

1 **KINETIC PARAMETERS FOR THE THERMAL CRACKING OF**  
2 **SIMPLE HYDROCARBONS: FROM LABORATORY TO**  
3 **GEOLOGICAL TIME-TEMPERATURE CONDITIONS**

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10 **Abstract**

11 Apparent kinetic parameters (frequency factor  $A$  and activation energy  $E_a$ ) were computed at  
12 700 bar and two temperatures: 200°C, a characteristic of deeply buried reservoirs of  
13 petroleum and 400°C an average temperature used for laboratory pyrolysis to simulate  
14 thermal cracking of petroleum. Several hydrocarbons were studied by simulation: several  
15 linear alkanes and one branched alkane, toluene, 2-methylnaphthalene, 1-,3-,5-  
16 trimethylbenzene, methylcyclohexane, butylcyclohexane, tetralin and decylbenzene. The  
17 calculations were performed using detailed kinetic models (mostly free-radical mechanisms)  
18 previously constructed. The computation of  $E_a$  does not depend on the global rate law, but we  
19 assumed a first-order reaction law to compute  $A$ . The apparent kinetic parameters of all  
20 hydrocarbons are slightly modified by temperature and rather strongly for tetralin and  
21 decylbenzene. The deviation, due to the use of Arrhenius law with the apparent kinetic  
22 parameters computed at 400°C for extrapolation to geological temperature, is important in  
23 most cases, even if the kinetic parameters do not strongly depend on temperature. In this  
24 study, most hydrocarbons appear more stable at 200°C than it could be inferred by using the  
25 extrapolation of the Arrhenius law from 400°C to 200°C, except for tetralin and decylbenzene  
26 which appear more reactive. The apparent kinetic parameters that were computed at 200°C  
27 could be implemented in global kinetic models used for the assessment of reservoirs.

## 28        1. Introduction

29        The prediction of the petroleum composition, submitted to thermal cracking over millions of  
30        years, requires a good knowledge of the reaction kinetics, particularly in deeply buried  
31        reservoirs, at high pressure (100-1000 bar) and high temperature ( $> 200^{\circ}\text{C}$ ). The apparent  
32        kinetic parameters (apparent frequency factor  $A$  and apparent activation energy  $E_a$ ) are  
33        commonly determined via pyrolysis in laboratory at high pressure of crude oils [i.e.1-13] or  
34        pure hydrocarbons [i.e. 14-24]. The geological time is compensated by higher temperatures  
35        ( $350\text{-}450^{\circ}\text{C}$ ) than those encountered in reservoir. The bulk kinetic models derived from  
36        experimental results are then computed using geological time-temperature conditions. Due to  
37        the complexity of the chemical composition of crude oils, it is easier to derive reliable kinetic  
38        parameters from pyrolysis of pure hydrocarbons or simple mixtures.

39        In order to determine precisely the main reaction pathways and obtain reliable predictions,  
40        detailed kinetic models for the thermal reaction of pure hydrocarbons or simple mixtures can  
41        also be set up. These detailed kinetic models consist of mechanisms made of elementary steps  
42        (mainly free-radical reactions) with their fundamental kinetic parameters [25-37]. That is the  
43        reason why they can be extrapolated in the pressure and temperature ranges of 100-1000 bar  
44        and  $150\text{-}450^{\circ}\text{C}$ , respectively: the main reaction pathways and the fundamental kinetic  
45        parameters remain unchanged. The main advantages of these models are their precision and  
46        the possibility for extrapolation. However, the main drawback is the size of the model (several  
47        hundred even thousand reactions) which prevents their integration in a basin model.

48        This paper aims to test if the apparent kinetic parameters are constant in the temperature range  
49        of  $200\text{-}400^{\circ}\text{C}$ , at high pressure, *i.e.* testing if the apparent kinetic parameters computed at  
50         $400^{\circ}\text{C}$  can be used at  $200^{\circ}\text{C}$  by extrapolation of the Arrhenius law in a reliable manner. A  
51        temperature equals to  $200^{\circ}\text{C}$  is typical of deeply buried oil reservoirs, and  $400^{\circ}\text{C}$  is an  
52        average temperature for laboratory pyrolysis. In this study, the selected hydrocarbons are  
53        linear and branched alkanes, several aromatic compounds (toluene, trimethylbenzene “TMB”,  
54        methylnaphthalene “MN”, decylbenzene “DB” and tetralin) and cycloalkanes  
55        (methylcyclohexane “MCH” and butylcyclohexane “BCH”). All free-radical mechanisms  
56        mentioned in this paper were constructed and validated based on experimental database  
57        recovered by our team for about 15 years ([25-28], [30-32], [34-35]). The detailed models are  
58        used to compute the apparent kinetic parameters at 700 bar and two temperatures:  $200^{\circ}\text{C}$  and

59 400°C. The kinetic parameters computed at 400°C and then 200°C are used to calculate and  
60 compare the conversion after 1, 10 or 100 million years at 200°C (depending on the reactivity  
61 of the compound), as well as the half-life. The relative scale of stability is set up and  
62 compared at both temperatures. This comparison is performed to highlight the influence of  
63 extrapolating kinetic parameters outside the range of temperatures in which they were  
64 computed. This work is the first step of the construction of bulk kinetic models reliable in the  
65 geological pressure and temperature conditions, and which could be implemented in complete  
66 basin models for petroleum exploration purposes.

## 67 2. Methodology for the determination of apparent kinetic parameters

68 Apparent kinetic parameters were first calculated at  $T_0 = 400^\circ\text{C}$  and then  $200^\circ\text{C}$ , at 700 bar.  
69 The computation of the apparent activation energy does not depend on the kinetic law, when  
70 it can be done directly by the rate, but the apparent frequency factor  $A$  does.  $A$  was computed  
71 assuming that cracking reactions follow the first-order kinetic law.

72 The models were computed by using the package of software CHEMKIN II [38], and more  
73 precisely SENKIN at constant pressure and temperature. It should be noted that this software  
74 uses the ideal gas law to calculate concentrations. Obviously, the ideal gas law is not valid at  
75 high pressure either. Consequently, all initial concentrations  $c_0$  were first calculated separately  
76 by using the Peng-Robinson equation of state [39] which is implemented in the software  
77 DIAGSIM [40]. The Peng-Robinson equation of state has been chosen because it is about two  
78 times more precise for the estimation of the densities than the other cubic equations of state as  
79 the Soave-Redlich-Kwong equation [41]. The calculation of the densities allows the  
80 calculation of the concentrations, and this calculation is performed to evaluate the ideal gas  
81 pressures that lead to the same initial concentrations than the Peng-Robinson equation of state  
82 computed at 700 bar. Simulation conditions are summarized in Table 1.

83 Simulations were performed and reached up to 30% conversion. The consumption rates  $r$   
84 were computed by CHEMKIN at 30% conversion, and by plotting  $\ln r$  vs  $-1/RT$  (where  $R$  is  
85 the ideal gas constant and  $T$  the temperature in K) with  $T = T_0, T_0 - 2, T_0 + 2$ , the slope  
86 corresponds to the apparent  $E_a$ . The first-order kinetic law ( $r = k \times c$ , with  $k$  the rate constant  
87 following the Arrhenius law,  $k$  in  $\text{s}^{-1}$  and  $c$  in  $\text{mol.m}^{-3}$ ) leads to:

$$88 \ln r = \ln A - E_a/RT + \ln c$$

89 The Y-intercept of the plots  $\ln r$  vs  $-1/RT$  allows the calculation of  $A$ . The results are presented  
90 first at laboratory temperature (400°C) and then at geological temperature (200°C). The  
91 uncertainty on the computed apparent kinetic parameters depends on: 1) the uncertainty of  
92 the intrinsic kinetic parameters of the elementary reactions in the kinetic models, 2) the  
93 choices made during the construction of the kinetic model (detailed vs lumped models), 3) the  
94 care given to the computations (strict respect of the method, especially iso-conversion), and  
95 for  $A$ : 4) the precision of the estimation of the initial concentrations. Considering the time and  
96 effort given to the calculations, the uncertainty on apparent  $E_a$  is estimated between  $\pm 0.5$   
97 kcal/mol and a factor of 2 for the apparent  $A$ .

98  
99

### 3. Kinetic parameters at laboratory temperature (400°C)

100 The hydrocarbons selected for this study are divided into three groups: alkanes (linear and  
101 branched), methylaromatics (toluene, methylnaphthalene and trimethylbenzene) and other  
102 hydrocarbons, such as naphthenes (methylcyclohexane and butylcyclohexane), decylbenzene  
103 and tetralin.

#### 104 3.1 Alkanes

105 The apparent kinetic parameters of alkanes were calculated using the detailed kinetic model of  
106 Burklé-Vitzthum et al. [30]. This model was constructed based on a lumping technique,  
107 taking into account any linear alkanes between  $C_1$  and  $C_{32}$  as well as model compounds of  
108 branched alkanes, as reactants and any crossed reactions, and making up 13 206 free radical  
109 reactions. Alkanes with several lengths ranging from  $C_5$  to  $C_{25}$  were studied. Linear alkanes  
110 were mostly studied, but the influence of branched structures was investigated by comparing  
111  $nC_{15}$  to  $isoC_{15}$ .

112 Table 2 presents the results at 400°C. It leads to the following comments: 1) kinetic  
113 parameters of  $n$ - and  $iso$ -alkanes do not exhibit any significant differences due to their similar  
114 structure and reactivity, 2) all activation energies are comprised between  $67.0 \pm 0.5$  and  
115  $69.1 \pm 0.5$  kcal/mol for any alkane chain length and they increase almost linearly with the chain  
116 length (Figure 1), 3) the frequency factor increases with the alkane length.

117 Activation energies obtained from the calculations are similar for any chain length  
118 investigated, proving the consistency of the lumping methods results with those of detailed

119 mechanisms. A mean value of 68.5 kcal/mol could be proposed, for every alkane for any  
120 structure and length ( $n > 4$ ).

121 For  $nC_8$ , the detailed model (91 free radical reactions) of Lannuzel et al. [33] was also tested,  
122 and leads to  $A = 6.5 \times 10^{16} \text{ s}^{-1}$  and  $E_a = 67.6 \text{ kcal/mol}$ , as well as the extended model of  
123 Bounaceur et al. [26] (182 free radical reactions) which leads to  $A = 6.6 \times 10^{15} \text{ s}^{-1}$  and  $E_a =$   
124  $64.0 \text{ kcal/mol}$ . The very simple model of Burklé-Vitzthum et al. [29] in the case of  $nC_{16}$  (13  
125 free radical reactions), which was constructed using a particularly drastic lumping method,  
126 leads to  $A = 2.0 \times 10^{15} \text{ s}^{-1}$  and  $E_a = 61.8 \text{ kcal/mol}$ .

127 The former mechanisms were all constructed and validated at high pressure (700 bar) and  
128 temperatures in the range of 300 - 400°C.

129 Another detailed model [42] was also tested: it was constructed to describe the cracking of  
130  $nC_{12}$  in jet fuel conditions (typically 1 bar and 800°C). In these conditions, short radicals  
131 formed via decompositions are predominant, whereas in our conditions (lower temperature  
132 and very high pressure), the predominant radicals are those corresponding to the reactant,  
133 which are formed via H-transfers [26]. Indeed high pressures favor bimolecular reactions, on  
134 the contrary to low pressures that favor unimolecular reactions. Therefore, the free radical  
135 reactions, that determine the cracking rates, are different and consequently the apparent  
136 kinetic parameters are totally different, even when the mechanism is computed at 400°C and  
137 700 bar:  $A = 4.5 \times 10^{13} \text{ s}^{-1}$  and  $E_a = 55.9 \text{ kcal/mol}$  ( $A = 1.6 \times 10^{12} \text{ s}^{-1}$  and  $E_a = 50.4 \text{ kcal/mol}$  when  
138 the mechanism is computed at  $P = 1 \text{ bar}$  and  $T = 800^\circ\text{C}$ ). This example illustrates the  
139 importance of selecting kinetic parameters that are determined in conditions close to those in  
140 which they are used.

141 Table 3 gathers the experimental literature data concerning  $n$ -alkanes in similar  $T$  and  $P$   
142 conditions, but the results appear very scattered:  $59 < E_a < 74 \text{ kcal/mol}$  and  $3 \times 10^{13} < A <$   
143  $7 \times 10^{22} \text{ s}^{-1}$ . The activation energy seems to increase with pressure, but no clear conclusion can  
144 be drawn. Agreement is obtained with Béhar and Vandenbroucke [14] and Khorasheh and  
145 Gray [20] for  $nC_{16}$  lumped in a drastic manner, and our results are consistent with those of  
146 Song et al. [46] and Zhou and Crynes [45].

147 *3.2 Methyларomatics*

148 Three methylaromatics were studied: toluene [32], 1-methylnaphthalene [35] and 1, 3, 5-  
149 trimethylbenzene [31]. Table 2 gathers the results obtained via computation of the models.  
150 The apparent activation energies of the studied methylaromatics are very close, due to their  
151 similar structures. The apparent frequency factor of trimethylbenzene is higher than that of  
152 toluene due to the reaction path degeneracy.

153 Literature data concerning several aromatics is presented in Table 3. Some data was found at  
154 the same  $P$ - $T$  conditions. Concerning toluene, literature activation energies are much higher  
155 than those calculated from our mechanisms, but they were determined at higher temperatures  
156 and at atmospheric pressure. In these conditions, the cracking mechanisms are different [32];  
157 particularly the bimolecular initiations are negligible while they are of utmost importance at  
158 high pressures. Indeed, activation energies of bimolecular initiations are 30 kcal/mol lower  
159 than unimolecular ones, but they become important only at high pressure and therefore the  
160 apparent activation energy is much lower at high pressure than at low pressure. Agreement  
161 between experimental [31] and computed activation energy of trimethylbenzene is obtained.  
162 Concerning 1-methylnaphthalene, the experimental activation energy [24] is far lower than the  
163 computed value. The determination of the experimental apparent activation energy requires  
164 the determination or the assumption of the kinetic law, which is not needed for the  
165 computation of  $E_a$  by our kinetic model using detailed mechanisms (the software CHEMKIN  
166 II computes directly the rates according to the reaction mechanisms implied). The  
167 determination of the experimental  $E_a$  is generally performed by assuming a first-order law but  
168 we demonstrated elsewhere [35] that this assumption is not always true. This may explain the  
169 differences between the experimental apparent activation energy and the computed value,  
170 especially for methylnaphthalene.

### 171 *3.3 Other hydrocarbons*

172 Methylcyclohexane [27], butylcyclohexane [34], decylbenzene [28] and tetralin [25] were  
173 also studied. The apparent kinetic parameters are in Table 2. The apparent activation energies  
174 are in the same order of magnitude at 400°C. Butylcyclohexane appears more reactive than  
175 methylcyclohexane. This is due to the cracking of the butyl side chain which is easier than the  
176 cracking of the methyl side chain. The literature data concerning naphthenes are presented in  
177 Table 3. No literature data was found for methylcyclohexane, but the computed kinetic

178 parameters of butylcyclohexane are very close to the experimental ones determined for  
179 butylcyclohexane [52, 53] and propylcyclohexane [54].

180 The computed results of tetralin are far from the values obtained by Yu and Eser [55],  
181 although they were determined in neighboring conditions. This difference can be due to the  
182 assumption of a first-order kinetic law to determine the experimental value. According to our  
183 model, it appears that this assumption is not justified [25]. Another sensitive factor is the  
184 amount of impurities [25] (particularly dihydronaphthalene) which greatly influences the  
185 reactivity of tetralin and consequently its apparent kinetic parameters. This point is not clearly  
186 addressed in the experimental study by Yu and Eser [54] and so the comparison is difficult. In  
187 our case, the computations were performed with 0.1% of dihydronaphthalene. The mechanism  
188 of 2-ethyltetralin is strongly influenced by the cracking of the side chain and so the  
189 comparison with tetralin would not be consistent.

190 Concerning long-chain ( $n \geq 4$ ) alkylbenzene in which the long chain allows the retroene  
191 reaction [28], no agreement is obtained with the literature data, except with the activation  
192 energy of Leigh and Szwarc [50] and the frequency factor of Yu and Eser. These results are  
193 surprising but once again, literature data were generally determined at higher temperatures  
194 and much lower pressures than in our study, making comparisons quite difficult.

### 195 *3.4 Relative thermal stability at 400°C*

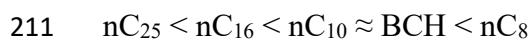
196 To establish a relative scale of stability of the investigated hydrocarbons at 400°C, the  
197 conversion  $X$  is plotted as a function of temperature  $T$  assuming a first-order reaction law  
198 (Figure 2). The pyrolysis duration is set to 10 hours, which is a typical duration for pyrolysis  
199 in the laboratory. Characteristic S-shaped curves are obtained (Figure 2). The scale of stability  
200 between compounds is determined as following: the higher the temperature to reach 50%  
201 conversion, the more stable the compound. The following scale is obtained from the least to  
202 the most stable compound:

203  $DB < C_{15} < MN < BCH < Tetralin < MCH \approx TMB < toluene$

204 It should be noted that: 1) long alkanes ( $C_{15}$ ) belong to compounds with low stability, 2)  
205 cycloalkanes are more stable than  $C_{15}$ , 3) toluene and trimethylbenzene are the most stable  
206 compounds and behave very differently from methylnaphthalene and decylbenzene, 4)

207 alkylbenzenes are increasingly reactive as their side chain length increases, 5) the stability of  
208 the aromatic compounds strongly depends on their structure.

209 The relative stability of butylcyclohexane and alkanes was studied elsewhere [34] and the  
210 relative scale of stability was:



212 This is in agreement with our results.

213 The relative stability of dodecylbenzene, 9-methylphenanthrene and  $nC_{25}$  was evaluated by  
214 Béhar et al. (2008), and it should be noted that our results concerning decylbenzene (instead  
215 of dodecylbenzene), methylnaphthalene (instead of 9-methylphenanthrene) and  $nC_{15}$  (instead  
216 of  $nC_{25}$ ) are in agreement with their results at 400°C.

#### 217 **4. Kinetic parameters at geological temperature (200°C)**

##### 218 *4.1 Alkanes and methylaromatics*

219 Alkanes and methylaromatics kinetic parameters seem to be almost constant between 400°C  
220 and 200°C (Table 2) because the free-radical reactions that control the thermal cracking  
221 remain the same. The apparent activation energy of alkanes does not depend on the chain  
222 length at 200°C on the contrary to 400°C and an average value of  $69.3 \pm 0.5$  kcal/mol can be  
223 proposed for any chain length. Nevertheless, the apparent rate constant  $k$  increases with the  
224 chain length and its evolution can be compared to the correlation of Watanabe et al. [56]  
225 validated at 200-700°C and 1-150 bar. Watanabe et al. [56] proposed a correlation that plots  
226  $k(nC_i)/k(nC_{16})$  as a function of the chain length  $i$  ( $3 \leq i \leq 32$ ) regardless of the temperature  
227 and the pressure. The correlation was validated up to  $i \approx 20$ . We have chosen arbitrarily  $C_{15}$  as  
228 a reference alkane instead of  $C_{16}$  in this study, so we plotted  $k(nC_i)/k(nC_{15})$  [26, 29, 30, 33] at  
229 400°C and 200°C as a function of  $i$  and we compared it to the correlation established for  
230  $k(nC_i)/k(nC_{16})$  (Figure 3). Agreement with the model [56] was obtained at up to  $i = 16$  for  
231 both temperatures.

##### 232 *4.2 Other hydrocarbons*

233 Table 2 shows that the behavior of decylbenzene and tetralin strongly depends on  
234 temperature. In the case of decylbenzene, the previous mechanistic study [28] revealed that  
235 the mechanism includes free-radical reactions as well as the molecular retroene reaction. The



236 importance of the retroene reaction increases as temperature decreases because its activation  
237 energy (54 kcal/mol) is much lower than the global cracking activation energy of the free-  
238 radical (70 kcal/mol). Indeed, at low temperature, the apparent activation energy almost  
239 corresponds to the activation energy of the retroene reaction.

240 The apparent activation energy of tetralin strongly decreases when the temperature decreases  
241 but so does the frequency factor. This is due to the bimolecular initiations whose kinetic  
242 parameters are much lower than those of unimolecular initiations. So their importance  
243 increases when temperature decreases and they strongly influence the apparent kinetic  
244 parameters, leading to lower frequency factors and lower activation energies.

#### 245 *4.3 Relative thermal stability at 200°C*

246 A relative scale of stability is established at 200°C in the same manner as previously for  
247 400°C, but the duration is set to 100 million years (Figure 4) which is characteristic of  
248 geological duration. The apparent kinetic parameters computed at 200°C are used. A different  
249 scale of stability is obtained at 200°C (Figure 4) in comparison to 400°C (Figure 2),  
250 essentially because the apparent activation energies are different for each compound, but also  
251 because of the modification of some kinetic parameters with temperature:

252  $DB < MN < C_{15} \approx TMB < \text{tetralin} \approx BCH \approx \text{toluene} < MCH$

253 One major difference between both scales of stability is the increase of the relative stability of  
254 methylcyclohexane when temperature decreases. Another major difference is the strong  
255 increase of the reactivity of decylbenzene with the decrease of temperature due to the retroene  
256 reaction as explained above.

257 The comparison between the relative scales of stability at low and high temperatures appears  
258 more difficult for the other compounds. In order to highlight the influence of the  
259 modifications of the apparent kinetic parameters, we also computed the relative scale of  
260 stability at 200°C, but with the kinetic parameters computed at 400°C, and we obtained:

261  $DB \approx MN < TMB < C_{15} < \text{toluene} < BCH \approx MCH \approx \text{tetralin}$

262 The major mistake that would be done by using at 200°C the kinetic parameters computed at  
263 400°C, would lie in the overestimation of the stability of tetralin and decylbenzene, in  
264 comparison to the other compounds.

## 265        5. Discussion

266        The apparent kinetic parameters of decylbenzene, methylcyclohexane and tetralin clearly  
267        depend on temperature. For all other compounds, the differences between the kinetic  
268        parameters computed at 400°C and 200°C could seem negligible. In order to quantify the  
269        consequences of these minor changes, we calculated the conversion at 200°C after 1, 10 or  
270        100 million years (depending on the reactivity of the hydrocarbon) by using the apparent  
271        kinetic parameters computed at 400°C and 200°C and assuming a first-order law (Table 4).  
272        We also calculated the half-life with both apparent kinetic parameters (Figures 5, 6 and 7).

### 273        5.1. Alkanes

274        Table 4 and Figure 5 show that simulating at 200°C the apparent kinetic parameters  
275        computed at 400°C would lead to a higher reactivity (increase of conversion and decrease of  
276        half-life) than using the kinetic parameters computed at 200°C, especially for short alkanes  
277        ( $nC_8$ ). The difference is less important when the length of alkanes increases and there is no  
278        difference for long alkanes ( $nC_{25}$ ).

279        Differences could be explained by on the weight of decomposition reactions via  $\beta$ -scission of  
280        the short radicals among all decomposition reactions. Indeed, the decomposition reactions are  
281        rate-limiting at high pressures and low temperatures [26] and they greatly influence the  
282        apparent activation energy. The kinetic parameters of the decomposition of short radicals  
283        depend more on temperature (higher activation energy) than the kinetic parameters of the  
284        other radicals [26]. Moreover, for short alkanes, the weight of decomposition of short radicals  
285        among all decomposition reactions is more important than for long-chain alkanes: short  
286        radicals are more concentrated during the thermal cracking of short alkanes than of long  
287        alkanes. That is why the apparent activation energy of short alkanes depends more on  
288        temperature than that of long alkanes. As a consequence, the errors made when simulating at  
289        200°C the kinetic parameters determined at 400°C lies on the fact that the alkane studied is  
290        short.

### 291        5.2. Monoaromatics and naphthenes

292        The computed conversions and half-lives are presented in Table 4, and Figure 6. All these  
293        compounds appear to be more stable at 200°C than at 400°C. With the apparent kinetic  
294        parameters computed, the conversions are lower while the half-lives higher at 200°C than at

295 400°C, although the apparent kinetic parameters do not seem to depend on temperature except  
296 for methylcyclohexane. In order to evaluate more precisely the temperature-dependence of the  
297 apparent first-order kinetic constant  $k$  of methylcyclohexane, we computed  $k(T)$  in the  
298 temperature range of 200-400°C and we plotted  $k(T)$  vs  $1/RT$  ( $R$ : ideal gas constant and  $T$ :  
299 temperature) in a logarithmic scale (markers in Figure 8). By considering the apparent kinetic  
300 parameters ( $A$  and  $E_a$ ) computed at 400°C, we extrapolated the Arrhenius law from 400°C to  
301 low temperature and the extrapolated rate constant is referred to  $k(400)$ . Then we plotted  
302  $k(400)$  vs  $1/RT$  (straight lines in a logarithmic scale in Figure 8). It clearly appears that the  
303 deviation between  $k(400)$  and  $k(T)$  increases when the temperature decreases and that  $k(T) <$   
304  $k(400)$ .

### 305 5.3. Decylbenzene and tetralin

306 The computed conversions and half-lives are presented in Table 4 and Figure 7. Unlike the  
307 other hydrocarbons studied, decylbenzene and tetralin appear to be more reactive at 200°C  
308 than at 400°C. When the apparent kinetic parameters are computed at 200°C the conversions  
309 are higher and the half-lives lower than at 400°C. In the case of decylbenzene, as explained  
310 above, this increasing reactivity is due to the retroene reaction whose activation energy is  
311 particularly low and whose importance increases when the temperature decreases [28]. The  
312 increasing reactivity of tetralin is probably related to the bimolecular initiations whose  
313 importance also increases when the temperature decreases because of their lower activation  
314 energy than that of unimolecular initiations [25].

315 The apparent kinetic parameters of decylbenzene and tetralin (Table 2) strongly depend on  
316 temperature. In order to evaluate precisely the temperature-dependence of the apparent first-  
317 order kinetic constant  $k$  of decylbenzene and tetralin, we computed  $k(T)$  in the temperature  
318 range of 200-400°C and we plotted  $\ln k(T)$  vs  $1/RT$  as well as  $\ln k(400)$  vs  $1/RT$  (Figure 8). It  
319 also clearly appears that the deviation between  $k(400)$  and  $k(T)$  increases when the  
320 temperature decreases and that  $k(T) > k(400)$ .

## 321 6. Conclusion

322 In this paper we computed the apparent kinetic parameters (frequency factor  $A$  and activation  
323 energy  $E_a$ ) of several pure hydrocarbons on the basis of detailed kinetic models previously  
324 constructed and validated. The purpose was to evaluate the effects of extrapolating data  
325 obtained in a range of temperatures to another. Indeed, thermal cracking reactions take place

326 at 150°C - 200°C for millions of years in geological reservoirs and so all laboratory data are  
327 obtained at about 400°C during some hours or days and then extrapolated.

328 The calculations were performed at 700 bar and at 400°C and then 200°C. The computation of  
329  $E_a$  is independent from the kinetic rate law with our method, but the computation of A is not:  
330 we postulated a first-order kinetic law which is of use in the vast majority of geochemistry  
331 publications. In the present study, most of the compounds exhibit apparent kinetic parameters  
332 that are slightly or even highly different between 400°C and 200°C. Extrapolating apparent  
333 kinetic parameters obtained at 400°C, even though they are only slightly different from those  
334 at 200°C, can lead to significant differences in terms of half-lives or conversion after several  
335 millions of years. All hydrocarbons studied here (alkanes, monoaromatics and naphthenes)  
336 appear to be more stable at 200°C than when extrapolating the Arrhenius law from 400°C to  
337 200°C, except decylbenzene and tetralin which appear much more reactive. To conclude, any  
338 extrapolation beyond the original range of temperatures used to determine the apparent kinetic  
339 parameters has to be considered very carefully, in the geochemistry field, but also in any other  
340 application fields of thermal cracking. We previously showed [26] that the same is true for the  
341 extrapolation beyond the original range of pressure because the importance of bimolecular vs  
342 unimolecular reactions can be completely modified.

343 To complete the discussion, two more points should be addressed: the choice of the kinetic  
344 rate law and the effects of co-reactants. In most studies, a first-order kinetic law is postulated,  
345 although it clearly appears that this choice is erroneous for some compounds; for instance, the  
346 order is close to  $\frac{1}{2}$  for pure alkanes [57] and 2 for pure methyl-naphthalene [35]. It would be  
347 interesting to lead a complete study on the best choice of kinetic order for each hydrocarbon  
348 and address the consequences of an erroneous choice on hydrocarbon stability. The second  
349 point is the effects of co-reactants on the apparent kinetic parameters. We showed on several  
350 systems ([29, 33, 58, 59]) that some co-reactants can completely change the stability of  
351 hydrocarbons and consequently the apparent kinetic parameters are also very likely to be  
352 modified. ~~Therefore, the studies on pure hydrocarbons have a limited range of application.~~ In  
353 order to construct bulk kinetic models of mixtures meant to represent oil fractions, a similar  
354 study as the one presented here should be performed with binary mixtures: correlations  
355 between the apparent kinetic parameters and the content of co-reactants could be obtained and  
356 then used in bulk kinetic models.

357

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513 **List of tables**

514 Table 1: Simulation conditions deduced from Peng Robinson equation of state.

	<b>Initial concentration (mol/m<sup>3</sup>) 700 bar - 400°C</b>	<b>Ideal gas pressure (bar) equivalent to 700 bar at 400°C</b>	<b>Initial concentration (mol/m<sup>3</sup>) 700 bar - 200°C</b>	<b>Ideal gas pressure (bar) equivalent to 700 bar at 200°C</b>
<i>n</i> C <sub>5</sub> H <sub>12</sub>	6948	389	8319	327
<i>n</i> C <sub>6</sub> H <sub>14</sub>	6280	351	7330	288
<i>n</i> C <sub>8</sub> H <sub>18</sub>	5208	291	5878	231
<i>n</i> C <sub>12</sub> H <sub>26</sub>	3602	202	3903	153
<i>n</i> C <sub>14</sub> H <sub>30</sub>	3038	170	3250	128
<i>n- or iso</i> C <sub>15</sub> H <sub>32</sub>	2797	156	2976	117
<i>n</i> C <sub>16</sub> H <sub>34</sub>	2579	144	2731	107
<i>n</i> C <sub>25</sub> H <sub>52</sub>	1376	77	1420	56
<b>Toluene</b>	7470	418	8700	342
<b>1-Methyl-naphthalene</b>	5919	331	6494	255
<b>Trimethyl-benzene</b>	6002	336	6756	266
<b>Decylbenzene</b>	3080	172	3281	129
<b>Tetralin</b>	5793	324	6404	252
<b>Methyl-cyclohexane</b>	6526	365	7486	294
<b>Butylcyclohexane</b>	4744	265	5223	205

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521 Table 2: Computed apparent kinetic parameters.

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Hydrocarbons	$A$ ( $s^{-1}$ )	$E_a$ (kcal/mol)	$A$ ( $s^{-1}$ )	$E_a$ (kcal/mol)
	400°C		200°C	
<b><i>n</i>C<sub>5</sub>H<sub>12</sub> [30]</b>	$1.1 \times 10^{16}$	67.2	$5.3 \times 10^{16}$	69.2
<b><i>n</i>C<sub>6</sub>H<sub>14</sub> [30]</b>	$3.6 \times 10^{16}$	67.3	$1.6 \times 10^{17}$	69.3
<b><i>n</i>C<sub>8</sub>H<sub>18</sub> [30]</b>	$7.7 \times 10^{16}$	67.0	$1.8 \times 10^{17}$	69.2
<b><i>n</i>C<sub>8</sub>H<sub>18</sub> [26]</b>	$6.6 \times 10^{15}$	64.0	$1.7 \times 10^{18}$	71.8
<b><i>n</i>C<sub>8</sub>H<sub>18</sub> [33]</b>	$6.5 \times 10^{16}$	67.6		
<b><i>n</i>C<sub>12</sub>H<sub>26</sub> [30]</b>	$1.8 \times 10^{17}$	68.6	$2.2 \times 10^{17}$	69.3
<b><i>n</i>C<sub>14</sub>H<sub>30</sub> [30]</b>	$1.6 \times 10^{17}$	68.2	$3.0 \times 10^{17}$	69.2
<b><i>n</i>C<sub>15</sub>H<sub>32</sub> [30]</b>	$1.8 \times 10^{17}$	67.8	$5.2 \times 10^{17}$	69.3
<b><i>iso</i>C<sub>15</sub>H<sub>32</sub> [30]</b>	$1.7 \times 10^{17}$	67.9	$5.3 \times 10^{17}$	69.3
<b><i>n</i>C<sub>16</sub>H<sub>34</sub> [29]</b>	$2.0 \times 10^{15}$	61.8		
<b><i>n</i>C<sub>25</sub>H<sub>52</sub> [30]</b>	$6.5 \times 10^{17}$	69.1	$7.3 \times 10^{17}$	69.2
<b>Toluene</b>	$2.9 \times 10^{10}$	54.4	$4.0 \times 10^{10}$	55.4
<b>1-Methylnaphtalene</b>	$1.3 \times 10^{14}$	59.3	$1.8 \times 10^{14}$	60.2
<b>Trimethylbenzene</b>	$8.8 \times 10^{11}$	56.1	$1.4 \times 10^{12}$	57.3
<b>Decylbenzene</b>	$1.2 \times 10^{18}$	67.0	$4.9 \times 10^{12}$	54.0
<b>Tetralin</b>	$7.4 \times 10^{15}$	67.5	$2.6 \times 10^{13}$	61.2
<b>Methylcyclohexane</b>	$1.9 \times 10^{14}$	63.7	$1.5 \times 10^{15}$	66.4
<b>Butylcyclohexane</b>	$4.7 \times 10^{16}$	68.4	$7.6 \times 10^{16}$	69.4

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527 Table 3: Literature data.

528

Compound	E (kcal/mol)	A (s <sup>-1</sup> )	Ref.	T-P conditions
<i>n</i> C <sub>9</sub> H <sub>20</sub>	63	2.0×10 <sup>14</sup>	Kunzru et al., [42]	650-750°C
<i>n</i> C <sub>10</sub> H <sub>22</sub>	60	3.6×10 <sup>21</sup>	Yu and Eser, [43]	400-450°C 21 bar
<i>n</i> C <sub>12</sub> H <sub>26</sub>	62	2.3×10 <sup>22</sup>	Yu and Eser, [43]	400-450°C 18 bar
<i>n</i> C <sub>12</sub> H <sub>26</sub>	65.2	3.7×10 <sup>16</sup>	Zhou and Crynes, [44]	250-440°C 92-103 bar
<i>n</i> C <sub>14</sub> H <sub>30</sub>	63	7.2×10 <sup>22</sup>	Yu and Eser, [43]	400-450°C 16 bar
<i>n</i> C <sub>14</sub> H <sub>30</sub>	67.6	7.2×10 <sup>18</sup>	Song et al., [45]	450°C 20-90 bar
<i>n</i> C <sub>16</sub> H <sub>34</sub>	59.2	3.1×10 <sup>13</sup>	Blouri et al., [46]	350-440°C 20 bar
<i>n</i> C <sub>16</sub> H <sub>34</sub>	59.6	3.1×10 <sup>14</sup>	Ford, [19]	330-420°C Liquid phase
<i>n</i> C <sub>16</sub> H <sub>34</sub>	61.2	1.5×10 <sup>15</sup>	Khorasheh and Gray, [20]	380-450°C 139 bar
<i>n</i> C <sub>16</sub> H <sub>34</sub>	74	3.0×10 <sup>19</sup>	Jackson et al., [15]	300-370°C 150-600 bar
<i>n</i> C <sub>25</sub> H <sub>52</sub>	68.2	6.1×10 <sup>17</sup>	Behar and Vandenbroucke, [14]	325-425°C 120-800 bar

<b>Toluene</b>	66.0	$3.3 \times 10^{13}$	Domke et al., [47]	540-700°C 1 bar
<b>Toluene</b>	70.1	$4.8 \times 10^{14}$	Clark and Price, [48]	?
<b>1-methylnaphthalene</b>	47.4	$7.9 \times 10^9$	Leininger et al., [24]	380-450°C 100 bar
<b>Trimethylbenzene</b>	58.0	$1.1 \times 10^{12}$	Fusetti et al., [31]	395-475°C 100 bar
<b>Ethylbenzene</b>	62.3	$4.7 \times 10^{13}$	Domke et al., [47]	540-700°C 1 bar
<b>Butylbenzene</b>	65	$3.0 \times 10^{14}$	Leigh and Szwarc, [49]	?
<b>Butylbenzene</b>	52.9	$1.1 \times 10^{12}$	Freund and Olmstead, [50]	505-650°C
<b>Butylbenzene</b>	57	$4.0 \times 10^{17}$	Yu and Eser, [51]	400-450°C 20-85 bar
<b>Dodecylbenzene</b>	53.3	$1.3 \times 10^{13}$	Behar et al., [17]	325-425°C 140 bar
<b>Pentadecylbenzene</b>	55.5	$1.1 \times 10^{14}$	Savage and Klein, [21]	375-450°C 1-9.7 bar
<b>9- Methylphenanthrene</b>	49.0	$4.5 \times 10^{10}$	Behar et al., [13]	375-450°C 120 bar
<b>Dibenzothiophene</b>	59.0	$1.9 \times 10^{11}$	Dartiguelongue et al., [18]	375-500°C 100 bar
<b>2-Dodecyl-9,10- dihydrophenanthrene</b>	$54.5 \pm 9.1$	$4.0 \times 10^{13}$	Savage and Baxter, [23]	375-450°C
<b>Tetralin</b>	58.0	$3.5 \times 10^{12}$	Yu and Eser, 1998	425-475°C

				23-75 bar
<b>2-Ethyltetralin</b>	53.5	$5.0 \times 10^{12}$	Savage and Klein, 1988	375-450°C
<b>Propylcyclohexane</b>	67.6	$2.6 \times 10^{16}$	Widegren and Bruno, [53]	350-450°C 3500 bar
<b>Butylcyclohexane</b>	65.0	$3.5 \times 10^{15}$	Yu and Eser, [54]	425-475°C 17-86 bar
<b>Butylcyclohexane</b>	69.0	$6.7 \times 10^{16}$	Rakatoalimanana et al., [52]	300-425°C 100 bar

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544 Table 4: Temperature-dependence of kinetic parameters of some hydrocarbons: influence on  
 545 conversion at 200°C.

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	<i>n</i> C <sub>8</sub> H <sub>18</sub>		<i>n</i> C <sub>15</sub> H <sub>32</sub>	
	First-order law using A and Ea calculated at 400°C	First-order law using A and Ea calculated at 200°C	First-order law using A and Ea calculated at 400°C	First-order law using A and Ea calculated at 200°C
<b>Conversion after 10 million years</b>	91%	45%	93%	78%
	<i>n</i> C <sub>25</sub> H <sub>52</sub>		Toluene	
	First-order law using A and Ea calculated at 400°C	First-order law using A and Ea calculated at 200°C	First-order law using A and Ea calculated at 400°C	First-order law using A and Ea calculated at 200°C
<b>Conversion after 10 million years</b>	91%	91%	49%	27%
	1-Methylnaphtalene		Trimethylbenzene	
	First-order law using A and Ea calculated at 400°C	First-order law using A and Ea calculated at 200°C	First-order law using A and Ea calculated at 400°C	First-order law using A and Ea calculated at 200°C
<b>Conversion after 10 million years</b>	100%	100%	96%	77%
	DB		Tetralin	
	First-order law using A and Ea calculated at 400°C	First-order law using A and Ea calculated at 200°C	First-order law using A and Ea calculated at 400°C	First-order law using A and Ea calculated at 200°C
<b>Conversion after 100 million years</b>	100% (98% after 1 million years)	100% (100% after 1 million years)	78%	99%
	MCH		BCH	
	First-order law using A and Ea calculated at	First-order law using A and Ea calculated at	First-order law using A and Ea calculated at	First-order law using A and Ea calculated at

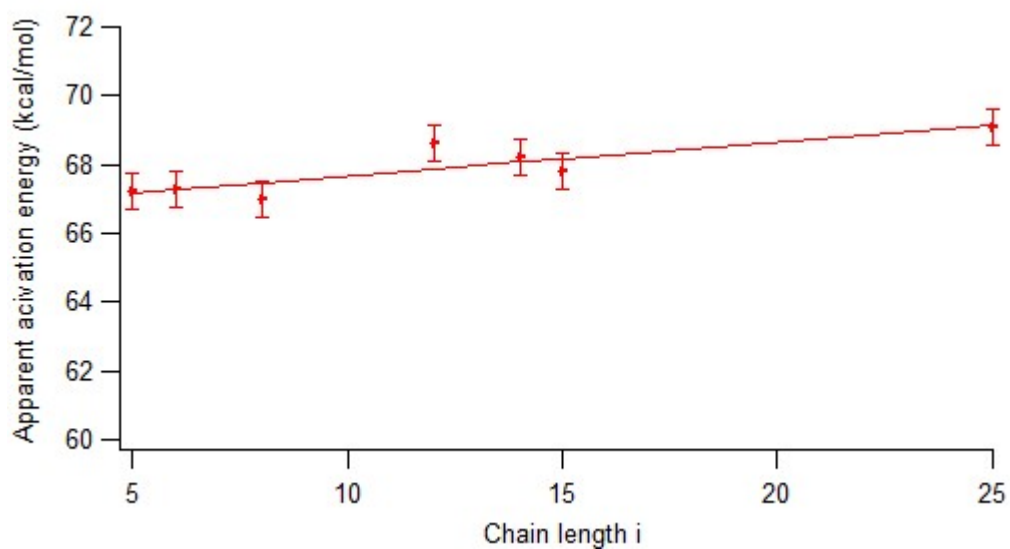


	400°C	200°C	400°C	200°C
<b>Conversion after 100 million years</b>	89%	63%	97%	87%

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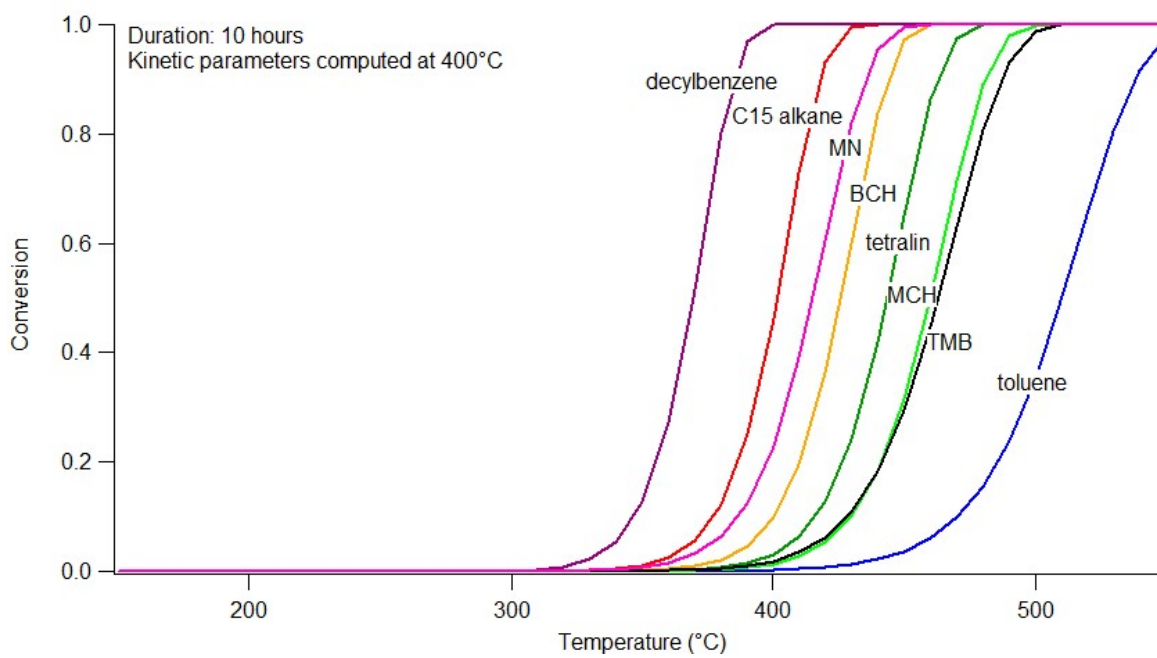
549 **Figure captions**



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551 Figure 1: Apparent activation energy computed at 400°C as a function of the chain length  $i$   
552 (linear regression:  $E_a$  (kcal/mol) =  $0.10 \times i + 66.7$ ).

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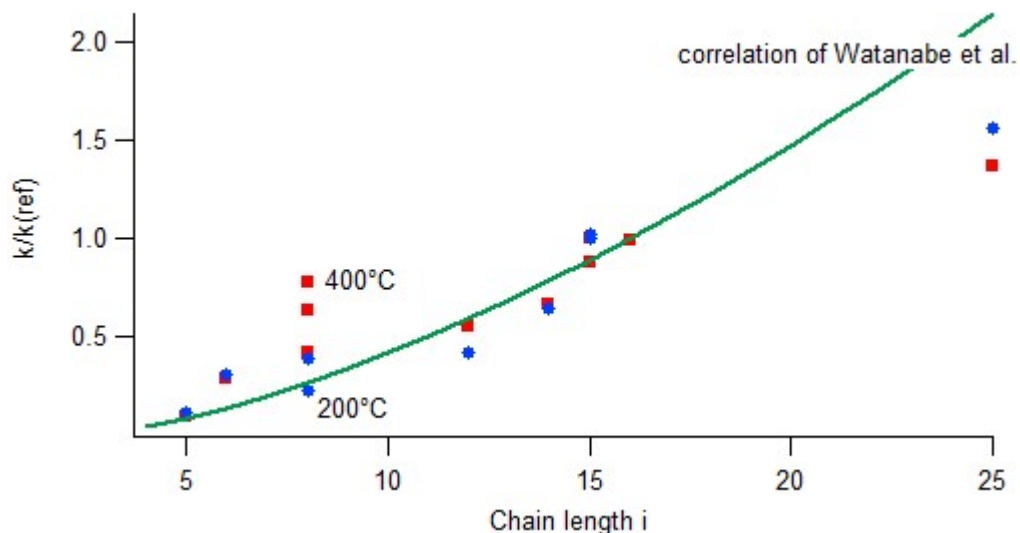
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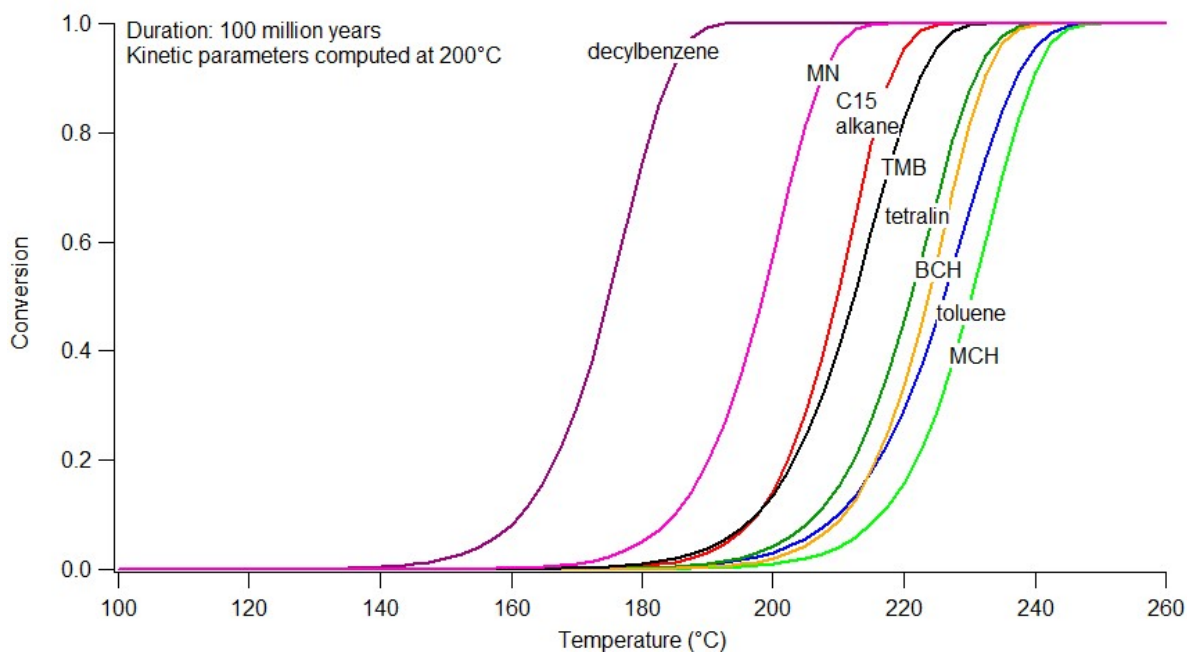
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560 Figure 2: Relative stability of the hydrocarbons determined by using the apparent kinetic  
561 parameters calculated at 400°C (pyrolysis duration set to 10 h).



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563 Figure 3: Comparison of the normalized rate constants [26, 29, 30, 33] (C<sub>15</sub> as the reference  
564 compound) as a function of the chain length to the correlation of Watanabe et al. [55] (square  
565 markers: 400°C and circle markers: 200°C).

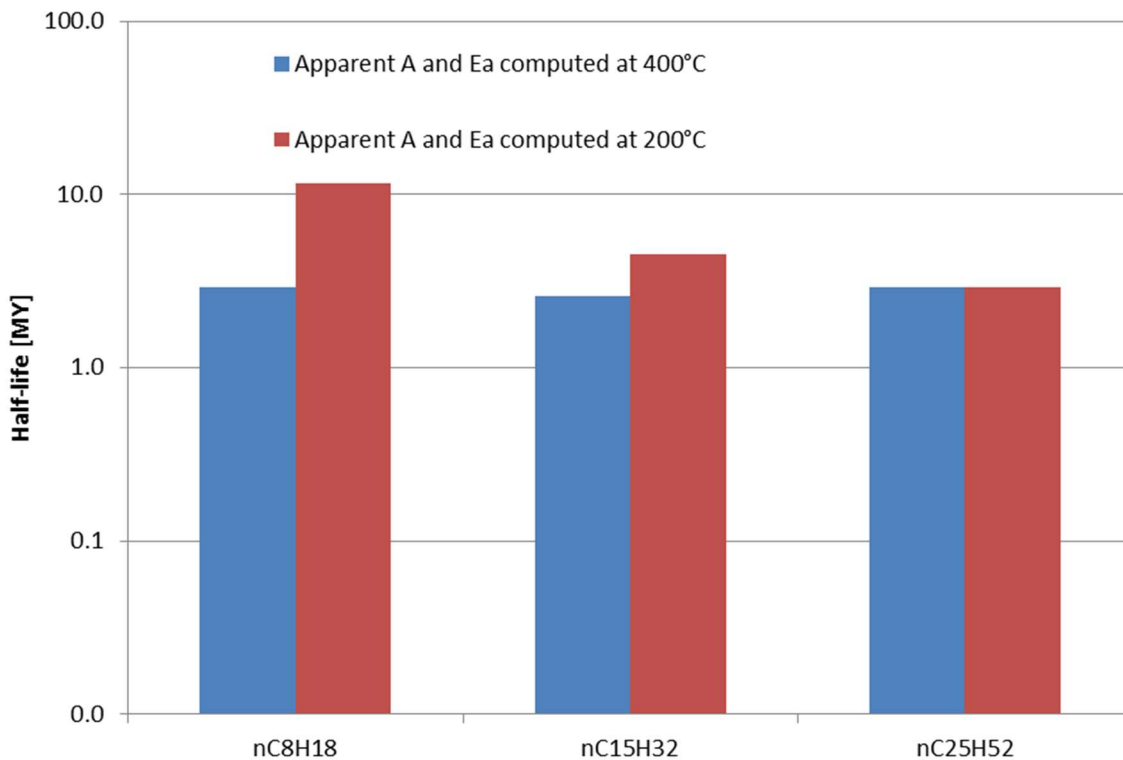


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567 Figure 4: Relative stability of the hydrocarbons determined by using the apparent kinetic  
568 parameters calculated at 200°C (pyrolysis duration set to 100 million years).

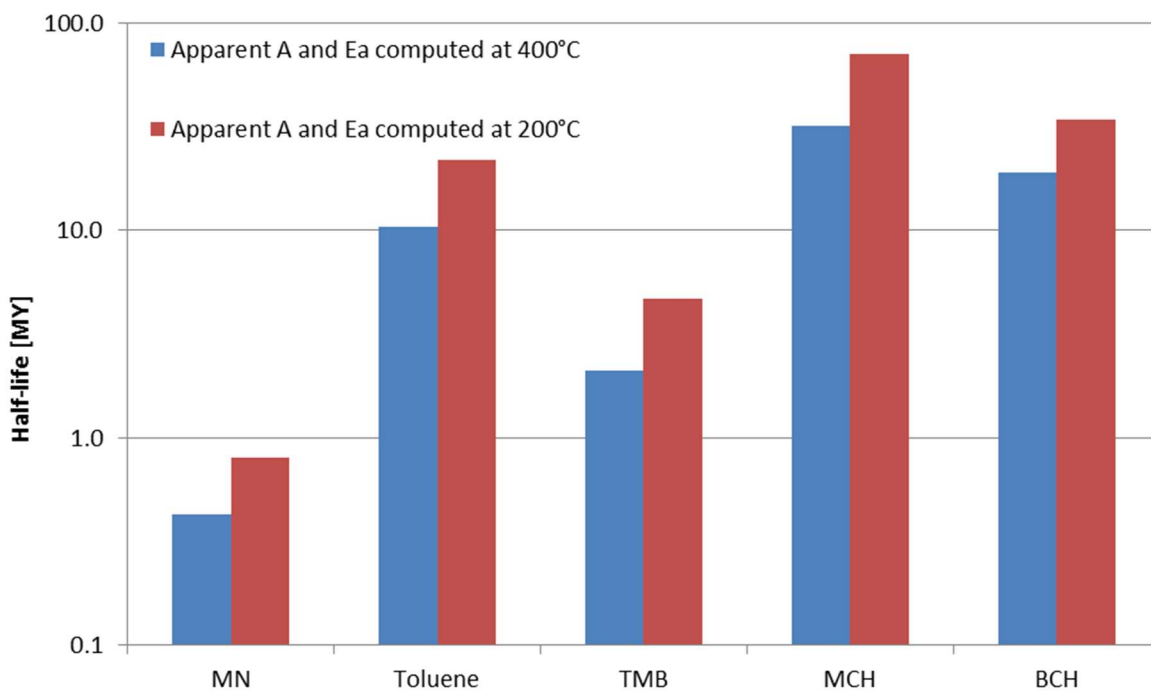
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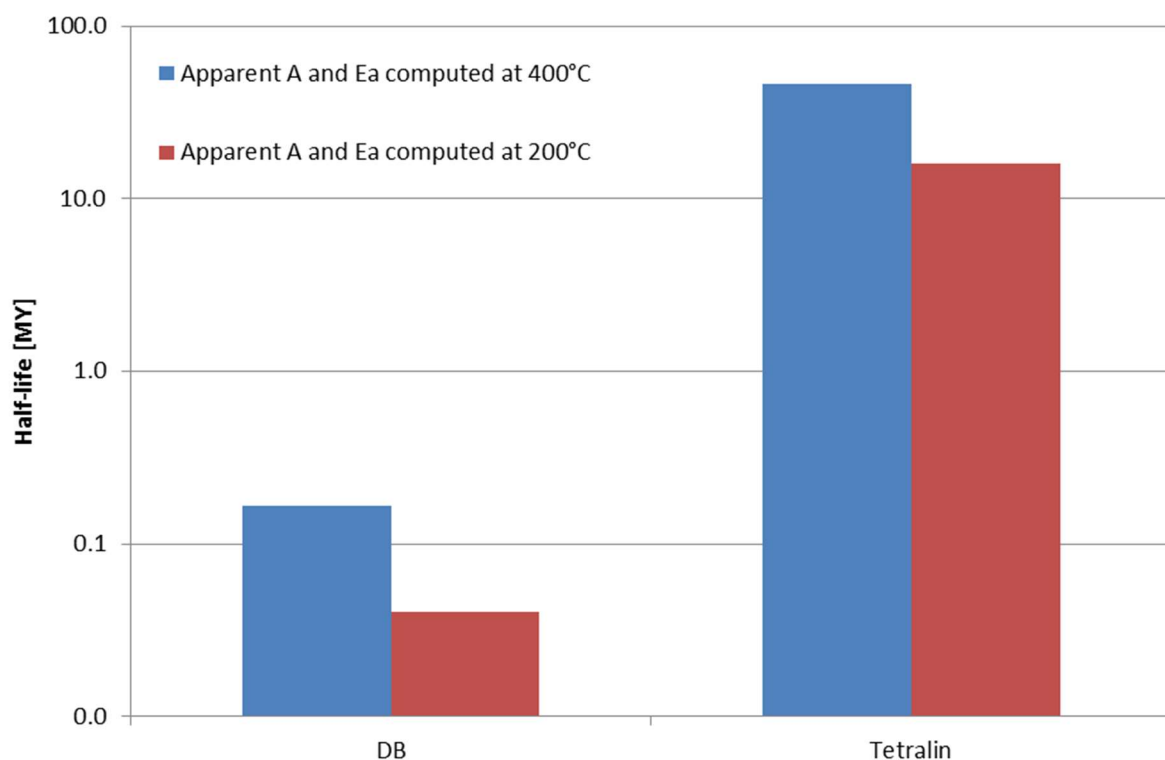
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572 Figure 5: Temperature-dependence of kinetic parameters of some alkanes: influence on half-  
 573 life at 200°C.



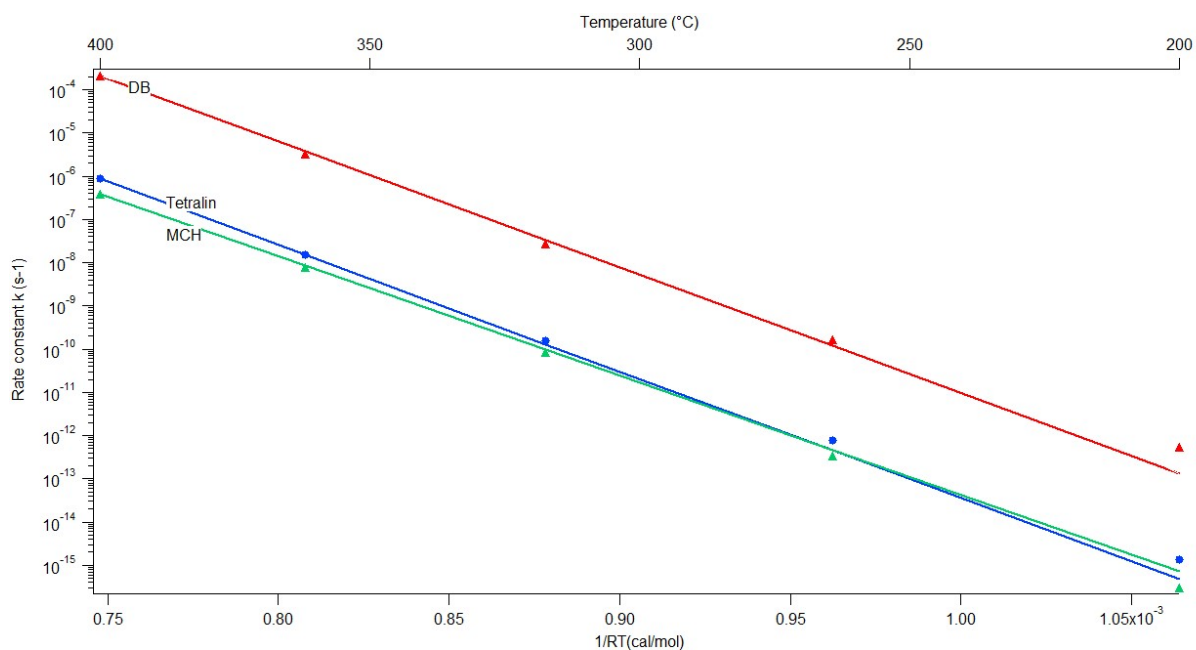
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575 Figure 6: Temperature-dependence of kinetic parameters of methylaromatics and naphthenes:  
 576 influence on half-life at 200°C.



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578 Figure 7: Temperature-dependence of kinetic parameters of decylbenzene and tetralin:  
579 influence on half-life at 200°C.



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581 Figure 8: Computed evolution of the rate constants (decylbenzene, methylcyclohexane and  
582 tetralin) as a function of temperature compared with the extrapolation of the Arrhenius law  
583 from 400°C.

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