants when R = Et or CH_2Ph than when R = H. However vinyl cyanide, which polymerises, reduces the rates of these oxidations (again an effect opposite to that with malonic acid) so that the independent existence of the above radicals and their further oxidation by
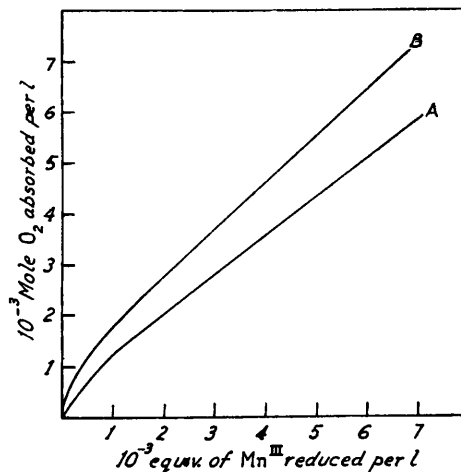

FIG. 6. *Oxidation of benzylmalonic acid under oxygen.*
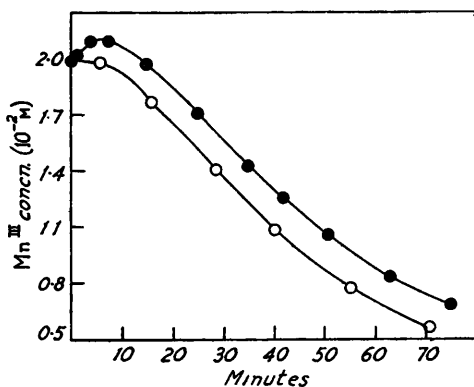
pH 1·39; *temp.* 25·0°. *Initial concns.*: *benzyl-malonic acid*, 6·59 × $10^{-2}M$; Mn^{2+}, 5·38 × $10^{-3}M$; *pyrophosphate*, 0·168M.

● *Iodometry.*
○ *Colorimetry.*

FIG. 7. *Oxidation of ethylmalonic acid under oxygen.*

pH 1·58; *temp.* 35·0°. *Initial concns.*: Mn^{III}, 1·07 × $10^{-2}M$; *ethylmalonic acid*, 4·65 × $10^{-2}M$; *pyro-phosphate*, 0·123M.

A, *No manganous sulphate added.*
B, *Initial concn. of manganous sulphate,* 1·78 × $10^{-3}M$.

Mn^{III} cannot be doubted. In prolonged oxidations of ethylmalonic acid as many as 6—7 equivalents of Mn^{III} can be consumed per mole of acid, which must thus be degraded in part beyond the stage of either methylpyruvic or propionic acid.
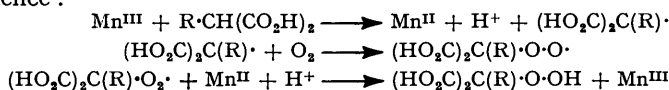
In the presence of oxygen the reduction of Mn^{III} by ethyl- and benzyl-malonic acid is retarded, and not accelerated. Moreover the addition of manganous sulphate then reduces the initial rate of reduction of Mn^{III} so much that there can be no doubt about the formation of hydroperoxides capable of oxidising Mn^{II} to Mn^{III}.

Fig. 6 which contrasts iodometric titrations with colorimetric measurements confirms this deduction for benzylmalonic acid, with which the thiosulphate titre even shows a slight increase during the first few minutes owing to the formation of a fairly constant amount of the hydroperoxide, $Ph\cdot CH_2\cdot C(O_2H)(CO_2H)_2$, which, as well as Mn^{3+}, is capable of oxidising I^- to I_2. Similar, but less pronounced, discrepancies between iodometric and colorimetric measurements have been observed with ethylmalonic acid. Comparative, oxygen uptake measurements with this substance (Fig. 7) show that the addition of manganous salt increases the ratio R (p. 2461), which at first has a very high value indicative of the presence of the hydroperoxide molecule $(HO_2C)_2C(Et)\cdot O\cdot OH$. After about 20 min.

R settles down to values somewhat below 1·0, as the hydroperoxide decomposes to products which are further oxidised by the manganic pyrophosphate. Though it is clear that the rates of formation and destruction of hydroperoxide are in balance for a long time no definite conclusions can, as yet, be drawn concerning the secondary reactions. Methanol and ethanol both retard slightly the rate of consumption of Mn^{III} in these ærobic oxidations.

The essential differences in behaviour between the unsubstituted and the substituted malonic acids in the presence of oxygen can be associated with the relative stabilities of substituted hydroperoxides, which are normally $R \cdot CH_2 \cdot O \cdot OH <$ $R_2CH \cdot O \cdot OH < R_3C \cdot O \cdot OH$. The electrophilic carboxyl groups of malonic acid must reduce the stability of the molecule $(HO_2C)_2CH \cdot O \cdot OH$ so much that its independent existence is in doubt, whereas the tertiary hydroperoxides $(HO_2C)_2CR \cdot O \cdot OH$, in which the electron-donating alkyl group partly neutralises the polar effect of the carboxyl group (see p. 2466), can certainly exist for periods of the order of minutes.

The effect of manganous salts on the ratio R can most simply be accounted for by the reaction sequence :

$$Mn^{III} + R \cdot CH(CO_2H)_2 \longrightarrow Mn^{II} + H^+ + (HO_2C)_2C(R) \cdot$$
$$(HO_2C)_2C(R) \cdot + O_2 \longrightarrow (HO_2C)_2C(R) \cdot O \cdot O \cdot$$
$$(HO_2C)_2C(R) \cdot O_2 \cdot + Mn^{II} + H^+ \longrightarrow (HO_2C)_2C(R) \cdot O \cdot OH + Mn^{III}$$

since there is little fall in Mn^{III} during the first few minutes in which the hydroperoxide concentration is building up to its equilibrium value (see Fig. 6).

This mechanism, which, corresponding to scheme (iii) for the ærobic oxidation of malonic acid, is an oxidation chain with respect to manganese ions, here seems to be more satisfactory than the usual chain reaction (scheme ii) in which the peroxy-radical is the hydrogen abstractor. Since however re-oxidation of Mn^{II} may be due to the subsequent peroxide decomposition

$$(HO_2C)_2C(R) \cdot O \cdot OH + Mn^{2+} \longrightarrow (HO_2C)_2C(R) \cdot O \cdot + Mn^{3+} + OH^- \quad . \quad . \quad (16)$$
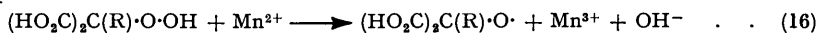
scheme (ii) cannot be rejected decisively.

Oxidation of Malonic Acid by Potassium Permanganate.—The reduction of cold potassium permanganate by malonic acid at pH $\ll 1$ is preceded by a noticeable induction period (cf. Sanyal and Dhar, *Z. anorg. Chem.*, 1924, **139**, 161), but the later course of the oxidation becomes obscured by the separation of brown manganese dioxide though the Guyard reaction

$$2MnO_4^- + 3Mn^{2+} + 2H_2O = 5MnO_2 + 4H^+$$
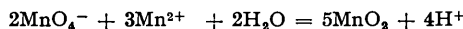
Initial addition of manganous sulphate eliminates the induction period almost completely (Fig. 8, curve B) but hastens also the formation of manganese dioxide. Conversely the addition of either fluoride or pyrophosphate anions greatly prolongs the induction period (Fig. 8, curves C and D) and completely prevents separation of manganese dioxide. Since these anions act by forming complexes with Mn^{2+}, and more especially Mn^{3+}, it can be concluded that in dilute acid the MnO_4^- anion itself does not attack un-dissociated malonic acid; in alkali, however, oxidation of the malonate anion is rapid (see Part I). If, however, manganic pyrophosphate solution is added initially, the acid permanganate oxidation of malonic acid commences immediately (Fig. 8, curve E). Presumably the malonic acid is first attacked by the Mn^{III}, by the reactions elucidated in the preceding sections of this paper, and the resulting organic radical is then rapidly oxidised by a permanganate anion :

$$\cdot CH(CO_2H)_2 + MnO_4^- \longrightarrow {}^+CH(CO_2H)_2 + MnO_4^{2-} \quad . \quad . \quad . \quad . \quad (17)$$

Manganate anions thus formed would very rapidly yield more Mn^{III}, or possibly Mn^{IV}, by the Guyard reaction and the equilibration $Mn^{4+} + Mn^{2+} \rightleftharpoons 2Mn^{3+}$. The two anions which, in the absence of added manganic or manganous salt, prolong the initial induction period are those which remove Mn^{3+} ions as inactive complexes, but since this complex formation is a mobile equilibrium, and not an irreversible process, oxidations through reactions (7) and (8) will eventually occur autocatalytically. As would be expected the autocatalytic oxidation accelerates much more rapidly when pyrophosphate has not been added to the solution.

Oxygen has a slight retarding effect on the reduction of permanganate. Consequently reaction (17) must be fast, though not sufficiently so to prevent the oxygen from deviating a small part of the oxidation of the radical $\cdot CH(CO_2H)_2$ through the path of the aerobic sequence discussed on pp. 2457 and 2466. The latter oxidation must involve a series of reactions which are slower than the permanganate oxidation of the initial radical *via* tartronic and formic acid.

Some Comments on Oxidising and Reducing Powers of Free Radicals.—This investigation has provided a striking instance of a radical-transfer process (reaction 3, p. 2458) whereby one organic compound (malonic acid) can catalyse the oxidation of substances (alcohols, ethers) which are not directly attacked by the oxidising agent at hand (Mn^{III}). Mackinnon and Waters previously (*loc. cit.*) called attention to an oxidation (by $\cdot OH$) whereby a free radical could effect a subsequent reduction, and analgous induced hydrogen-transfers involving free thiyl radicals have been described by Bickel and Kooyman (*Nature*, 1952, **170**, 211), Harris and Waters (*ibid.*, p. 212), and Barrett and Waters (*Discuss. Faraday Soc.*, 1953,

FIG. 8.   *Oxidations of malonic acid by potassium permanganate.*
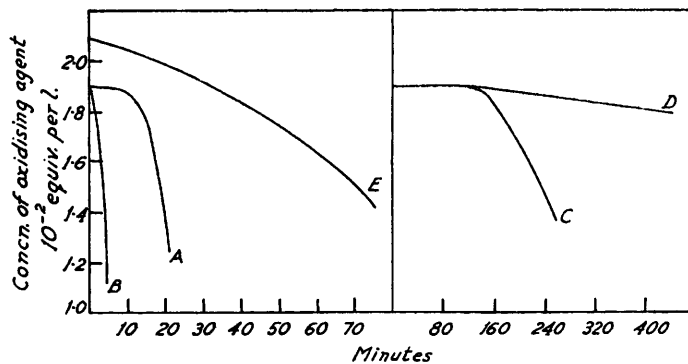
All mixtures had initial concns. of $KMnO_4$ $3 \cdot 81 \times 10^{-3}$M and of malonic acid $1 \cdot 60 \times 10^{-2}$M.   Temp. 20 0°.

A, *No added ions*; pH $1 \cdot 05$.
B, *Manganous sulphate* ($3 \cdot 89 \times 10^{-3}$M) *added*; pH $1 \cdot 05$.
C, *Potassium hydrogen fluoride* ($4 \cdot 99 \times 10^{-2}$M) *added*; pH $1 \cdot 35$.
D, *Sodium pyrophosphate* ($1 \cdot 48 \times 10^{-2}$M) *added*; pH $0 \cdot 96$.
E, *Manganic pyrophosphate* (Mn^{III}, $2 \cdot 16 \times 10^{-3}$M) *added* (*pyrophosphate as in D*); pH $0 \cdot 96$.

14, 221). The concept (Mackinnon and Waters, *loc. cit.*) that two oxidation–reduction potentials, $E_{(-e)}$ and $E_{(+e)}$ may be assigned to any free radical provides a rational explanation of available facts more explicitly than the more recent, though similar, vague suggestion of Kharasch, Zimmermann, Zimmt, and Nudenberg (*J. Org. Chem.*, 1953, **18**, 1045).

Thus the fact that the initial oxidation of malonic acid by manganic pyrophosphate is reversible places $E_{(+e)}$ for the radical $\cdot CH(CO_2H)_2$ at about $+1 \cdot 1$ v under the conditions used. Since this radical can also be oxidised by manganic pyrophosphate, $E_{(-e)}$ must have a slightly lower value. The reducibility of the $\cdot CH(CO_2H)_2$ radical probably explains why malonic acid is not detectably oxidised by hydroxyl radicals and also why it is not autoxidised. Consistent too with this view is the fact that ethyl α-bromomalonate is a typical " positive halogen " compound (Robertson and Waters, *J.*, 1947, 492; Ford and Waters, *J.*, 1952, 2240).

As indicated by earlier suggestions of Merz and Waters (*loc. cit.*) both the $E_{(+e)}$ and the $E_{(-e)}$ values of free radicals depend on the spatial distribution of the odd electron within the radical structure. Thus for malonic acid radical (I) the inductive effect of the two carboxyl groups must bring about absorption of the odd electron into the molecule and so diminish the tendency to further electron loss to a much greater extent than in the alcohol radical (II) ; but this electron drift similarly enhances the tendency to electron gain whereby the resonance-stabilised anion (III) is formed. The added resonance stability of (III) also favours energetically the reaction (I) $+$ e $\longrightarrow$ (III).

The polar control of oxidation–reduction potentials of radicals is further exemplified by the different behaviours of ethyl- and benzyl-malonic acid : the electron-repelling $(+I)$ alkyl groups tend to compensate for the loss of electron density at the point of free valency

(I)   (II)   (III)   (IV)
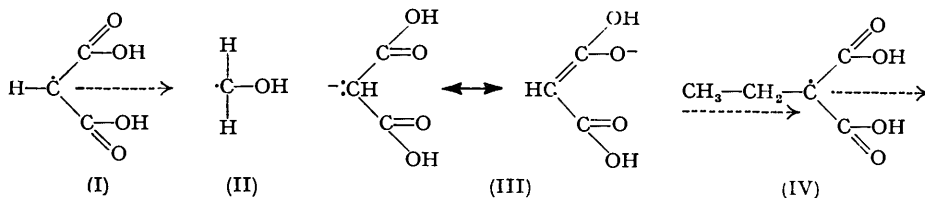
due to the carboxyl groups and thereby diminish the oxidising powers of the radicals $(HO_2C)_2C(Alk)\cdot$ (IV), compared with (I).

EXPERIMENTAL

Materials.—Sodium manganic pyrophosphate solutions were prepared from " AnalaR " reagents as described in Part II. Malonic acid was recrystallised to constant m. p. (134°) from ether–benzene containing 5% of light petroleum (b. p. 40—60°) ; ethylmalonic (m. p. 112°) and benzylmalonic acids (m. p. 120°) were recrystallised from benzene. Sodium mesoxalate was prepared by the method of Chattaway and Harris (*J.*, 1922, 2703), and glyoxylic acid by that of Benedict (*J. Biol. Chem.*, 1909, **6**, 51).

Kinetic Measurements.—The reactions under nitrogen were carried out as described in Part III. With reactions in oxygen the rate of reduction of Mn^{III} was shown to be independent both of the rate of stirring and of the rate of flow of oxygen through the solution. All iodometric titrations were performed as rapidly as possible so as to minimise errors due to dissolution of oxygen. Colorimetric measurements of Mn^{III} were made by the standard method with a " Spekker " photo-electric absorptiometer with a green (No. 604) filter. Rates of absorption of oxygen were measured, by the standard procedure (Milton and Waters " Quantitative Micro-analysis," Ed. Arnold & Co., London, 1949), in Warburg micro-respirometers, shaken at 35°, at 130 oscillations per minute. Filter paper saturated with 20% potassium hydroxide solution was placed in the central cups of the reaction vessels to absorb the carbon dioxide evolved during the reaction. The buffered manganic pyrophosphate solution was placed in the main reaction vessel, and a known excess of malonic acid solution in the side-arm. Every determination was repeated at least 3 times.

Induced Polymerisation.—The polyvinyl cyanide formed during the anærobic oxidation of malonic acid in presence of vinyl cyanide was washed 12 times with distilled water at the centrifuge and was then dried at 100°. The infra-red spectrum showed, besides the very strong CN absorption band at 4·5 μ, distinct absorptions at 3·01 μ (OH group) and at 5·81 and 5·96 μ (C=O) which confirm the presence of carboxyl groups at the ends of the polymer chains.

Identification of Oxidation Products.—Formic acid, from the anærobic oxidations of malonic, tartronic, and glyoxylic acids was separated from the dilute aqueous reaction mixtures by passing a slow stream of air through warmed solutions and freezing their vapours in carbon dioxide–ethanol traps. The collected liquids gave chromotropic acid reactions for formic acid (Feigl, " Spot Tests," Elsevier, Amsterdam, 1949, p. 397).

Induced Oxidations.—The anærobic oxidation of malonic acid by manganic pyrophosphate in the presence of manganous ions was accelerated by the compounds listed in Table 1 (p. 2458). The following factors indicate the extent of acceleration of the initial reaction at pH 1·08 of mixtures of 100 ml. of 0·0243N-manganic pyrophosphate, 10 ml. of M-malonic acid, and 5 ml. of solutions containing 0·025 moles of added substrate : formic acid 1·5, diethyl ether 2·1, vinyl cyanide 3·3, ethanol 3·7, *iso*propanol 4·9, maleic and fumaric acid 26 times. Volatile oxidation products were separated by a stream of air, as described above; aldehydes were characterised by the m. p. and mixed m. p. of their dimedone derivatives, and other products were identified by diagnostic spot tests.