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Mechanisms of Organic Oxidation and Reduction by Metal Complexes

Electron and ligand transfer processes form the basis for redox reactions of radicals and metal species.

Jay K. Kochi

The mechanisms of most organic reactions can be classified into two categories, according to the mode in which the covalent bond undergoes fission: heterolytic (ionic) and homolytic (free-radical). For reactions proceeding by way of an ionic path, the transition state of the reaction takes on polar character, and in some cases carbonium ions and carbanions are actually intermediates. The general concept of acids and bases based on the electron pair, proposed by G. N. Lewis (*1*), in large part forms the framework for discussions of the mechanisms of ionic reactions.

Homolytic reactions, on the other hand, result from symmetric cleavage of chemical bonds and involve free radicals as intermediates. Combination and disproportionation of radicals as well as addition and atom transfer are characteristic reactions of radicals. In contrast to ionic reactions, both solvent and polar effects are usually small in homolytic reactions.

This mechanistic dichotomy into ionic and radical mechanisms is preserved intact in a large number of organic reactions (*1*). Yet such a categorization into mechanistic types has not prevailed in the study of inorganic mechanisms. This is, in part, due to the prevalence of charged inorganic species, which would make such a classification meaningless. There has been, however, an attempt to systematize inorganic oxidation and reduction reactions into one-equivalent and two-equivalent changes which have formal analogies to the radical and ionic categories.

Organic chemistry has relied heavily on inorganic chemistry to provide numerous reagents to effect oxidation and reduction and to promote catalysis. Cross-fertilization of ideas between these two disciplines, moreover, has markedly increased in recent years. The study of the mechanism of oxidation and reduction of organic compounds by inorganic reagents provides an important example of this interrelationship, and it seems to focus attention on the growing inadequacy of the gross categorization of reaction mechanisms. For example, if the concept of one-equivalent changes between inorganic species is applied to organic intermediates, the strict heterolytic-homolytic dichotomy is vitiated. The interconversion of the series of species carbonium ion (R^+), free radical ($R\cdot$), and carbanion ($R:^-$)

$$R^+ \overset{\pm e}{\rightleftharpoons} R\cdot \overset{\pm e}{\rightleftharpoons} R:^-$$

results from one-equivalent changes. In this manner free radicals can be considered intermediates in ionic reactions, and vice versa.

Such a broadening of the base of organic reaction mechanisms is most likely to result from studies of oxidation and reduction reactions. Inorganic oxidants and reductants are particularly useful in these investigations because of the variety of elemental species available with a multiplicity of oxidation states. Chromium (VI) and manganese (VII), as chromate and permanganate, are usually included in the repertory of readily available and useful oxidants (*2*). Oxidation of organic compounds with these reagents has been well examined and found generally to involve overall three-equivalent reduction of the oxidants to Cr^{III} and Mn^{IV} species, respectively. Mechanistic studies have shown that a number of intermediate metastable oxidation states such as Cr^{V} and Cr^{IV} and Mn^{VI} and Mn^{V} must be included in the reaction sequence. The manner in which each of these species in turn reacts with the organic intermediates is still incompletely understood.

An alternative initial approach to delineating the mechanisms of these complex reactions is to deal directly with the usual organic intermediate—the free radical. In such a case, the oxidation and reduction process is constrained to a one-equivalent change, and a more manageable number of species (organic and inorganic) is involved. In the following discussion, oxidation and reduction mechanisms involving organic free radicals and metal complexes are presented, with Cu^{II}, Cr^{II}, and Pb^{IV} used as illustrative examples.

The products, when metal salts and complexes are used as oxidants of free radicals, are highly dependent on the nature of the anion or ligand associated with the metal moiety. For example, cupric chloride and bromide oxidize alkyl radicals to the corresponding alkyl halides, but cupric sulfate or acetate under the same conditions pro-

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duces predominantly alkenes. Both reactions are qualitatively much faster than other usual reactions of free radicals, such as atom transfer and addition. Thus, in practice, such metal complexes inhibit some free-radical chain processes, such as polymerization, and greatly affect others, such as autoxidation.

As prototypes, cupric sulfate and cupric chloride represent two characteristic oxidants of free radicals by metal complexes (*3*). The oxidation of free radicals by metal salts such as cupric sulfate, perchlorate, or carboxylates to yield elimination products (alkenes) is described as an electron transfer process. The oxidation of free radicals by metal complexes such as cupric chloride to yield substitution products derived from the metal oxidant is described as a ligand transfer process. These definitions are analogous to outer-sphere and inner-sphere mechanisms which have been developed by inorganic chemists (*4*). In the discussion which follows, these two mechanisms are described in greater detail.

Electron Transfer Oxidation of Alkyl Radicals by Copper (II)

In aqueous solutions ethyl radicals are oxidized by Cu^{II} sulfate to ethylene, and Cu^{I} is formed according to the reaction

$$CH_3CH_2\cdot + Cu^{II} \rightarrow CH_2 = CH_2 + Cu^{I} + H^{+} \qquad (1)$$

The oxidation also proceeds readily in various nonaqueous solvents, such as benzene, acetonitrile, chloroform, and acetic acid, with carboxylato Cu^{II} oxidants. A number of other alkyl radicals (R·) are similarly oxidized, in excellent yields, to alkenes. However, as one progresses from primary to secondary to tertiary alkyl radicals, the high yields of alkene are supplanted by increasing amounts of substitution product RX, in which X represents other nucleophilic species present in solution (*5*). Thus, *t*-butyl radicals are oxidized by Cu^{II} in aqueous solution to a mixture of isobutylene and *t*-butyl alcohol, while, in methanol, isobutylene and *t*-butyl methyl ether are the products. *t*-Butyl acetate is formed along with isobutylene in acetic acid solutions. The amount of substitution product relative to alkene depends on the radical and, to a lesser extent, on the solvent. Its formation is promoted by protic solvents, and the yield qualitatively parallels the stability of the corresponding carbonium ion, as illustrated in Fig. 1.

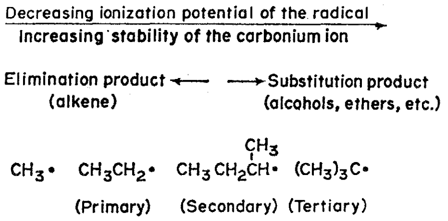
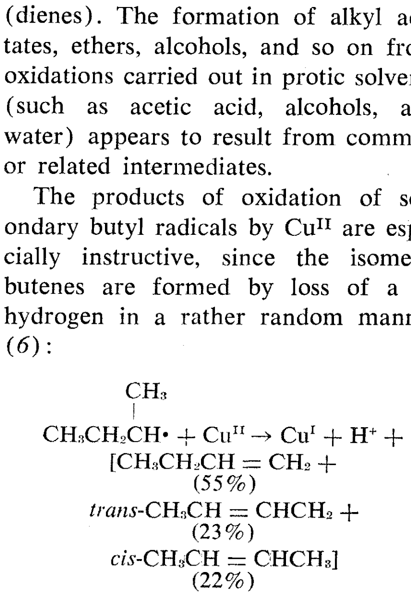


Fig. 1. Variation of elimination versus substitution in the oxidation of alkyl radicals by Cu^{II} acetate.

The partitioning of free radicals between elimination and substitution products of oxidation correlates roughly with the ionization potentials (Table 1) of the free alkyl radical in the gas phase. The amount of substitution product relative to elimination product produced from the oxidation of alkyl radicals by cupric sulfate increases as the ionization potential of the radical decreases. Or, as a corollary, the relationship can be stated as follows: the more stable the radical or the incipient carbonium ion formed by oxidation, the greater the probability that the products will be those expected from a carbonium ion formed by a classical solvolytic route.

Thus, while primary *n*-butyl radicals are oxidized by cupric acetate in glacial acetic acid to butene-1, exclusively, *t*-butyl radicals under the same conditions yield isobutylene (80 percent) as well as *t*-butyl acetate (20 percent). Further, the benzylic 1-phenylpropyl radical is oxidized largely to the substitution product 1-phenylpropyl acetate (80 percent), and the elimination product (*cis*- and *trans*-propenylbenzene) is minor (20 percent). Finally, allylic radicals such as the 2-butenyl radical yield, under these conditions, only acetates and no elimination product (dienes). The formation of alkyl acetates, ethers, alcohols, and so on from oxidations carried out in protic solvents (such as acetic acid, alcohols, and water) appears to result from common or related intermediates.

The products of oxidation of secondary butyl radicals by Cu^{II} are especially instructive, since the isomeric butenes are formed by loss of a β-hydrogen in a rather random manner (*6*):

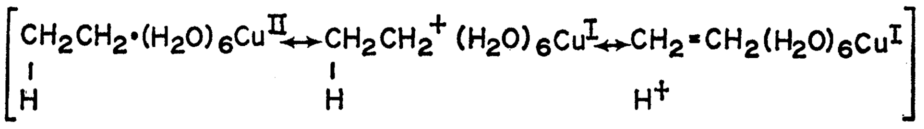


$$\begin{aligned} &CH_3CH_2\overset{\displaystyle CH_3 \atop |}{C}H\cdot + Cu^{II} \rightarrow Cu^{I} + H^{+} + \\ &\quad [\underset{(55\%)}{CH_3CH_2CH = CH_2} + \\ &\quad \underset{(23\%)}{trans\text{-}CH_3CH = CHCH_3} + \\ &\quad \underset{(22\%)}{cis\text{-}CH_3CH = CHCH_3}] \end{aligned} \qquad (2)$$

A number of investigators have commented on elimination reactions from *s*-butyl derivatives by various methods, including deamination of *s*-butylamine, solvolysis of *s*-butyl tosylate, Ag^{I}-assisted acetolysis of *s*-butyl halide, pyrolysis of *s*-butyl acetate, bimolecular base-induced elimination of *s*-butyl halides and tosylate, disproportionation of *s*-butyl radical, rearrangement of methyl ethyl carbene, and other methods, in attempts to determine electronic and conformational effects in elimination reactions. With the exception of the oxidation of *s*-butyl radical by Cu^{II}, none of these methods yields even an approximation to a statistical distribution of butene isomers [1-butene (60 percent), *trans*-butene-2 (20 percent), and *cis*-butene-2 (20 percent)]. The thermodynamic distribution at 30°C [butene-1 (2 percent), *trans*-butene-2 (75 percent), *cis*-butene-2 (23 percent)] varies widely from both our experimental and the statistical distribution and indicates that the stabilities of the oxidation products (butenes) do not control the direction of oxidative elimination of the β-hydrogen from an *s*-butyl radical by Cu^{II}. Thus we conclude that, in the transition state of the oxidation, the scission of the β-hydrogen bond has not progressed significantly, and that the primary driving force is derived by electron transfer.

Table 1. Estimated rate constants for the oxidation of alkyl radicals by Cu^{II} at 57°C.

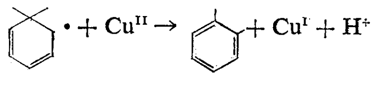
Alkyl radical	E (kcal/mole)	Log A	Log k_h (57°C)	k_h/k_e	Log k_e	Ionization potential (electron volts)
$CH_3CH_2\cdot$	5.9	8.1	4.2			8.80
$CH_3CH_2CH_2\cdot$	6.7	8.3	3.90	1.8×10^{-4}	7.64	8.69
CH_3CHCH_3	6.3	8.3	4.15	2.8×10^{-4}	7.70	7.90
$CH_3CH_2CH_2CH_2\cdot$	5.4	7.9	4.35	2.0×10^{-4}	8.05	8.64
$(CH_3)_2CHCH_2\cdot$	6.5	8.7	4.42	5.3×10^{-4}	7.70	8.35
$CH_3CH_2CHCH_3$	4.9	7.7	4.47	3.9×10^{-4}	7.88	7.93
$(CH_3)_3C\cdot$	4.3	7.5	4.70	9.2×10^{-5}	8.74	7.43



$$\left[\underset{\displaystyle H}{CH_2}CH_2\cdot(H_2O)_6Cu^{II} \leftrightarrow \underset{\displaystyle H}{CH_2}CH_2^{+}(H_2O)_6Cu^{I} \leftrightarrow \underset{\displaystyle H^{+}}{CH_2}{=}CH_2(H_2O)_6Cu^{I}\right]$$

Fig. 2. A transition state for oxidative elimination.

The importance of carbonium ions or carbonium-ion character in the transition state of the oxidation can further be ascertained from rearrangement studies. Neopentyl radical having no β-hydrogen does not itself rearrange, but on oxidation with Cu^{II} in glacial acetic acid it yields a mixture of rearranged elimination products, 2-methyl butene-1 and butene-2, as well as *t*-amyl acetate, diagnostic of cationic intermediates. Similarly Walling (*7*) has observed the following rearrangement:



(3)

Each of the homoallylic C_4H_7 radicals (cyclopropylcarbinyl, cyclobutyl, and allylcarbinyl) is oxidized by Cu^{II} in acetic acid to mixtures of rearranged C_4H_7 acetates characteristic of the homoallylic C_4H_7 cations.

The rate of oxidation of alkyl radicals by Cu^{II} can be determined relative to a hydrogen atom transfer reaction from various active hydrogen donors, especially aldehydes (*8*). Both reactions are second-order, and the competitive kinetic method allows determination of the ratio of second-order rate constants, k_e/k_h, for oxidation relative to hydrogen transfer. To obtain the absolute rates of these reactions it is necessary to know k_h independently. These hydrogen-transfer rates are known only for reactions in the gas phase. For example, the reduction of *n*-butyl radical by *n*-valeraldehyde

$$CH_3(CH_2)_2CH_2\cdot + CH_3(CH_2)_3CHO \xrightarrow{k_h} CH_3(CH_2)_2CH_3 + CH_3(CH_2)_3CO\cdot \quad (4)$$

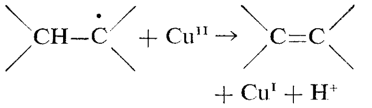
has an extrapolated rate constant of 2.2 × 10^4 liters per mole per second at 57°C (*9*). If it is *assumed* that this value is unchanged in solution, the second-order rate constant for the oxidation of *n*-butyl radical by Cu^{II},

$$CH_3CH_2CH_2CH_2\cdot + Cu^{II} \xrightarrow{k_e} CH_3CH_2CH{=}CH_2 + Cu^{I} + H^{+} \quad (5)$$

in aqueous acetic acid can be evaluated as 1 × 10^8 liters per mole per second. Further, the rates of oxidation of a series of other alkyl radicals can be calculated on such a simplifying basis, since the corresponding values of k_h have been determined. The rates of oxidation of several alkyl radicals calculated in this manner are listed in Table 1, together with hydrogen-transfer constants at 57°C.

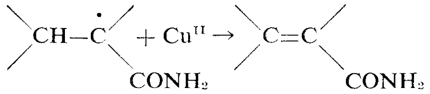
The high rates of the oxidation reaction are derived mainly from the low energy of activation. The temperature variation of k_h/k_e yielded a difference in activation energy between hydrogen transfer and oxidation, $E_h - E_e$, of 2.0 kilocalories per mole. The activation parameters for the transfer reaction 4 in the gas phase are: $E_4 = 5.4$ kilocalories per mole; $\log A_4 = 7.9$. The application of these values to solution yields values of $E_5 = 3.4 \pm 0.5$ kilocalories per mole and $\log A_5 = 10.1 \pm 0.5$ for the oxidation reaction 5. Thus, the second-order rate constants for the oxidation of alkyl radicals by Cu^{II} are close to the diffusion-controlled limit. These large rate constants reflect the versatility of the Cu^{II} oxidation reaction in precluding the usual second-order reactions of free alkyl radicals. The usefulness of Cu^{II} salts as trapping agents for free alkyl radicals and in synthetic applications is apparent.

From the data of Table 1 one can see that the second-order rate constants for the oxidation of alkyl radicals by Cu^{II}



(6)

is generally in the range 10^8 liters per mole per second. These values compare with the second-order rate constant of 10^3 liters per mole per second determined by Dainton and his co-workers (*10*) for the oxidation of an α-carbamido- radical by Cu^{II} in aqueous solution.



(7)

It is thought that the difference in rate between these analogous reactions is due to the presence of an electron-withdrawing group ($CONH_2$) in the α-position of the latter radicals. Earlier qualitative studies have shown that electron-withdrawing groups indeed hinder the oxidation of radicals by electron transfer. These studies were based on product isolation, which is, at best, a very insensitive tool. The deleterious effect of such substituents has also been observed in the enhanced ionization potentials of the substituted radicals. Thus, the ionization potential (in the gas phase) of alkyl radicals progressively increases with substitution of cyano groups. The phenomenon is undoubtedly related to the destabilization of the cationic transition state or intermediate through substitution with electron-withdrawing groups.

Formally, there are several mechanisms which can be postulated for the oxidation of free radicals by Cu^{II}. In a two-step process the oxidation is postulated to proceed by way of a carbonium ion intermediate which results from the transfer of the electron from the radical to the Cu^{II} species. In a subsequent step the alkyl carbonium ion is partitioned between alkene, by proton loss, or substitution product, by solvation.

$$CH_3CH_2\cdot + Cu^{II} \rightarrow CH_3CH_2^{+} + Cu^{I} \quad (8)$$

$$CH_3CH_2^{+} \longrightarrow \begin{cases} \longrightarrow CH_2{=}CH_2 + H^{+} & (9) \\ \xrightarrow{HX} CH_3CH_2X + H^{+} & (10) \end{cases}$$

An alternative mechanism to that which postulates a discrete carbonium ion as an intermediate is one in which the reaction between the free radical and Cu^{II} occurs by way of two concurrent reactions proceeding through *similar* transition states. In aqueous solutions the transition state for elimination may be depicted as in Fig. 2. A similar transition state can be constructed for the oxidative substitution reaction.

The choice between the discrete carbonium ion mechanism and the synchronous one-step processes is at present a difficult and, to a certain extent, a moot one. In essence these paths differ by the extent to which the β-hydrogen bond is broken in the rate-limiting step. In the former case this bond is largely intact throughout the oxidation step, while in the latter case the breaking of the β-hydrogen bond and the electron transfer from the radical to the metal ion are concerted.

There is evidence that both of these mechanisms are operative to a degree in reactions involving individual free radicals. The carbonium ion mechanism is most satisfactory with certain free radicals which produce relatively stable carbonium ions. The oxidation of tertiary alkyl radicals, benzylic radicals, and allylic radicals to yield mainly substitution products, the rearrangement of neopentyl- and homoallylic-type radicals on oxidation, and the low selectivity observed in oxidative elimination of *s*-butyl radicals are in harmony with the postulation of carbonium ion intermediates. Similarly, there is no evidence for ketene intermediates in the oxidation of acyl radicals. The *t*-butyl radicals formed from several sources show a pronounced propensity for forming, primarily, products which correspond to the solvolytic reaction of the *t*-butyl cation, together with some elimination product, isobutylene. This rationale, however, is difficult to apply to primary alkyl systems like ethyl radical. The high yield of elimination product is not consistent with the postulation of an intermediate carbonium ion, at least not one of the type usually found in organic reactions.

The oxidation of primary radicals to afford unsaturated products in high yields has no analogy in classical carbonium ion reactions. In the classical reactions, substitution rather than elimination is the usual course of reaction. The presence of ethylene as a major product from the oxidation of ethyl radical can, on this basis, be taken as evidence against the formation of ethyl cation in solution. However, with the exception of the recently examined alkyl carbonium ions in extremely acidic media (*11*), there exists no good model system for the study of reactions of free primary carbonium ions in solution, and extrapolation from classical methods of carbonium ion generation is tenuous. Energetically, a reaction such as reaction 8 is found to be highly endothermic if the large ionization potential of primary alkyl radicals is considered (*12*), and reasonable estimates of the solvation energy of primary carbonium ions are used in the calculation. The thermodynamic value of the Cu^{II}—Cu^{I} couple is completely inadequate for explaining the high energy of ionization of primary radicals in solution. The high rates and low activation energies are incompatible with such an energetically unfavorable process as the simple electron transfer oxidation of a primary alkyl radical presented in reaction 8.

Regardless of whether the oxidation of free radicals by cupric salts proceeds by a one-step concerted mechanism or a two-step carbonium ion mechanism, it is clear from the reactivities of free radicals that a sizable amount of positive charge is incorporated by the organic moiety in the transition states. Thus, oxy and thiyl radicals are unaffected by cupric salts. Even radicals such as *n*-butoxy and cyclohexyloxy, which possess β-hydrogens, appear not to be oxidized by cupric salts. The inability of cupric salts to oxidize even alkoxy radicals, which possess removable β-hydrogens, indicates that a considerable driving force for the oxidation is derived from the removal of the electron from the free radical center. The ionization potential of oxygen (13.64 volts) is greater than that of carbon (11.26 volts), and the transfer of an electron from an oxy radical would be expected to be more difficult than that from an analogous carbon center. Morever, the selectivity is not restricted to oxy and thiyl radicals, since, as other experiments have shown, negatively substituted carbon radicals show different reactivities toward various cupric salts. Thus the determining factor in an oxidation of a free radical is the ease with which the system can achieve cationic character in the transition state, and the contribution to the driving force from the loss of a β-proton cannot be very large. The contribution of electron transfer in the transition state of a one-step mechanism for oxidation is included in Fig. 2.

The rapidity of the oxidation of alkyl radicals by Cu^{II} is similar to some electron-transfer reactions between inorganic species. The kinetics of fast electron-transfer reactions have been recently measured by relaxation methods. Electron transfer between cationic phenanthroline Fe^{II} complexes and anionic Ir^{IV} or Ru^{IV} complexes have second-order rate constants of the order of 10^9 liters per mole per second, close to the diffusion-controlled limit (*13*). These "outer-sphere" electron-transfer reactions in many respects represent optimum examples, according to the theory developed by Marcus (*14*). The second-order rate constant for an electron-transfer process, according to Marcus, is given by the expression

$$k = Ze^{[-W(R)\ +\ m^2\lambda]/kT} \tag{11}$$

where Z is the collision frequency between two uncharged reactants in solution, $W(R)$ is the coulombic work involved in bringing the reactants together in the transition state, and $m^2\lambda$ is a term related to the work necessary to reorganize the coordination shell around the reactant ion. The reaction between alkyl free radicals and metal ion oxidants should also represent optimum examples of facile electron-transfer reactions. Qualitatively, electrostatic repulsion between reactants is absent, owing to the neutrality of the free radical. Only the reorganization of the coordination shell around the metal oxidant is important, since free radicals, by and large, involve minimal changes in solvation in attaining the transition state. The reorganization energy associated with the Cu^{I}—Cu^{II} transformation appears not to be large. For example, the less favorable "inner-sphere" exchange between Cu^{I} chloride and Cu^{II} chloride has an estimated second-order rate constant of 5×10^5 liters per mole per second (*15*). The low free energy of activation of Cu^{II} oxidations of alkyl radicals resembles, in some respects, the facile coupling of these radicals in solution.

Oxidation of Free Radicals by Ligand Transfer

If ethyl radicals are oxidized by cupric chloride in aqueous solutions, ethyl chloride is produced, according to the reaction

$$\begin{aligned} CH_3CH_2\cdot + Cu^{II}Cl_2 \rightarrow \\ CH_3CH_2Cl + Cu^{I}Cl \end{aligned} \tag{12}$$

No ethyl alcohol is formed, and the amount of ethylene, which is a side product, increases at only low concentrations of cupric chloride. Similarly, the oxidation in glacial acetic acid produces only ethyl chloride and no acetate. The amount of ethyl chloride relative to ethylene is dependent on the nature of the Cu^{II} species. In aqueous solutions cupric chloride is dissociated according to the following reactions (*16*):

$$\begin{aligned} Cu^{++} + Cl^- \rightleftharpoons CuCl^+ \ K_{CuCl^+} = 1.0 \\ CuCl^+ + Cl^- \rightleftharpoons CuCl_2 \ K_{CuCl_2} = 0.2 \end{aligned} \tag{13}$$

Both $CuCl_2$ and $CuCl^+$ are capable of oxidizing alkyl radicals to alkyl chloride, whereas $Cu^{II}{}_{aq}$ produces ethylene. Addition of chloride salts enhances alkyl chloride formation by the mass-action effect of chloride ion on the equilibria shown in reactions 13. Chlorocuprates ($CuCl_n{}^{-n+2}$, $n > 2$) are significant in aqueous solutions only at high concentrations, due to their low formation constants [$K_{CuCl_3{}^-} = 0.04$, $K_{CuCl_4{}^{--}} = 0.01$]. However, Cu^{II} and chloride ions are more weakly solvated in such nonaqueous solvents as acetonitrile and acetic acid, and, as a result, the formation constants of the chlorocupric complexes are much higher than they are in aqueous solutions (*17*). For example, the formation constants in acetonitrile are as follows:

$K_{\mathrm{CuCl}^+} = 10^{9.7}$, $K_{\mathrm{CuCl}_2} = 10^{7.9}$,
$K_{\mathrm{CuCl}_3^-} = 10^{7.1}$, and $K_{\mathrm{CuCl}_4^{--}} = 10^{3.7}$.

Consequently, oxidation of alkyl radicals in nonaqueous solvents such as acetic acid and acetonitrile produces only alkyl chlorides and very little alkene, even at low concentrations of oxidant.

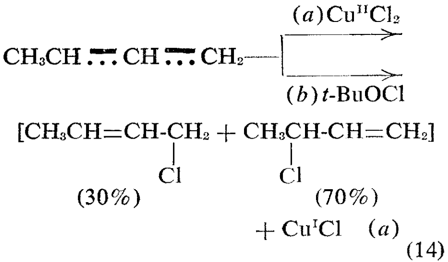
A large number of different chlorocupric species exist in solution, and determination of individual rate constants for oxidation by each of these species is difficult. On a molar basis, cupric chloride is more effective than cupric acetate as an oxidant for most alkyl radicals. This is especially true in nonaqueous solutions in which cupric acetate exists largely as an inactive dimer ($K_{\mathrm{dimer}} = 10^4$ in acetic acid) (*18*). Consequently alkyl chlorides usually predominate over alkenes when mixtures of acetate and chloride are used as ligands for the $\mathrm{Cu^{II}}$ oxidant. In glacial acteic acid the averaged rate constant for all chlorocupric species involved in the ligand transfer oxidation of *n*-butyl radical is about 10^5 to 10^6 liters per mole per second.

Those alkyl radicals which are likely to give products of rearrangement from electron transfer oxidation show no or little evidence of such a tendency when subjected to the action of a typical ligand transfer oxidant such as cupric chloride. Thus, neopentyl radicals are oxidized exclusively to neopentyl chloride. Cyclobutyl chloride and allylcarbinyl chloride are the predominant products from the oxidation of cyclobutyl and allylcarbinyl radicals, respectively.

Though oxidation of alkyl radicals by $\mathrm{Cu^{II}}$ oxidants proceeds readily by both electron transfer and ligand transfer processes, the transition state of each process has unique characteristics. Carbonium-ion character dominates electron transfer oxidation and leads, predominantly, to production of alkenes from alkyl radicals, production of unique substitution products from allylic and benzylic radicals, and rearrangement from neopentyl and homoallylic radicals. Ligand transfer oxidation, on the other hand, gives the same class of products irrespective of the nature of the free radical; organic chlorides are the major products from $\mathrm{Cu^{II}}$ chloride, bromides from $\mathrm{Fe^{III}}$ bromide, thiocyanates from $\mathrm{Fe^{III}}$ thiocyanates. In formal terms, the ligands are transferred from the metal complex to the free radical as an atom or free radical, and an overall one-equivalent change results. The transition state, shown in Fig. 3, for the transfer of chlorine from cupric chloride to an ethyl radical involves significantly less carbonium-ion character than the transition state for electron transfer oxidation. For chlorine ligand transfer, the transition state is more akin to that of an atom transfer in a typical free radical reaction. Thus, the ambident butenyl radical gives the same ratio of α-methallyl chloride and crotyl chloride in its reaction with the ligand transfer oxidant cupric chloride that it gives in its reaction with a chlorine atom transfer agent, *t*-butyl hypochlorite (*19*).

$$\left[\mathrm{CH_3CH_2\cdot Cl{-}Cu^{II}Cl} \longleftrightarrow \mathrm{CH_3CH_2{-}ClCu^{I}Cl}\right]$$

Fig. 3. Transition state for ligand transfer oxidation.



$$\mathrm{CH_3CH \overset{\cdots}{\text{—}} CH \overset{\cdots}{\text{—}} CH_2} \begin{cases} \xrightarrow{(a)\mathrm{Cu^{II}Cl_2}} \\ \xrightarrow[(b)t\text{-}\mathrm{BuOCl}]{} \end{cases}$$
$$\left[\mathrm{CH_3CH{=}CH{-}\underset{\displaystyle |\atop \displaystyle Cl}{CH_2}} + \mathrm{CH_3\underset{\displaystyle |\atop \displaystyle Cl}{CH}{-}CH{=}CH_2}\right]$$
$$(30\%) \qquad (70\%)$$
$$\begin{aligned} &+ \mathrm{Cu^{I}Cl} \quad (a) \\ &+ t\text{-}\mathrm{BuO}\cdot \quad (b) \end{aligned} \qquad (14)$$

The charge distribution on the butenyl moiety in the transition states of the two reactions must be similar, and both must be quite distinct from an electron transfer transition state.

This analogy between an oxidative process involving ligand transfer with a metal complex and the usual free radical atom transfer process is not unique but further illustrates the unity between reactions which span the traditional gap between an inorganic species and an organic chain transfer agent. With the former, ligand transfer merely results in a one-equivalent reduction of the metal complex to its lower oxidation state. With organic atom transfer reagents, a new free radical (such as *t*-BuO·), is generated, to propagate a chain process. Potentially a variety the metal complex to its lower oxidation states can be generated through the judicious use of organic intermediates.

The distinction between electron transfer and ligand transfer routes is further amplified in a study of substituted alkyl radicals. α-Cyanoisopropyl radicals are effectively oxidized by cupric chloride to α-chloroisobutyronitrile in a typical ligand transfer oxidation (*20*). Cupric acetate under the same circumstances is ineffective as an oxidant. Similarly, alkyl radicals with other electron-withdrawing alpha substituents such as carboxy, acetyl, carbalkoxy, chloro, cyano, and nitro, are rapidly oxidized by cupric chloride to the corresponding chloro compound, while cupric acetate is largely ineffective in trapping these radicals. It is the electron transfer reagent cupric acetate which suffers when electron-withdrawing substituents are placed in the alpha position. Thus, the rate of oxidation of *n*-butyl radical by cupric acetate is diminished by a factor of 10^4 to 10^5 by the presence of an α-carbamido substituent. It is apparent that cupric and ferric halides are useful, whereas cupric acetate is not an effective inhibitor for most vinyl polymerizations, since the latter usually involve monomers with electron-withdrawing functional groups.

The discrimination which negatively substituted alkyl radicals display toward ligand transfer and electron transfer oxidants provides further support for the formulation of the cationic transition state A in electron transfer processes.

$$\mathrm{R\cdot + Cu^{II}Y \rightarrow \underset{\mathbf{A}}{[R\cdot YCu^{II} \longleftrightarrow R^{+}YCu^{I}]} \rightarrow}$$
$$\mathrm{Cu^{I}} + \text{carbonium products} \qquad (15)$$

In a ligand transfer process the direct transfer of an atom or radical from the metal oxidant to the carbon radical moderates the development of charges in the transition state, such as state B.

$$\mathrm{R\cdot + Cu^{II}X \rightarrow \underset{\mathbf{B}}{[R\cdot XCu^{II} \longleftrightarrow R{-}XCu^{I}]} \rightarrow}$$
$$\mathrm{R{-}X + Cu^{I}} \qquad (16)$$

The inability of the carbon moiety to accommodate a positive charge is thus a less severe limitation in ligand transfer processes. The formation of carbonium ions is not favored by electron-withdrawing substituents in the α-position, and this effect is manifested by the fact that oxidation of such radicals by electron transfer oxidants is more difficult than their oxidation by ligand transfer oxidants.

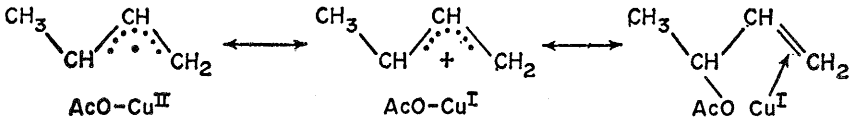


Fig. 4. Transition state for oxidation of allylic radicals by $Cu^{II}OAc_2$.

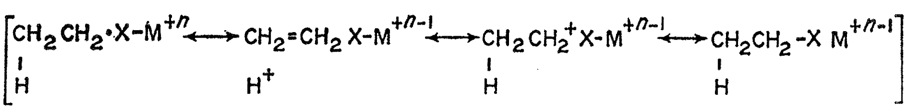
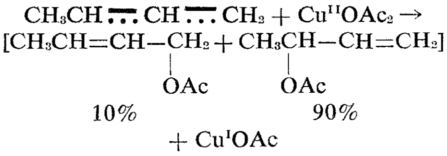


Fig. 5. Generalized transition state for oxidation of free radicals by metal complexes.

The oxidation of allylic radicals represents a unique bridge between distinct electron transfer and ligand transfer mechanisms (*21*). The oxidation of butenyl radicals by Cu^{II} acetate



$$CH_3CH \overset{\cdots}{-} CH \overset{\cdots}{-} CH_2 + Cu^{II}OAc_2 \rightarrow [\underset{10\%}{CH_3CH{=}CH{-}\underset{OAc}{CH_2}} + \underset{90\%}{CH_3\underset{OAc}{CH}{-}CH{=}CH_2}] + Cu^{I}OAc$$

in a variety of solvents leads to a mixture of allylic acetates which is quite distinct from the distribution of allylic chlorides generated by ligand transfer, discussed above. The predominance of the α-methallyl acetate isomer is unusual, since it is the thermodynamically less stable isomer. Further, the composition of this mixture differs from that (approximately 50 percent for each isomer) obtained by Ag^I-induced acetolysis of either isomeric butenyl chloride. The latter method appears, heretofore, to be the best available kinetic route to butenyl cations.

The composition of the butenyl acetate mixture derived from the oxidation of butenyl radicals can be systematically altered to the composition obtained from acetolysis of the butenyl chlorides simply by altering the Cu^{II} oxidant. Formation of phenanthroline, bipyridyl, pyridine, and acetonitrile complexes of Cu^{II} facilitates the electron transfer process, and butenyl cations are intermediates much like those derived from the solvolysis reaction. On the other hand, the high specificity for the α-methallyl isomer obtained from the oxidation of butenyl radical by uncomplexed cupric acetate is postulated to proceed via a transition state (Fig. 4) which has both electron transfer and ligand transfer character, with the former predominating. Part of the driving force for preference of the α-methallyl isomer relative to the crotyl isomer is derived through interaction of Cu^I with the terminal double bond.

Various ligands, usually halides and pseudohalides, have been transferred to free radicals from Cu^{II} complexes. These include chloride, bromide, thiocyanate, cyanide, and azide. The compounds Fe^{III}, Cr^{VI}, Pt^{IV}, and Pb^{IV}, among the more common metal species in higher oxidation states, have been used in preliminary studies. Potentially the available metal species and associated ligands for study are inexhaustible. No doubt more extensive studies will provide insight and subtle variations on these reactions.

Finally, the classification of these oxidative processes of free radicals into discrete categories is arbitrary. The designation of the mechanism as a two-step carbonium ion, a one-step oxidative elimination, or a one-step ligand transfer reaction may be too arbitrary. It is conceivable that all these reactions proceed through related transition states in which the resonance contributions from electron transfer and ligand transfer, as indicated in Fig. 5, vary not only with the structure of the free radical but also with the identity of the metal species and counter-ions.

The concept of ligand transfer presented in these studies is not unique to organic free radicals. It is derived in large part from the classic studies of Taube and his co-workers, who were the first to delineate the mechanism of inorganic oxidation and reduction reactions as outer-sphere (electron transfer) and inner-sphere (ligand transfer) processes (*22*). They postulated that the transition state of the latter involves a ligand-bridged activated complex. In the example

$$\begin{aligned}&(NH_3)_5Co^{III}Cl^* + Cr^{II} \rightarrow \\ &\quad \underset{\text{Bridged activated complex}}{[(NH_3)_5Co\cdot\cdot Cl\cdot\cdot Cr]^{++}} \rightarrow \\ &5NH_3 + Co^{II} + Cr^{III}Cl^* \qquad (17)\end{aligned}$$

the identity of the chlorine in both the substitution-stable oxidant and the product was demonstrated by the use of isotopic tracers. The inner-sphere mechanism has been established for a variety of wholly inorganic oxidation-reduction reactions, but is difficult to confirm in substitution-labile systems. For such systems, a study of free radicals in conjunction with metal complexes may aid in the general solution of this problem, because of the variety of structural and electronic parameters as well as stabilities available in organic systems. The following discussion of Cr^{II} as a reducing agent and Pb^{IV} as an oxidizing reagent represents such an attempt.

Reduction of Organic Halides by Chromium (II)

Chromium (II) compounds are convenient reagents for use as reducing agents because they are readily converted to Cr^{III}. Benzyl halides are reduced by Cr^{II} salts in a variety of solvents, according to reaction 18:

$$PhCH_2X + 2Cr^{II} \rightarrow PhCH_2Cr^{++} + CrX^{++} \qquad (18)$$

The reaction is first-order in each reactant, and the rate of formation of benzylchromium cation is the same as the rate of disappearance of benzyl halide (*23*). Since the stoichiometry of the activation process is different from the stoichiometry of the net reaction (reaction 18), the reaction cannot occur by a single-step mechanism. Furthermore, the same benzylchromium species is formed quantitatively from Cr^{II} and phenyl-*t*-butyl hydroperoxide

$$\begin{aligned}&PhCH_2C(CH_3)_2OOH + 2Cr^{II} \rightarrow \\ &\quad PhCH_2Cr^{++} + (CH_3)_2CO + Cr^{III}, \qquad (19)\end{aligned}$$

and from *bis*-phenylacetyl peroxide. We formulate these reductions as proceeding by way of a common mechanism. The two-step sequence, for example, for the reduction of benzyl chloride is:

$$PhCH_2Cl + Cr^{II} \xrightarrow{\text{slow}} PhCH_2\cdot + CrCl^{++} \qquad (20)$$

$$PhCH_2\cdot + Cr^{II} \xrightarrow{\text{fast}} PhCH_2Cr^{++} \qquad (21)$$

The initial rate-determining step in these mechanisms is the one-equivalent reduction of the organic substrate by Cr^{II} to generate Cr^{III} products and a free radical. Thus, an alkyl radical is generated from an organic halide, and an oxy radical is generated from a peroxide; both react in fast subsequent steps and can be detected as intermediates by means of trapping experiments. The formation of the substitution-stable chlorochromic ion $CrCl^{++}$ is further indication that chlorine is being transferred directly from benzyl chloride to the chromous ion in the rate-determining step (reaction 20), since such a chlorochromic complex is not formed from aquochromic ion and chloride ion under the conditions of these reactions.

Formally, the slow step (reaction 20) in these reactions is analogous to the atom-transfer reactions of chromous ion with transition-metal halide complexes. It is the *microscopic reverse* process for the oxidation of free radicals by the ligand transfer mechanism; for example,

$$CH_3\cdot + Cu^{II}Cl_2 \rightarrow CH_3Cl + Cu^{I}Cl \quad (22)$$

As such, the electronic requirements for the two processes should be the same. Thus, a measure of these electronic effects in atom-transfer reactions can be obtained by comparing the relative rates of reaction of various substituted benzyl halides. In Table 2, the relative rates of reduction are compared with the Hammett σ-parameter. Also included are rates of pyrolysis of benzyl bromides in the gas phase obtained by Szwarc *et al.* (*24*), using the toluene carrier method.

$$ArCH_2\text{—}Br \overset{\Delta}{\rightarrow} ArCH_2\cdot + Br\cdot \quad (23)$$

It is apparent that there is no correlation between the relative rates of reduction and the Hammett linear free-energy relationship. A similar lack of correlation has been observed in the polarographic reduction of a variety of substituted benzyl halides (*25*). All substituents increase the rate of reduction relative to the unsubstituted compound, albeit in relatively small degree, and the effect is much the same as effects obtained by Szwarc from pyrolysis studies. This parallel behavior provides further support for the homolytic dissociation of the benzyl-halogen bond in the rate-controlling step.

The relatively small polar effect in the chlorine atom-transfer reaction and in the pyrolysis of the bromides is in marked contrast to the strong Hammett correlation and the pronounced polar effect in hydrogen atom-transfer reactions from substituted toluenes by atoms and free radicals (*26*). The implications of this, in a consideration of electron transfer in free-radical chain transfer reactions

$$ArCH_3 + X\cdot \rightarrow ArCH_2\cdot + HX \quad (24)$$

(where X = Cl, Br, *t*-BuO, Cl_3C) has been discussed elsewhere (*27*).

Table 2. Relative rates of reduction and pyrolysis of benzyl halides.

Relative rates of benzyl bromide reduction*	Relative rates of benzyl bromide pyrolysis†	Hammett σ-parameter	Substituent
1.00	1.00	0.000	H
1.04 (1.02)‡	1.02	− 0.069	*m*-Me
1.33 (1.6)‡		− 0.170	*p*-Me
1.07	1.13	+ 0.373	*m*-Cl
(1.26)‡	1.3	+ 0.227	*p*-Cl
1.17	1.24	+ 0.391	*m*-Br
1.43	1.24	+ 0.232	*p*-Br
1.04 (0.95)‡		+ 0.115	*m*-MeO

* Temperature, 27.5 ± 0.5°C; aqueous ethanol. † 800°K; Toluene carrier method. ‡ Corresponding value for benzyl chlorides are given in parentheses.

Table 3. Reduction by Cr^{II}.

Oxidant	k_2 (liter $mole^{-1}$ sec^{-1})	T (°C)	ΔH^* (kcal/mole)	ΔS^* (entropy unit)
$(NH_3)_5CrF^{++}$	2.7×10^{-4}	25	13.4	−30
$(NH_3)_5CrCl^{++}$	5.1×10^{-2}	25	11.1	−23
$(NH_3)_5CrBr^{++}$	3.2×10^{-1}	25	8.5	−33
$(NH_3)_5CrI^{++}$	5.5	25		
$(H_2O)_5CrF^{++}$	2.6×10^{-2}	27	13.7	−20
$(H_2O)_5CrCl^{++}$	5×10^{-2}	0		
$PhCH_2Cl$	3.2×10^{-3}	27	14.0	−14.3
$PhCH_2Br$	4.1×10^{-1}	27		
$PhCH_2I$	1.8			

The slow halogen atom-transfer step (reaction 20) is fastest with iodide and slowest with chloride, the relative rates for the iodide, bromide, and chloride being 555, 124, and 1, respectively. This order follows the bond dissociation energies of the benzyl derivatives benzyl iodide, benzyl bromide, and benzyl chloride—39, 51, and 60 kilocalories per mole, respectively—and chain-transfer reactivities of alkyl halides toward carbon radicals.

There is a similarity of reaction type between the reduction of benzyl halides by chromous ion and the reduction of halopentamminechromic complexes in aqueous perchloric acid, studied by Taube and Ogard (*28*). Some results are summarized in Table 3. As discussed above, the view that the latter reactions proceed by way of a halogen-bridged activated complex is generally accepted. For the reduction of benzyl halides, a transition state such as C is proposed.

$$[ArCH_2 \cdots Cl \cdots Cr]^{++}$$

C

The values of the activation ethalpies for the reactions of the pentamminechromic complexes are similar to those for the benzyl halide reactions. The entropies of activation, however, are slightly higher, presumably because the reaction of the pentamminechromic complex involves rearrangement in the coordination sphere of the oxidizing agent, leading to the loss of the ammonia ligands.

$$Cr(H_2O)_5F^{++} + {}^{*}Cr^{++} \rightarrow {}^{*}Cr(H_2O)_5F^{++}\ Cr^{++} \quad (25)$$

The reaction in which a ligand is exchanged between isotopes has been studied, and the activation parameters have been determined to be $\Delta H^* =$ 13.7 kilocalories per mole, $\Delta S^* = -20$ entropy units, at ionic strength $\mu =$ 1.0.

The smaller change in ΔS^* might be accounted for by the fact that the oxidant, benzyl chloride, has a lower requirement of bond stretching and ligand rearrangement than an inorganic aquo ion has. Although a study of entropies of activation is by no means conclusive evidence in the elucidation of a mechanism, it is instructive to note that the entropies of activation for the halogen abstraction reaction decrease in the order $(NH_3)_5CrX^{++} >$ $(H_2O)_5CrX^{++} > C_6H_5CH_2X$, in accordance with decreasing requirements (NH_3, H_2O, solvent) of ligand participation in the transition state.

The facile reduction of alkyl halides to alkanes by way of alkylchromium intermediates also has synthetic application to a variety of derivatives. Thus, vicinal dihalides are reductively eliminated by Cr^{II}, quantitatively, to alkenes (29) by two successive halogen atom transfers. Similarly, β-halohydrins and β-haloamines, together with their derivatives, produce alkenes in excellent yield, with the more potent ethylenediamine Cr^{II} as reducing agent. In the latter case, halogen atom transfer,

Table 4. Reduction of cations (M^{x+}) by the polyacrylamide radical ($m_j\cdot$) (see *10*). $E_n°$, Standard reduction potential, in electron volts; k_t, second-order rate constant; A_t, preexponential constant; E_t, activation energy.

M^{x+}	Redox potential* [$E_n°$ (V)]	k_t (liter mole^{-1} sec^{-1})†	A_t (liter mole^{-1} sec^{-1})	E_t (kcal mole^{-1})
Fe^{3+}	-0.771	2.8×10^3	1.45×10^5	2.35 ± 0.6
Cu^{++}	-0.153	1.17×10^3	1.1×10^7	5.4 ± 1.3
$CeOH^{3+}$	-1.7	3.2×10^3	2×10^5	2.45
Hg^{++}	(-0.9 for Hg_2^{++}/Hg^{++}); unknown for Hg^+/Hg^{++})	1.05	4.2×10^4	6.2 ± 1.0
Tl^{3+}	E° for Tl^{++}/Tl^{3+} unknown, but $-0.79 > E° > -1.7$	0.34	21	2.5 ± 0.4
Ag^+	-0.8	0		

* Acid media. † At 25°C. In 1.0N $HClO_4$.

followed by hydroxy or amino radical transfer to Cr^{II}, presumably prevails. 1,3-Dihalides generate cyclopropanes with this reagent, and it has been postulated that geminal dihalides generate carbenoid intermediates (*30*).

In addition to Cr^{II}, a number of other reducing metal complexes effect halogen atom transfer from organic halides. Recently Menapace and Kuivila (*31*) described the reduction of alkyl halides by organo-tin hydrides as a free-radical chain process. For this discussion, the relevant chain-transfer step which they postulated as

$$R_3Sn^{III}\cdot + ArCH_2{-}X \rightarrow R_3Sn^{IV}X + ArCH_2\cdot \quad (26)$$

is akin to reaction 20 in the reductions by chromous ions. Other studies indicate that silicon and lead in group IV also perform such functions. Furthermore, pentacyanocobaltate (II), many transition metal carbonyls, and Fe^{II} porphyrins reduce organic halides by a similiar ligand transfer process.

Finally, we feel that the concept of atom or ligand transfer has general validity for a variety of organic systems other than halides and peroxides. The application of metal complexes in particularly low and uncommon oxidation states to organic compounds with other functional groups is clearly desirable.

Oxidative Decarboxylation of Acids by Lead (IV)

Lead (IV) compounds have been considered classic examples of two-equivalent oxidants, since Pb^{II} derivatives invariably result. Recent kinetic studies of the oxidation of Co^{II} and Ce^{III} by Pb^{IV} show overall second-order kinetics (*32*). This suggests that Pb^{III} is an intermediate, and that the reduction of Pb^{IV} may consist of successive one-equivalent changes.

Pb^{IV} acetate is a useful reagent for the decarboxylation of aliphatic acids (RCO_2H) (*3*). This facile reaction

$$Pb^{IV}OAc_4 + RCO_2H \rightarrow Pb^{II}OAc_2 + ROAc + HOAc + CO_2 \quad (27)$$

has been shown to be a radical chain process which is initiated photochemically or thermally with peroxides and is strongly inhibited by oxygen or phenols (*33*). Alkyl radicals have been demonstrated to be *prime* intermediates in the decarboxylation, and the presence of Pb^{III} is also strongly implicated. Subsequent to fast metathesis, the pertinent propagation sequence in the chain decomposition is:

$$Pb^{III}(O_2CR)_3 \rightarrow Pb^{II}(O_2CR)_2 + R\cdot + CO_2 \quad (28)$$

$$R\cdot + Pb^{IV}(O_2CR)_4 \rightarrow [R^+]O_2CR + Pb^{III}(O_2CR)_3 \text{ etc.} \quad (29)$$

Reaction 29 represents the oxidation of the alkyl radical derived from the decarboxylation of the carboxylate moiety (reaction 28). It shows characteristics of an electron transfer oxidation in that alkyl acetates, in addition to alkenes, are formed in the presence of acetic acid. Furthermore, neopentyl radicals derived from *t*-butylacetic acid yield rearranged *t*-amyl acetate as well as the 2-methylbutenes. Cyclobutane carboxylic acid affords a mixture of homoallylic esters. Thus, Pb^{IV} carboxylates qualitatively resemble Cu^{II} carboxylates in their behavior toward alkyl radicals. However, primary alkyl radicals are not as effectively oxidized by Pb^{IV} esters as they are by Cu^{II} derivatives. Consequently, the decarboxylations of primary acids by Pb^{IV} are slow and exhibit short chain lengths, and they yield relatively large amounts of alkane by hydrogen transfer of the intermediate primary alkyl radicals with solvent.

Furthermore, Cu^{II} carboxylates markedly catalyze the decarboxylation of acids by Pb^{IV}. Even primary acids under these conditions afford alkenes in the excellent yields characteristic of the efficient trapping of primary radicals by Cu^{II}. We explain the Cu^{II} catalysis in terms of the incursion of two facile reactions (30 and 31) to replace the slower Pb^{IV} oxidation (reaction 29) in the propagation sequence

$$R\cdot + Cu^{II} \rightarrow [R^+] + Cu^I \quad (30)$$

$$Cu^I + Pb^{IV} \rightarrow Cu^{II} + Pb^{III} \text{ etc.} \quad (31)$$

At present it is difficult to explain why Cu^{II} is so much more efficient than Pb^{IV} in the oxidation of alkyl (especially primary) radicals. We estimate the oxidation potential for the Pb^{IV}-Pb^{III} couple to be at least 1.5 volts, as compared to less than 0.2 volt for the Cu^{II}-Cu^I couple in aqueous solutions. Earlier, Dainton pointed out the inadequacy of employing a thermodynamic argument to describe such a kinetic process (*10*), in order to explain his results on the oxidation of the α-polyacrylamide radical. Some of his data are given in Table 4. No doubt far more quantitative data on a variety of metal systems and free radicals are needed before a cogent formulation of this general problem can be generated.

The oxidation of alkyl radicals by Pb^{IV} esters which is represented by reaction 29 is, formally, an electron transfer process. If the formulation given above concerning the duality of oxidation mechanisms is applicable, oxidation of alkyl radicals by ligand transfer should also be possible. If halide salts are added to a mixture of Pb^{IV} acetate and carboxylic acid, a smooth decarboxylation ensues, and alkyl halides are formed, in excellent yields, in place of the usual products of oxidation (*34*). Mixed halocarboxylato Pb^{IV} species formed by metathesis with Pb^{IV} esters,

$$Pb^{IV}(O_2CR)_4 + X^- \leftrightharpoons Pb^{IV}X(O_2CR)_3 + O_2CR^-, \quad (32)$$

are, we feel, responsible for transferring halogen to the free radical in a ligand transfer process.

$$R\cdot + ClPb^{IV}(O_2CR)_3 \rightarrow RCl + Pb^{III}(O_2CR)_3 \text{ etc.} \quad (33)$$

In this respect the parallel between Cu^{II} and Pb^{IV} complexes toward alkyl radicals is striking. Electron transfer oxidation with carboxylate esters of both metal ions produces qualitatively the same products. Further, halo Pb^{IV} species, like Cu^{II} halides, afford the same alkyl halides from a variety of alkyl radicals. Such rearrangement-prone moieties as neopentyl and the homoallylic systems are converted intact.

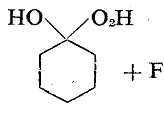
Synthesis with Free Radicals and Metal Complexes

The rapidity of many of the reactions of free radicals and metal complexes can be utilized to effect organic syntheses. Metal salts and complexes, particularly salts and complexes of cobalt and manganese, are effective catalysts for autoxidation of a variety of organic substrates. This is due to the fact that they effect oxidation and reduction of free radicals as intermediates and, in turn, oscillate between several oxidation states (*35*).

Hydrogen peroxide and its organic derivatives are catalytically decomposed by a variety of transition metal ions, and the free radicals thus generated participate in further oxidation and reduction reactions (*36*). An example is the novel use of Fenton's reagent [Fe^{II} and H_2O_2] to add two azido moieties to alkenes by the following suggested mechanism (*37*):

$$\begin{array}{ll}
Fe^{II} + H_2O_2 & \rightarrow Fe^{III} + HO\bullet \\
HO\bullet + Fe^{III}N_3 & \rightarrow Fe^{III} + N_3\bullet \\
N_3\bullet + {>}{=}{<} & \rightarrow N_3\text{-}\overset{|}{\underset{|}{C}}\text{-}\overset{|}{\underset{|}{C}}\bullet \\
N_3\text{-}\overset{|}{\underset{|}{C}}\text{-}\overset{|}{\underset{|}{C}}\bullet + Fe^{III}N_3 & \rightarrow N_3\text{-}\overset{|}{\underset{|}{C}}\text{-}\overset{|}{\underset{|}{C}}\text{-}N_3 + Fe^{II}
\end{array}$$

The adduct of cyclohexanone and hydrogen peroxide reacts with Fe^{II} to produce the ω-carboxypentyl radical.



$$+ Fe^{II} \rightarrow HO_2C(CH_2)_4CH_2\bullet + Fe^{III}$$

Synthetic schemes have been devised to produce dodecanedioic acid by dimerization, or the C_{20}-diendioic acids by addition to butadiene followed by dimerization of the allylic radicals. The carboxypentyl radical can be oxidized to ω-hexenoic acid through the use of Cu^{II} sulfate in an electron transfer oxidation, or to a variety of ω-substituted hexanoic acids through the addition of suitable ligands, Y, such as Cl^-, Br^-, I^-, N_3, SCN^-, and cyanide.

$$HO_2C(CH_2)_4CH_2\bullet + Fe^{III}Y \rightarrow HO_2C(CH_2)_4CH_2Y + Fe^{II}$$

A variety of organic substrates RH are selectively oxidized by *t*-butyl peresters through the use of transition metal salts, particularly copper, as catalysts:

$$R{-}H + t\text{-}BuO_2Ac \xrightarrow{CuX_n} R{-}OAc + t\text{-}BuOH$$

It can be shown (*38*) that the relevant oxidation steps in the mechanism are:

$$\begin{array}{ll}
t\text{-}BuO_2Ac + Cu^{I} & \rightarrow t\text{-}BuO\bullet + Cu^{II}OAc \\
t\text{-}BuO\bullet + RH & \rightarrow t\text{-}BuOH + R\bullet \\
R\bullet + Cu^{II}OAc & \rightarrow t\text{-}ROAc + Cu^{I}
\end{array}$$

The telomerization of polyhalocarbon compounds such as carbon tetrachloride and chloroform with alkenes is catalyzed by Fe^{III} and Cu^{II} halides (*39*). With chloroform the adduct is different from that obtained by the usual photochemical or peroxide-induced addition. The structure of the products is consistent with the propagation sequence in which the metal halide acts in a ligand transfer capacity.

$$Fe^{II} + HCCl_3 \rightarrow Fe^{III}Cl + H\dot{C}Cl_2$$

$$H\dot{C}Cl_2 + {>}{=}{<} \rightarrow HCl_2C\text{-}\overset{|}{\underset{|}{C}}\text{-}\overset{|}{\underset{|}{C}}\bullet$$

$$HCl_2C\text{-}\overset{|}{\underset{|}{C}}\text{-}\overset{|}{\underset{|}{C}}\bullet + Fe^{III}Cl \rightarrow$$

$$HCl_2C\text{-}\overset{|}{\underset{|}{C}}\text{-}\overset{|}{\underset{|}{C}}Cl + Fe^{II} \text{ etc.}$$

The mechanism of the Sandmeyer reaction with aromatic diazonium ions involves a series of oxidation and reduction reactions of $Cu^{I}{-}Cu^{II}$ halides with phenyl radicals and is a catalytic process (*40*). Homologation with alkenes is known as the Meerwein reaction and is a useful method for synthesizing aralkyl halides. Particularly interesting examples of a homogeneous autoxidation catalyst are the Cu^{II} amine complexes which have been used by Brackman (*41*) to effect oxidation of a variety of alcohols and ketones. The same catalyst can also be used to hydroxylate and polymerize phenols.

Summary

The mechanism of many organic oxidation and reduction reactions can be described in terms of the formation and reaction of free radicals with metal complexes. Redox (trace-metal) catalysis also involves the oxidation and reduction of radical intermediates with a metal species which oscillates between several oxidation states (*4*).

The oxidation and reduction of free radicals with metal complexes follow two general mechanisms, electron transfer and ligand transfer. Direct analogy exists with wholly inorganic descriptions of outer-sphere and inner-sphere processes.

In an electron transfer or outer-sphere mechanism the redox process is derived largely by transfer of an electron from reductant to oxidant, with only indirect contributions from the solvent and ligand. Carbonium ion intermediates and transition states are important considerations, and the scission of the β-hydrogen bond is minor during oxidation of alkyl radicals to alkenes. In contrast, ligand transfer or inner-sphere mechanism demands maximum involvement of the ligand in the transition state. Free-radical character prevails; cationic contributions from the organic moiety are minimal.

Oxidation and reduction are conjugate processes. In an electron transfer mechanism the oxidation of alkyl radicals to carbonium ions is conceptually represented by a microscopic reverse reaction in which a carbonium ion is reduced to an alkyl radical. A similar duality exists in the interconversion of carbanions and free radicals by metal complexes.

The reversibility of the ligand transfer process is easier to observe. For example, the chlorine-transfer oxidation of alkyl radicals is represented by a microscopic reverse reduction of alkyl chlorides to alkyl radicals by cuprous chlorides. A ligand transfer counterpart of the *reduction* of radicals

$$R\bullet + Cu^{II}Cl_n \leftrightharpoons R\text{-}Cl + Cu^{I}Cl_{n-1}$$

can also be described.

Hopefully, these simple redox mechanisms will be utilized in rationalizing complex reactions and formulating new syntheses. The limited number of examples cited in this short review represent only an introduction to the vast area of chemical research to be tapped in the study of the mechanisms and the synthetic utility of oxidation-reduction reactions and catalysis.

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Nutrient Cycling

Small watersheds can provide invaluable information about terrestrial ecosystems.

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Life on our planet is dependent upon the cycle of elements in the biosphere. Atmospheric carbon dioxide would be exhausted in a year or so by green plants were not the atmosphere continually recharged by CO_2 generated by respiration and fire (*1*). Also, it is well known that life requires a constant cycling of nitrogen, oxygen, and water. These cycles include a gaseous phase and have self-regulating feedback mechanisms that make them relatively perfect (*2*). Any increase in movement along one path is quickly compensated for by adjustments along other paths. Recently, however, concern has been expressed over the possible disruption of the carbon cycle by the burning of fossil fuel (*3*) and of the nitrogen cycle by the thoughtless introduction of pesticides and other substances into the biosphere (*4*).

Of no less importance to life are the elements with sedimentary cycles, such as phosphorus, calcium, and magnesium. With these cycles, there is a continual loss from biological systems in response to erosion, with ultimate deposition in the sea. Replacement or return of an element with a sedimentary cycle to terrestrial biological systems is dependent upon such processes as weathering of rocks, additions from volcanic gases, or the biological movement from the sea to the land. Sedimentary cycles are less perfect and more easily disrupted by man than carbon and nitrogen cycles (*2*). Acceleration of losses or, more specifically, the disruption of local cycling patterns by the activities of man could reduce existing "pools" of an element in local ecosystems, restrict productivity, and consequently limit human population. For example, many agriculturalists, food scientists, and ecologists believe that man is accelerating losses of phosphorus and that this element will be a critical limiting resource for the functioning of the biosphere (*1, 5*).

Recognition of the importance of these biogeochemical processes to the welfare of mankind has generated intensive study of such cycles. Among ecologists and foresters working with natural terrestrial ecosystems, this interest has focused on those aspects of biogeochemical cycles that occur *within* particular ecosystems. Thus, information on the distribution of chemical elements and on rates of uptake, retention, and release in various ecosystems has been accumulating (*6*). Little has been done to establish the role that weathering and erosion play in these systems.

Yet, the rate of release of nutrients from minerals by weathering, the addition of nutrients by erosion, and the loss of nutrients by erosion are three primary determinants of structure and function in terrestrial ecosystems. Further, with this information it is possible to develop total chemical budgets for ecosystems and to relate these data to the larger biogeochemical cycles.

It is largely because of the complex natural interaction of the hydrologic cycle and nutrient cycles that it has not been possible to establish these relationships. In many ecosystems this interaction almost hopelessly complicates the measurement of weathering or erosion. Under certain conditions, however, these apparent hindrances can be turned to good advantage in an integrated study of biogeochemical cycling in small watershed ecosystems.

It is the function of this article (i) to develop the idea that small watersheds can be used to measure weathering and erosion, (ii) to describe the parameters of watersheds particularly suited for this type of study, and (iii) to discuss the types of nutrient-cycling problems that this model renders susceptible to attack. Finally (iv), the argument is developed that the watershed ecosystem provides an ideal setting for studies of ecosystem dynamics in general.

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