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Activation of Carbon–Hydrogen Bonds by Metal Complexes: Mechanistic, Kinetic and Thermodynamic Considerations

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Abstract

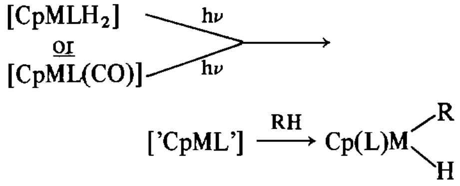
The activation and functionalization of carbon–hydrogen bonds under homogeneous conditions continues to be an important and challenging objective. Only in recent years have a few examples of such C–H bond activation, especially at aliphatic carbon centers, been reported and these involve, for the most part, stoichiometric rather than catalytic reactions; general approaches with widespread applicability still are lacking. Drawing upon the limited insights provided by these examples, and upon the pertinent information derived from related studies on the reverse process (*i.e.*, metal complex-promoted C–H bond *formation*) and on the catalytic activation of other saturated molecules (notably H_2), as well as from recent determinations of metal–carbon and metal–hydrogen bond energies, the mechanistic, kinetic and thermodynamic aspects of C–H bond activation at metal centers are analyzed. It is concluded that thermodynamic constraints, notably those associated with the characteristic weakness of metal–carbon bonds, are of dominant importance in limiting the reactivities of metal complexes toward C–H bonds.

Introduction

The homogeneous activation of C–H bonds by metal complexes, especially at saturated carbon centers, has long been recognized as an important and challenging objective [1]. Only in recent years have significant advances been made toward realizing this objective and toward achievement of an understanding of the factors that influence the mechanisms, rates and reactivity patterns associated with C–H bond activation. Even now, such successes as have been achieved involve, for the most part, *stoichiometric* reactions. *Catalytic* activation of C–H bonds resulting in useful and selective functionalization of saturated hydrocarbons remains a challenging and elusive objective [2].

Examples of recently reported homogeneous intermolecular activation of saturated hydrocarbons by metal complexes are depicted by eqn. 1–4 [2c, 2e, 3–7]:

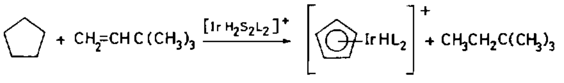
$$\begin{tikzcd} & & RD \quad (1a) \\ RH \arrow[r, dash, "Pt^{II}Cl_n"] & {} \arrow[ru, "DOAc"] \arrow[rd, "Pt^{IV}Cl_n"'] & \\ & & RCl \quad (1b) \end{tikzcd}$$



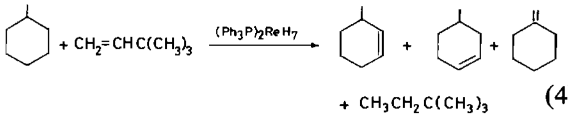
$$[CpMLH_2] \text{ or } [CpML(CO)] \xrightarrow{h\nu} [\text{'}CpML\text{'}] \xrightarrow{RH} Cp(L)M\begin{smallmatrix}R\\H\end{smallmatrix}$$

(2)

[$Cp = C_5H_5$ or C_5Me_5; M = Ir or Rh; L = CO or PMe_3]



[S = acetone; $L = PPh_3$] (3)



(4)

Because of the small number of systems in which the activation of C–H bonds has been directly observed, the information that can be derived from the study of such systems also is limited. However, considerable insights into the mechanistic, kinetic and thermodynamic aspects of C–H bond activation have been provided for some time by studies on related phenomena, including (a) the microscopic reverse of the C–H activation process, *i.e.*, metal complex-promoted C–H bond *formation* [8], (b) the activation of aromatic C–H bonds and of strained hydrocarbons which exhibit greater reactivities toward metal complexes than aliphatic

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hydrocarbons [9], (c) the activation of other saturated molecules, notably H_2 [10, 11], and (d) the determination of metal–hydrogen and metal–alkyl bond dissociation energies [12].

The present article is concerned with these themes as they relate to the mechanisms, kinetics and thermodynamics of C–H bond activation, particularly in saturated hydrocarbons.

Only the *direct* activation of C–H bonds by metal centers will be considered although it should be noted that there also are a variety of other ways in which metal complexes may be indirectly involved in C–H bond activation. Examples include H atom abstraction from hydrocarbons by OH radicals generated by reaction of Fe^{2+} with H_2O_2 [13], or by alkyl radicals generated by metal–alkyl bond dissociation (*e.g.*, coenzyme B_{12}-promoted rearrangements) [14], as well as by oxometal complexes (cytochrome P_{450} and non-enzymatic analogs thereof) [15]. Such processes appear to be particularly widespread and important in biochemical systems.

Activation of H_2 and Other Saturated Molecules

Whereas examples of the activation of saturated hydrocarbons by metal complexes have only recently been identified and still are rare, the activation of other saturated molecules, including H_2, is widespread and has been extensively studied [10, 11]. Several mechanisms for the activation of saturated molecules at metal centers have been demonstrated and are listed in Table I.

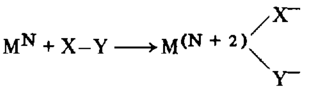
TABLE I. Mechanisms of Activation of Saturated Molecules.

1. *Electrophilic displacement*

$M^N + X\text{–}Y \longrightarrow M^NX^- + Y^+$

$[Cu^{2+} + H_2 \longrightarrow CuH^+ + H^+]$

2. *One-center oxidative addition*



$M^N + X\text{–}Y \longrightarrow M^{(N+2)}\begin{smallmatrix} X^- \\ Y^- \end{smallmatrix}$

$[Rh^ICl(PPh_3) + H_2 \longrightarrow Rh^{III}H_2Cl(PPh_3)_3]$

3. *Two-center oxidative addition*

$2M^N \text{ (or } M_2^N) + X\text{–}Y \longrightarrow M^{(N+1)}X^- + M^{(N+1)}Y^-$

$[2Co^{II}(CN)_5{}^{3-} + H_2 \longrightarrow 2Co^{III}H(CN)_5{}^{3-}]$

4. *Homolytic displacement*

$M^N + X\text{–}Y \longrightarrow M^{(N+1)}X^- + Y^-$

$[Co^{II}(CN)_5{}^{3-} + CH_3I \longrightarrow Co^{III}(CN)_5I^{3-} + CH_3\cdot]$

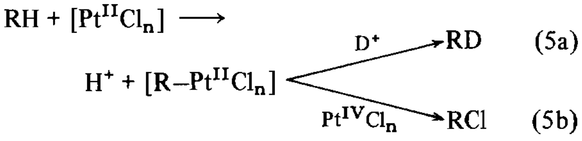
5. *Nucleophilic displacement*

$M^N + X\text{–}Y \longrightarrow M^{(N+2)}X^- + Y^-$

$[Mn^{-I}(CO)_5{}^- + CH_3I \longrightarrow Mn^I(CO)_5CH_3 + I^-]$

Of these mechanisms, only the first three (*i.e.*, electrophilic displacement and 1- and 2-center oxidative addition) have been identified for H_2 and are expected to be applicable to C–H bonds. Because of the weakness of M–H and M–R bonds, one-center homolytic displacement is expected to be thermodynamically highly unfavorable for H_2 and C–H bonds and is not observed. (However, the reverse reaction, *e.g.*, $L_nM\text{–}H + R\cdot \rightarrow L_nM\cdot + RH$, is facile and widespread.) [16a]. Nucleophilic displacement requires a good anionic leaving group such as halide or carboxylate and, thus, is not expected (nor found) to be applicable to H_2 or C–H bonds.

Of the known examples of C–H bond activation in saturated hydrocarbons, reactions (1a) and (1b) probably proceed via initial electrophilic displacement [2c, 2e] (eqn. 5) whereas the reactions depicted by eqn. 2 clearly involve oxidative addition. For the reactions depicted by eqn. 3 and 4 the initial mode of activation of the hydrocarbon has not been identified.



$$RH + [Pt^{II}Cl_n] \longrightarrow H^+ + [R\text{–}Pt^{II}Cl_n] \begin{cases} \xrightarrow{D^+} RD & (5a) \\ \xrightarrow{Pt^{IV}Cl_n} RCl & (5b) \end{cases}$$

All the recognizable potential routes to activating C–H bonds by metal complexes (Table I) involve the initial formation of metal–alkyls and/or hydrides, *i.e.*, the trading off of M–H and/or M–C bonds for C–H bonds. Thus, the thermodynamic constraints that govern such processes are dependent upon the magnitudes of the M–H and M–C bonds involved. Accordingly, it is useful to consider the available information concerning such bond dissociation energies.

Metal–Hydrogen Bond Dissociation Energies

Some of the available data concerning metal–hydrogen bond dissociation energies, both for simple diatomic metal hydrides and for complex hydrides, are listed in Table II. Most of these bond dissociation energies are seen to lie close to 60 kcal/mol and no pronounced trends with regard to the nature of the metal or of the ligands are observed. Few known metal–hydrogen bond dissociation energies exceed 65 kcal/mol.

Metal–Alkyl Bond Dissociation Energies

Fewer data are available concerning transition metal–alkyl bond dissociation energies and only recently have reliable and reasonably general methods

TABLE II. Transition Metal–Hydrogen Bond Dissociation Energies.

Metal hydride	M–H bond dissociation energy (Kcal/mol)	Ref.
$[Cr{-}H]_g$	66, 42	19a/16b
$[Mn{-}H]_g$	55	19a
$[Co{-}H]_g$	39	20
$[Ni{-}H]_g$	68	19a
$[Cu{-}H]_g$	66	19a
$[Ag{-}H]_g$	53	19a
$[Au{-}H]_g$	74	19a
$[Pt{-}H]_g$	80	19a
$[Cr^+{-}H]_g$	35	16c
$[V^+{-}H]_g$	50	16d
$[Mn^+{-}H]_g$	53	16c
$[Fe^+{-}H]_g$	58	16c
$[Co^+{-}H]_g$	52	16c
$[Ni^+{-}H]_g$	43	16c
$[Zn^+{-}H]_g$	60	16c
$[H{-}Co(CO)_4]$	58	16e
$[H{-}Co(CN)_5{}^{3-}]_{aq}$	58	16f
$[H_2{-}IrCl(CO)(PPh_3)_2]$	60 (mean)	16g
$[H{-}V(CO)_6{}^+]$	56	16h
$[H{-}Cr(CO)_6{}^+]$	58	16h
$[H{-}Mo(CO)_6{}^+]$	65	16h
$[H{-}W(CO)_6{}^+]$	64	16h

TABLE III. Cobalt–Alkyl Bond Dissociation Energies [17].

Compound[a]	Co–R bond dissociation energy (Kcal/mol)
$(4\text{-}NH_2\text{-}py)(DH)_2Co{-}CH(CH_3)C_6H_5$	21
$(4\text{-}CH_3\text{-}py)(DH)_2Co{-}CH(CH_3)(C_6H_5)$	20
$(py)(DH)_2Co{-}CH(CH_3)C_6H_5$	20
$(4\text{-}NC\text{-}py)(DH)_2Co{-}CH(CH_3)C_6H_5$	18
$(Imidazole)(DH)_2Co{-}CH(CH_3)C_6H_5$	21
$(PMe_2Ph)(DH)_2Co{-}CH(CH_3)C_6H_5$	24
$(P^nBu_3)(DH)_2Co{-}CH(CH_3)C_6H_5$	21
$(PEtPh_2)(DH)_2Co{-}CH(CH_3)C_6H_5$	19
$(PPh_3)(DH)_2Co{-}CH(CH_3)C_6H_5$	17
$(py)(SALOPH)Co{-}CH_2CH_2CH_3$	25
$(py)(SALOPH)Co{-}CH(CH_3)_2$	20
$(py)(SALOPH)Co{-}CH_2C(CH_3)_3$	18
$(py)(SALOPH)Co{-}CH_2C_6H_5$	22

[a] py = pyridine, DH_2 = dimethylglyoxime, SALOPH = N,N′-disalicylidene-*o*-phenylenediamine.

for determining such bond dissociation energies become available [12]. The most extensive measurements to date have been made on various cobalt–alkyl compounds [17], the choice reflecting, at least in part, the relevance of such compounds as models for coenzyme B_{12} and the role of cobalt–carbon bond homolysis in the mechanisms of coenzyme B_{12}-promoted reactions [16a]. Values of cobalt–alkyl bond dissociation energies are summarized in Table III.

The known values of cobalt–alkyl bond dissociation energies are seen to lie in the range 18–25 kcal/mol. Several known manganese–alkyl bond dissociation energies (*i.e.*, of $R{-}Mn(CO)_5$ or $R{-}Mn(CO)_4(PR_3)$) are of comparable magnitude [18, 19]. Thus, such values seem typical, at least of first row transition metal–alkyls.

An extensive body of evidence suggests that the characteristic weakness of such transition metal–alkyl bonds, notably in octahedral complexes, is due to steric crowding and that steric factors play a dominant role in influencing transition metal–alkyl bond dissociation energies [12, 17]. The following evidence supports these conclusions:

1. The data in Table III reveal a marked inverse dependence of the Co–R bond dissociation energy on the size of either the alkyl group or of the *trans*-axial ligand, L (*cf.* $[CH_3CH_2CH_2{-}Co(SALOPH)py]$ *vs.* $[(CH_3)_3CCH_2{-}Co(SALOPH)py]$ and $[C_6H_5(CH_3)CH{-}Co(DH)_2(PMe_2Ph)]$ *vs.* $[C_6H_5(CH_3)CH{-}Co(DH)_2(PPh_3)]$). Electronic factors also influence these trends but their effects are relatively small compared with those of steric factors [17d].

2. The typically large difference between M–H and M–C bond dissociation energies (commonly > 30 kcal/mol *vs.* 15–20 kcal/mol for $D_{C-H} - D_{C-C}$) is difficult to interpret in terms of electronic factors. The most plausible explanation is in terms of sterically-induced weakness of the metal alkyl bond. Consistent with this is the observation that dissociation energies of sterically uncrowded 'bare' metal alkyls (*i.e.*, having no other ligands) typically are higher than those of coordinately saturated (ligated) alkyls and the difference between M–H and M–C bond dissociation energies in such compounds typically is small, *e.g.*, $D_{Co-CH_3} = 41$ kcal/mol, $D_{Co-H} = 39$ kcal/mol, $D_{Co^+-CH_3} = 61$ kcal/mol, $D_{Co^+-H} = 52$ kcal/mol [20].

3. Structural studies on a range of $[R{-}Co(DH)_2L]$ compounds (DH_2 = dimethylglyoxime) reveal marked lengthening of the Co–C and Co–L (L = H_2O, pyridine, PPh_3, *etc.*) bonds with increasing size of R or L, as well as significant sterically-induced distortions of the dimethylglyoxime ligands. Thus, the Co–C bond length in $[R{-}Co(DH)_2py]$ is 1.998, 2.060 and 2.085 Å for R = CH_3, $CH_2C(CH_3)_3$ and $CH(CH_3)_2$, respectively [21]. The corresponding Co–N bond lengths are 2.068, 2.081 and 2.099 Å, respectively [21].

At this stage there are virtually no reliable data concerning metal–alkyl bond dissociation energies of conventional (*i.e.*, ligated) complexes of the second and third row transition metal series. It is possible that such metal–alkyl bonds will turn out to be stronger than those of the first transition series but convincing evidence for this is lacking. In particular it should be noted that (a) transition metal–hydrogen bond dissociation energies (which are expected to parallel the corresponding metal–alkyl bond energies) do not appear to exhibit a significant increasing trend in going from the first transition series to the second and third (Table II) and (b) gas phase determinations of bond dissociation energies in 'bare' transition metal–alkyl ions also fail to reveal a consistent increasing trend of bond dissociation energies in going from the first to the second series. Thus, $D_{M^+—CH_3}$ = 58 *vs.* 54 kcal/mol for Fe^+–R *vs.* Ru^+–R, 61 *vs.* 47 kcal/mol for Co^+–R *vs.* Rh^+–R and 48 *vs.* 59 kcal/mol for Ni^+–R *vs.* Pd^+–R [22]. As elaborated later, the widespread and facile reductive elimination of alkanes from *cis*-hydridoalkyl compounds of the second and third row, as well as first row, transition metals also attests to the typical weakness of metal–alkyl bonds in such compounds.

Considerably higher Th–R bond dissociation energies (up to 80 kcal/mol) have been estimated for some thorium alkyl compounds $[(C_5Me)_5ThR_2]$ but these are based on assumed values of the corresponding Th–OR bond dissociation energies, the reliability of which is unclear [23].

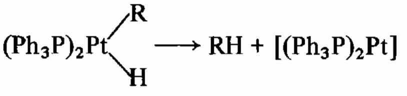
C–H Bond-Forming Reductive Elimination Reactions

Although examples of the activation of aliphatic C–H bonds by metal complexes through routes such as those depicted in Table I still are rare, examples of the reverse processes, *i.e.*, C–H *bond-forming* reactions at metal centers, are widespread and have been extensively studied [8]. Eqn. 6–9 depict examples of such reactions that correspond to the microscopic reverse of the first four entries in Table I.

Electrophilic Displacement

$$RHg^+ + H^+ \longrightarrow RH + Hg^{2+} \qquad (6)$$

One-Center Reductive Elimination [24]



$$(Ph_3P)_2Pt\begin{smallmatrix}R\\H\end{smallmatrix} \longrightarrow RH + [(Ph_3P)_2Pt] \qquad (7)$$

Two-Center Reductive Elimination [18]

$$[R\text{-}Mn(CO)_5] + [HMn(CO)_5] \longrightarrow RH + [Mn_2(CO)_{10}] \qquad (8)$$

Homolytic Displacement [18]

$$[H\text{-}Mn(CO)_5] + R\cdot \longrightarrow RH + [\cdot Mn(CO)_5] \qquad (9)$$

The widespread and facile occurrence of all these processes demonstrates not only that they are thermodynamically favorable but also that the kinetic barriers associated with them typically are small. An inescapable corollary is that the constraints associated with the reverse processes, *i.e.*, the activation of aliphatic C–H bonds by typical metal complexes (such as the products of reactions 6–9) are thermodynamic rather than kinetic. Thus, the free energy profile for reaction (7), which occurs rapidly at –20 °C, must resemble that of Fig. 1.

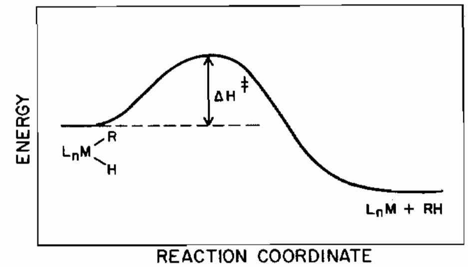
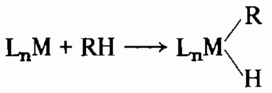


Fig. 1. Schematic energy *vs.* reaction coordinate profile for the reductive elimination of an alkane from a *cis*-hydrido-alkyl metal complex.

The conclusion that such C–H bond-forming reductive elimination reactions typically are exothermic is not surprising and, indeed, is predicted by the pattern of metal–alkyl and metal–hydrogen bond energies previously discussed. Thus, choosing $D_{M—H}$ = 60 kcal/mol, $D_{M—R}$ = 30 kcal/mol and $D_{R—H}$ = 100 kcal/mol, a reaction such as that depicted by eqn. 10 can be deduced to be endothermic by *ca.* 10 kcal/mol. Similar reasoning leads to the conclusion that the oxidative addition of unstrained saturated C–C bonds [*e.g.*, $L_nM + C_2H_6 \rightarrow L_nM(CH_3)_2$] also is likely to be thermodynamically highly unfavorable. Such processes are, however, known for strained hydrocarbons such as cyclopropane [9]:



$$L_nM + RH \longrightarrow L_nM\begin{smallmatrix}R\\H\end{smallmatrix} \qquad (10)$$

It should be noted that, although few, there are some exceptions to this generalization, *i.e.*, *cis*-alkyl hydrides which do *not* undergo spontaneous intramolecular reductive elimination of alkanes and apparently are thermodynamically stable. These include *cis*-$OsH(CH_3)(CO)_4$ (which undergoes intermolecular rather than intramolecular reductive elimination) [25], and the products of reaction (2), (*i.e.*, $(C_5Me_5)(L)MHR$, M = Rh or Ir, L = CO or PMe_3) [3–5].

Mechanisms of Activation of Saturated Hydrocarbons

Electrophilic Displacement

This would appear to be the most promising approach to the activation of saturated hydrocarbons for two reasons.

(a) The low reactivity of saturated hydrocarbons has its origin, at least in part, in the absence of low-lying unoccupied orbitals. Hence, the transition states in reactions of such molecules often involve the promotion of electrons into antibonding orbitals. Accordingly, their reactivities are expected to be highest toward species having low-lying unoccupied orbitals, *i.e.*, electrophiles. Similar considerations apply to other saturated molecules such as H_2 and it is significant that electrophilic displacement, exemplified by eqn. 11, was the first mode of homogeneous catalytic activation of H_2 to be identified [26]:

$$Cu^{2+} + H_2 \longrightarrow CuH^+ + H^+ \qquad (11)$$

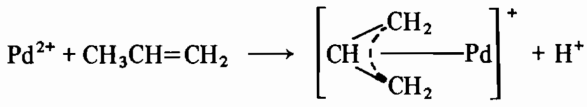
(b) Electrophilic dissociation of C–H bonds, in accord with eqn. 12, is not as severely limited as is oxidative addition by thermodynamic constraints associated with the weakness of M–H and M–C bonds. Although such constraints still play a role, the thermodynamic driving force for processes such as (12) may also be influenced by stabilization of the leaving group, *i.e.*, H^+ or R^+. In line with this the rates of electrophilic splitting of H_2 by metal ions such as Cu^{2+} (eqn. 11), Ag^+ and Hg^{2+} are enhanced by bases which serve to stabilize the released proton [10, 11]:

$$L_nM^+ + RH \longrightarrow \begin{cases} L_nMR + H^+ & (12a) \\ \text{or} & \\ L_nMH + R^+ & (12b) \end{cases}$$

Electrophilic attack by metal ions on *activated* (*e.g.*, aromatic or allylic) C–H bonds is a familiar process, *e.g.* [27, 28],

$$[Rh(OEP)]^+ + C_6H_6 \longrightarrow [C_6H_5Rh(OEP)] + H^+ \qquad (13)$$

(OEP = octaethylporphyrin)



$$Pd^{2+} + CH_3CH{=}CH_2 \longrightarrow [CH(CH_2)_2Pd]^+ + H^+$$

(14)

Electrophilic attack on aliphatic C–H centers is expected to be more difficult but is precedented. As previously noted, the activation of alkanes by Pt^{II} (eqn. 5) probably proceeds through such a mechanism. The hydrogen transfer reactions of cycloalkanes promoted by cationic iridium complexes (eqn. 3) also may involve initial electrophilic attack although other mechanisms (*e.g.*, C–H oxidative addition) cannot be ruled out. Finally, the activation of CH_4 by lanthanide complexes (eqn. 15), for which oxidative addition is an unlikely process, may well involve electrophilic attack [29].

$$(\eta^5\text{-}C_5H_5)_2M\text{-}CH_3 + {}^{13}CH_4 \rightleftarrows (\eta^5\text{-}C_5H_5)_2M\text{–}{}^{13}CH_3 + CH_4 \qquad (15)$$

[M = Lu or Y]

One-Center Oxidative Addition

As already noted, a principal constraint on the activation of C–H bonds by oxidative addition (eqn. 10) is thermodynamic, reflecting the relative weakness of M–H and M–C bonds, compared with C–H bonds. Thus, strategies to overcome this constraint should seek to maximize the M–H and M–C bond energies. Since M–H bond dissociation energies seem to be relatively invariant, this effectively reduces to trying to maximize the M–C bond dissociation energy.

One strategy for accomplishing this is to design coordinately unsaturated complexes that are highly unstable. The observation [25] that *cis*-$OsH(CH_3)(CO)_4$ resists intramolecular reductive elimination of CH_4, suggests that the oxidative addition of CH_4 to $[Os(CO)_4]$ should be thermodynamically favorable and that the latter should be capable of activating saturated hydrocarbons under mild conditions. Unfortunately, such species do not appear to be readily thermally accessible and typically must be generated photochemically as in the examples of eqn. 2 (*i.e.*, [CpML], M = Rh or Ir).

Since steric factors appear to play a dominant role in determining the strength of metal–alkyl bonds, the most promising candidates for C–H bond activation would appear to be the least sterically hindered metal complexes. In accord with this, 'bare' metal atoms or ions such as Co^+, Fe^+ and Ni^+ have been found to activate alkanes in the gas phase through processes such as those depicted by eqn. 16 which, however, typically are endothermic [20]:

$$Co^+ + C_2H_6 \longrightarrow \begin{cases} \underset{\text{(postulated)}}{[H\text{–}Co\text{–}C_2H_5]^+} \longrightarrow [CoH]^+ + C_2H_5 & (16a) \\[2ex] \underset{\text{(postulated)}}{[CH_3\text{–}Co\text{–}CH_3]^+} \longrightarrow [CoCH_3]^+ + CH_3 & (16b) \end{cases}$$

Also in line with this is the observation that the liganded transition metal complexes that have been found to be effective in activating aliphatic C–H bonds, *i.e.*, CpML (Cp = C_5H_5 or C_5Me_5, M = Rh or Ir, L = CO or PMe_3) have very low coordination numbers, *i.e.*, 2 [3–5]. (While Cp is a 6-electron donor it probably is not as sterically demanding as three 2-electron donor ligands.)

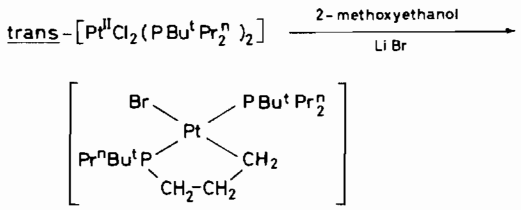
Our assertion that the kinetic barrier for C–H oxidative addition is low (*i.e.*, that the failure of most 16-electron complexes to activate C–H bonds by oxidative addition is due to *thermodynamic* constraints) is further supported by the following observations:

(a) For those metal complexes that do activate C–H bonds by oxidative addition the kinetic selectivity typically is small. Thus, the relative rate-constants for the oxidative addition of various C–H bonds to $[(\eta^5\text{-}C_5Me_5)(PMe_3)Ir]$ (normalized to the C–H bonds of cyclohexane) are cyclopropane (2.6), cyclopentane (1.6), cyclohexane (1.0), cyclooctane (0.1), neopentane (1.1), benzene (3.9), a remarkably small spread of only about one order of magnitude [3a]. We conclude that the more widespread facility of metal complexes to oxidatively add aromatic C–H bonds, compared with aliphatic ones, is due to thermodynamic, rather than kinetic, factors, *i.e.*, metal–aryl bonds are stronger than metal–alkyl bonds. Thus, $D_{C_6H_5-Rh(C_5Me_5)(PMe_3)} - D_{CH_3-Rh(C_5Me_5)(PMe_3)} = 13$ kcal/mol [5a]. Further, in line with this, we do not consider that prior coordination of an unsaturated hydrocarbon to a metal center (*e.g.*, of an arene or alkene through a C=C bond), *even where such coordination actually is observed* [5a] contributes significantly to the reactivity of the C–H bond, *i.e.* to the rate of the C–H activation process.

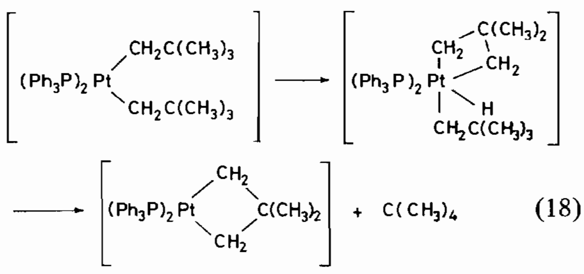
(b) The reactivities of $[(\eta^5\text{-}C_5Me_5)IrL]$ complexes toward alkanes seem to be relatively insensitive to the nature of L, *i.e.*, whether L = CO or PMe_3, suggesting that the kinetic barrier in both cases is small [3–5]. (It is possible that the difference in the electronic influences of CO and PMe_3 is compensated by steric differences.)

(c) The reactivity of $[(\eta^5\text{-}C_5Me_5)Ir(CO)]$ toward CH_4 to form $[(\eta^5\text{-}C_5Me_5)(CO)IrH(CH_3)]$ exhibits only a small temperature dependence and persists to temperatures as low as 12 K [30]. This implies a very small activation barrier for the oxidative addition of CH_4.

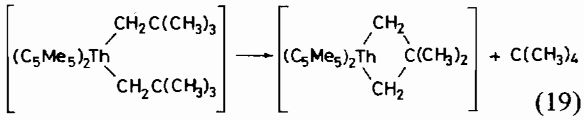
The commonly perceived greater facility of *intramolecular* C–H oxidative addition (exemplified by eqn. 17), compared with *intermolecular* oxidative addition, also seems attributable, at least in part, to steric factors. The apparent reactivity difference for such intramolecular oxidative addition, compared with the corresponding intramolecular reactivity, toward a 'neat' hydrocarbon solvent, would seem to be greater than can readily be accounted for simply in terms of entropic effects. Instead, we propose that (in addition to a more favorable entropy of activation) replacement of two ligands by a bidentate chelating ligand may reduce the interligand repulsion and, thus, strengthen the metal–alkyl bond. Such an effect has been invoked [31, 32] to account for the thermodynamic driving force that drives reactions such as (18) and (19), in which there is no net change in the number of M–C and C–H bonds. In these cases the steric destabilization of the bis(neopentyl)metal reactants probably is unusually severe because of the steric bulk of the neopentyl groups.



(17)

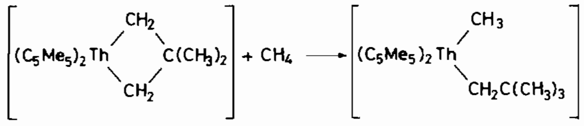


(18)



(19)

Using a smaller alkane, which gives rise to a less sterically hindered dialkyl, reaction (19) apparently can be 'reversed' [33], *i.e.*,



(20)

Consistent with our contention that the commonly perceived greater ease of intramolecular (compared with intermolecular) C–H oxidative addition has its origin in thermodynamic rather than kinetic differences, it is generally found that those complexes for which intermolecular C–H oxidative addition is thermodynamically allowed, do not exhibit a significant preference for intramolecular reaction. This is the case for oxidative addition both of aliphatic and of aromatic C–H bonds [3–5, 34].

Two-Center Oxidative Addition

$$2L_nM \text{ (or } M_2L_{2n}) + RH \longrightarrow L_nM\text{-}R + L_nM\text{-}H \qquad (21)$$

While known for H_2 (*e.g.*, $2[Co(CN)_5{}^{3-}] + H_2 \rightleftarrows 2[HCo(CN)_5{}^{3-}]$ or $[Co_2(CO)_8] + H_2 \rightleftarrows 2[H\text{-}Co(CO)_4]$) this type of process is rare for C–H bonds. The same type of thermodynamic constraints that disfavor one-center oxidative addition of C–H bonds to 16-electron complexes also disfavor the related two-center reactions, *i.e.*, oxidative addition to 17-electron complexes. Indeed, the thermodynamic constraints would be expected to be more severe since, in general, such 17-electron species capable of undergoing one-electron oxidative addition, such as $Mn(CO)_5$ or $CpFeL_5$, also would be expected to form reasonably stable metal–metal bonds whose dissociation would contribute an additional thermodynamic barrier to oxidative addition.

In accord with this the reactivities of binuclear complexes toward two-center oxidative addition have generally been found to be low. On the other hand, binuclear C–H bond-forming reductive elimination reactions, such as that depicted by eqn. 8, commonly are facile.

Among the few examples of binuclear C–H oxidative additions are the recently reported reactions (eqn. 22) of $[Rh_2(OEP)_2]$ (OEP = octaethylporphyrin) with hydrocarbons containing *activated* C–H bonds (R–H = $C_6H_5CH_2$–H or 9-10-dihydroanthracene) [35]:

$$[Rh_2(OEP)_2] + RH \longrightarrow [(OEP)Rh\text{-}R] + [(OEP)Rh\text{-}H] \qquad (22)$$

Conclusions

Only within the past decade have examples of the homogeneous intermolecular activation of aliphatic C–H bonds been identified and studied. However, examples of the microscopic reverse of such processes, *i.e.*, of the formation of C–H bonds from metal alkyls and/or hydrides, have long been familiar and their study has provided important insights into the mechanistic, kinetic and thermodynamic considerations that govern C–H bond activation by metal complexes. The development of general, widely applicable approaches to C–H bond activation, particularly those involving *catalytic* utilization of the metal complex and leading to useful and selective functionalization of C–H bonds, continues to be an important and challenging objective. Thermodynamic constraints, notably those associated with the characteristic weakness of transition metal–alkyl bonds, would appear to constitute the dominant obstacle to achieving this objective. Thus, importance must be accorded to the achievement of a better understanding of the factors that influence transition metal–alkyl bond energies and that might be manipulated to design organometallic complexes with the potential for forming stronger metal–alkyl bonds.

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References

1 (a) J. Halpern, *Adv. Chem. Ser.*, *70*, 1 (1966); (b) J. Halpern, *Discuss. Faraday Soc.*, *7*, 46 (1968).

2 (a) G. W. Parshall, *Acc. Chem. Res.*, *8*, 114 (1975); (b) D. E. Webster, *Adv. Organomet. Chem.*, *15*, 147 (1977); (c) A. E. Shilov and A. A. Shteinman, *Coord. Chem. Rev.*, *24*, 97 (1977); (d) G. W. Parshall, *Catalysis* (Chemical Society Specialist Periodical Report), *1*, 335 (1977); (e) A. E. Shilov, *Pure Appl. Chem.*, *50*, 725 (1978); (f) R. G. Bergman, *Science*, *223*, 902 (1984).

3 (a) A. H. Janowicz and R. G. Bergman, *J. Am. Chem. Soc.*, *105*, 3020 (1983); (b) M. J. Wax, J. M. Stryker, J. M. Buchanan, C. A. Kovac and R. G. Bergman, *J. Am. Chem. Soc.*, *106*, 1121 (1984); (c) R. A. Perlana and R. G. Bergman, *Organometallics*, *3*, 508 (1984).

4 (a) J. K. Hoyana and W. A. G. Graham, *J. Am. Chem. Soc.*, *104*, 3723 (1982); (b) *J. Am. Chem. Soc.*, *105*, 7190 (1983).

5 (a) W. D. Jones and F. J. Feher, *J. Am. Chem. Soc.*, *106*, 1650 (1984); (b) W. D. Jones and F. J. Feher, *Organometallics*, *2*, 562 (1983).

6 (a) R. H. Crabtree, M. F. Mellea, J. M. Mihelcic and J. M. Quirk, *J. Am. Chem. Soc.*, *104*, 107 (1982); (b) R. H. Crabtree, *Chemtech*, 506 (1982); (c) R. H. Crabtree, C. P. Parnell and R. J. Uriarte, *Organometallics*, *3*, 816 (1984).

7 (a) D. Baudry, M. Ephritikhine and H. Felkin, *J. Chem. Soc., Chem. Commun.*, 1243 (1980); (b) D. Baudry, M. Ephritikhine and H. Felkin, *J. Chem. Soc., Chem. Commun.*, 606 (1982); (c) D. Baudry, M. Ephritikhine, H. Felkin and J. Zakrewski, *J. Chem. Soc., Chem. Commun.*, 1235 (1982); (d) D. Baudry, M. Ephritikhine, H. Felkin and R. Holmes-Smith, *J. Chem. Soc., Chem. Commun.*, 788 (1983).

8 J. Halpern, *Acc. Chem. Res.*, *15*, 332 (1982), and refs. therein.

9 J. Halpern, in I. Wender and P. Pino (eds.), 'Organic Syntheses in Metal Carbonyls, Vol. 2', Wiley, New York, 1977, p. 705, and refs. therein.

10 J. Halpern, *Annu. Rev. Phys. Chem.*, *16*, 103 (1965) and refs. therein.

11 B. R. James, 'Homogeneous Hydrogenation', Wiley, New York, 1977, and refs. therein.

12 J. Halpern, *Acc. Chem. Res.*, *15*, 238 (1982) and refs. therein.

13 C. Walling, *Acc. Chem. Res.*, *8*, 125 (1975) and refs. therein.

14 J. Halpern, *Pure Appl. Chem.*, *55*, 1059 (1983) and refs. therein.

15 J. T. Groves, *Adv. Chem. Ser.*, *191*, 276 (1980) and refs. therein.

16 (a) J. Halpern, *Pure Appl. Chem., 51*, 2171 (1979); (b) L. Sallans, K. Lane, R. R. Squires and B. B. Freiser, *J. Am. Chem. Soc., 105*, 6352 (1983); (c) P. B. Armentrout, L. F. Halle and J. L. Beauchamp, *J. Am. Chem. Soc., 103*, 6501 (1981); (d) N. Aristov and P. B. Armentrout, *J. Am. Chem. Soc., 106*, 4065 (1984); (e) F. Ungvary, *J. Organomet. Chem., 36*, 363 (1972); (f) B. de Vries, *J. Catal., 1*, 489 (1962); (g) L. Vaska, *Acc. Chem. Res., 1*, 335 (1968); (h) A. E. Stevens and J. L. Beauchamp, *J. Am. Chem. Soc., 103*, 190 (1981).

17 (a) J. Halpern, F. T. T. Ng and G. L. Rempel, *J. Am. Chem. Soc., 101*, 7124 (1979); (b) F. T. T. Ng, G. L. Rempel and J. Halpern, *J. Am. Chem. Soc., 104*, 621 (1982); (c) T. T. Tsou, M. Loots and J. Halpern, *J. Am. Chem. Soc., 104*, 623 (1982); (d) F. T. T. Ng, G. L. Rempel and J. Halpern, *Inorg. Chim. Acta, 77*, L165 (1983).

18 (a) M. J. Nappa, R. Santi, S. P. Diefenbach and J. Halpern, *J. Am. Chem. Soc., 104*, 619 (1982); (b) M. J. Nappa, R. Santi and J. Halpern, *Organometallics, 4*, 34 (1985).

19 (a) J. A. Connor, *Top. Curr. Chem., 71*, 71 (1977); (b) J. A. Connor, M. T. Zafarani-Moattar, J. Bickerton, N. I. El Saied, S. Suradi, R. Carson, G. Aetakhin and H. A. Skinner, *Organometallics, 1*, 1176 (1982).

20 P. B. Armentrout and J. L. Beauchamp, *J. Am. Chem. Soc., 103*, 784 (1981).

21 N. Bresciani-Pahor, M. Calligaris, G. Nardin and L. Randaccio, *J. Chem. Soc., Dalton Trans.*, 2549 (1982) and refs. therein.

22 M. L. Mandlich, L. F. Halle and J. L. Beauchamp, *J. Am. Chem. Soc., 106*, 4403 (1984).

23 J. W. Bruno, T. J. Marks and L. R. Morss, *J. Am. Chem. Soc., 105*, 6824 (1983).

24 L. Abis, A. Sen and J. Halpern, *J. Am. Chem. Soc., 100*, 2915 (1978).

25 J. Norton, *Acc. Chem. Res., 12*, 139 (1979) and refs. therein.

26 J. Halpern, E. R. Macgregor and E. Peters, *J. Phys. Chem., 60*, 1455 (1956).

27 Y. Aoyama, T. Yoshida, K. Sakurai and H. Ogoshi, *J. Chem. Soc., Chem. Commun.*, 478 (1983).

28 P. M. Maitlis, 'The Organic Chemistry of Palladium, Vol. 1', Academic Press, New York, 1971, p. 176, and refs. therein.

29 P. L. Watson, *J. Am. Chem. Soc., 105*, 6491 (1983).

30 A. J. Rest, I. Whitwell, W. A. G. Graham, J. K. Hoyano and A. D. McMaster, *J. Chem. Soc., Chem. Commun.*, 624 (1984).

31 (a) G. M. Whitesides, *Pure Appl. Chem., 53*, 287 (1981); (b) R. Di Cosimo, S. S. Moore, A. F. Sowinski and G. M. Whitesides, *J. Am. Chem. Soc., 104*, 124 (1982); (c) J. A. Ibers, R. Di Cosimo and G. M. Whitesides, *Organometallics, 1*, 13 (1982).

32 J. W. Bruno, T. J. Marks and V. W. Day, *J. Am. Chem. Soc., 104*, 7357 (1982).

33 C. M. Fendrick and T. J. Marks, *J. Am. Chem. Soc., 106*, 2214 (1984).

34 W. D. Jones and F. J. Feher, *J. Am. Chem. Soc.*, in press.

35 (a) B. Wayland and K. J. Del Rossi, *J. Organomet. Chem.*, C27 (1984); (b) R. S. Paonessa, N. C. Thomas and J. Halpern, in preparation.