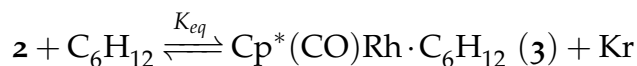1. The proposed mechanism for C-H bond activation of cyclohexane via photoin-
   duced reaction of Cp*Rh(CO)$_2$ (**1**) in liquid Kr solution is

$$\mathbf{1} \xrightarrow[\text{Liq}\cdot\text{Kr}]{h\nu} \text{Cp}^*(\text{CO})\text{Rh}\cdot\text{Kr } (\mathbf{2}) + \text{CO}$$

$$\mathbf{2} + \text{C}_6\text{H}_{12} \xrightleftharpoons{K_{eq}} \text{Cp}^*(\text{CO})\text{Rh}\cdot\text{C}_6\text{H}_{12} \ (\mathbf{3}) + \text{Kr}$$

$$\mathbf{3} \xrightarrow{k_2} \text{Cp}^*(\text{CO})\text{Rh}(\text{C}_6\text{H}_{11})(\text{H})$$

   (a) According to the authors, the mechanism can be analyzed using mass
       balance and a modified Michaelis-Menten model, assuming rapid pre-
       equilibrium, to give the following expression for the rate

$$\text{rate} = k_{\text{obs}}([\mathbf{2}] + [\mathbf{3}]) = \frac{k_2[\text{C}_6\text{H}_{12}]}{[\text{C}_6\text{H}_{12}] + \frac{[\text{Kr}]}{K_{eq}}}([\mathbf{2}] + [\mathbf{3}]).$$

       Verify that this expression is correct.

   (b) The authors confirmed their mechanism from the observed dependence of
       k_{obs} with $[\text{C}_6\text{H}_{12}]$. Plot the expected form of the graph for the variation k_{obs}
       with $[\text{C}_6\text{H}_{12}]$.

   (c) A second confirmation was from the ratio $k_{\text{obs}}(\text{C}_6\text{H}_{12})/k_{\text{obs}}(\text{C}_6\text{D}_{12})$ when
       the alkane concentration was high. Qualitatively analyze (a figure, maybe)
       how and why this ratio would provide confirmation for the mechanism.
       Although only a qualitative analysis is expected, you might find it useful to
       know that the $\text{C}-\text{H}$ vibrational frequency, ω_{CH}, is $3000\,\text{cm}^{-1}$.

   Answer: (a)

$$\text{Rate} = \frac{\text{d}[\text{Cp}^*(\text{CO})\text{Rh}(\text{C}_6\text{H}_{11})(\text{H})]}{\text{d}t} = k_2[\mathbf{3}]$$

$$K_{eq} = \frac{[\mathbf{3}][\text{Kr}]}{[\mathbf{2}][\text{C}_6\text{H}_{12}]}$$

$$[\mathbf{2}] = \frac{[\mathbf{3}][\text{Kr}]}{K_{eq}[\text{C}_6\text{H}_{12}]}$$

$$[\mathbf{2}] + [\mathbf{3}] = [\mathbf{3}]\left(1 + \frac{[\text{Kr}]}{K_{eq}[\text{C}_6\text{H}_{12}]}\right)$$

$$[\mathbf{3}] = \frac{1}{1 + \frac{[\text{Kr}]}{K_{eq}[\text{C}_6\text{H}_{12}]}}([\mathbf{2}] + [\mathbf{3}])$$

$$k_2[\mathbf{3}] = \frac{k_2[\text{C}_6\text{H}_{12}]}{[\text{C}_6\text{H}_{12}] + \frac{[\text{Kr}]}{K_{eq}}}([\mathbf{2}] + [\mathbf{3}]),$$

       which is the desired result.

(b) Compare the k_{obs} expression with the Michaelis-Menten expression $\frac{k_2[S][E]_0}{K_m+[S]}$. When $[C_6H_{12}]$ is small, k_{obs} varies linearly with $[C_6H_{12}]$. When $[C_6H_{12}]$ is high, k_{obs} becomes independent of $[C_6H_{12}]$, just like in the Michaelis-Menten case. In the figure below, replace "Reaction rate" with k_{obs} and "Substrate concentration" with $[C_6H_{12}]$.
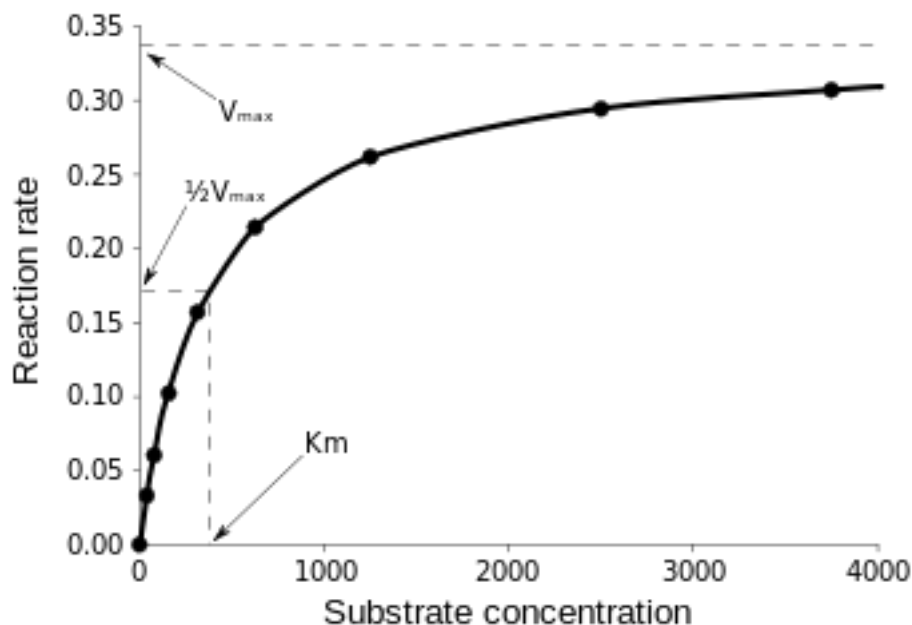
(c) When $[C_6H_{12}]$ is high, $k_{obs} = k_2$ implying that k_2 is rate determining. In this reaction the C−H bond is broken. Replacing C_6H_{12} with C_6D_{12} will show a large kinetic isotope effect. Draw a Lennard-Jones potential curve and show the C−H ground vibrational state with energy (ZPE) $1500\,\text{cm}^{-1}$ and C−D lower than that. The activation energy of C−D bond breakage is higher by this difference in the ZPE and hence the rate is slower.
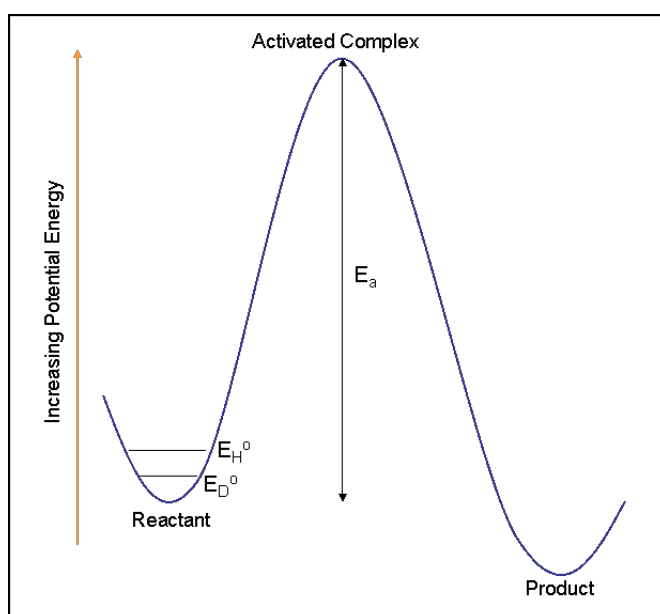
2. The unfolding and folding kinetics of the leucine zipper peptide GCN4-p1,

denoted N, is consistent with a model $N_2 \underset{k_f}{\overset{k_u}{\rightleftharpoons}} 2\,N$. The folding and unfolding rate constants are $4.2 \times 10^5 \, \text{M}^{-1}\text{s}^{-1}$ and $3.3 \times 10^{-3}\,\text{s}^{-1}$, respectively. A $6.0\,\mu\text{M}$ solution of N was prepared in a buffer. Addition of a compound ("denaturant") or dilution of the solution "instantaneously" led to the unfolding or folding, respectively, of the peptide which was followed as a function of time. Calculate the relaxation time observed in such an experiment.

Answer: This question is not as straight-forward as you imagined. You are given the initial concentration of the monomer, while what you need $\tau = 1/(k_u + 4k_f[N]_{eq})$ is the equilibrium concentration of the monomer. Gotcha! The expression that is required is:

$$\frac{1}{\tau^2} = k_u^2 + 8k_u k_f[N]_0$$

How does one get this expression from $\tau = 1/(k_u + 4k_f[N]_{eq})$ which you either derived or had memorized?

By mass balance, we have

$$[N]_0 = 2[N_2]_{eq} + [N]_{eq}$$

$$[N]_0 = [N]_{eq} + \frac{2[N]_{eq}^2 k_f}{k_u}$$

$$k_u[N]_0 = k_u[N]_{eq} + 2[N]_{eq}^2 k_f$$

$$\frac{1}{\tau^2} = (k_u + 4k_f[N]_{eq})^2$$

$$\frac{1}{\tau^2} = 8k_f\left(2k_f[N]_{eq}^2 + k_u[N]_{eq}\right) + k_u^2,$$
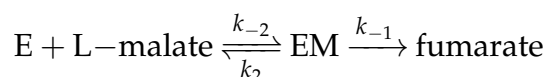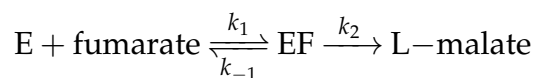
which yields the earlier written formula for $1/\tau^2$.

Substituting the measured values of the rate constants we obtain

$$\tau = 3.88\,\text{s}$$

3. The enzyme fumarase catalyzes the reversible hydration of fumarate to L-malate, $^-\text{OOCCH}=\text{CHCOO}^- + \text{H}_2\text{O} \rightleftharpoons {}^-\text{OOCCH}_2\text{CH(OH)COO}^-$. When fumarate is in excess the Michaelis constant and turnover number are $2.6\,\mu\text{M}$ and $1.0 \times 10^4\,\text{min}^{-1}$, while when L-malate is in excess they are $8.3\,\mu\text{M}$ and $7.9 \times 10^4\,\text{min}^{-1}$ respectively. Determine ALL the rate constants in the Michaelis-Menten mechanism for this reaction.
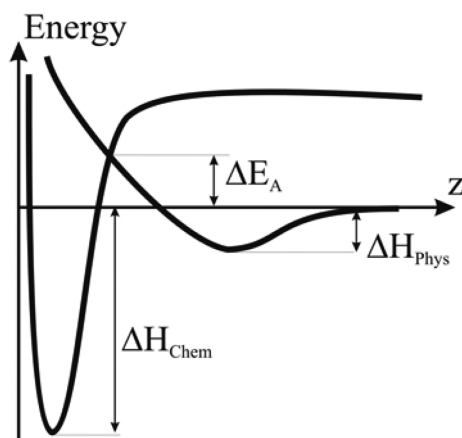
Answer:

$$\text{E} + \text{fumarate} \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} \text{EF} \overset{k_2}{\longrightarrow} \text{L}-\text{malate}$$

$$\text{E} + \text{L}-\text{malate} \underset{k_2}{\overset{k_{-2}}{\rightleftharpoons}} \text{EM} \overset{k_{-1}}{\longrightarrow} \text{fumarate}$$

$$K_m(\text{fumarate}) = \frac{k_{-1} + k_2}{k_1}$$

$$K_m(\text{L}-\text{malate}) = \frac{k_{-1}+k_2}{k_{-2}}$$

Hence, all the rate constants are $k_2 = 1.0 \times 10^4\,\text{min}^{-1}$, $k_{-1} = 7.9 \times 10^4\,\text{min}^{-1}$, $k_1 = \frac{k_{-1}+k_2}{K_m(\text{fumarate})} = 3.4 \times 10^{10}\,\text{M}^{-1}\text{min}^{-1}$, and $k_{-2} = \frac{k_{-1}+k_2}{K_m(\text{L}-\text{malate})} = 1.07 \times 10^{10}\,\text{M}^{-1}\text{min}^{-1}$.

4. The enthalpy of adsorption for H_2 adsorbed on a copper surface is $-54.4\,\text{kJ}\,\text{mol}^{-1}$. The activation energy for going from the physisorbed state to the chemisorbed state is $29.3\,\text{kJ}\,\text{mol}^{-1}$, and the curve crossing between these two potentials occurs at a potential of $21\,\text{kJ}\,\text{mol}^{-1}$. Draw a schematic representation of the variation of the potential of H_2 gas with distance from a copper surface. (Statutory Warning: Unlabelled figure kills!)

Answer:

In our case, $\Delta H_{\text{Chem}} = -54.4\,\text{kJ}\,\text{mol}^{-1}$, $\Delta E_A = 21\,\text{kJ}\,\text{mol}^{-1}$, and $\Delta H_{\text{Phys}} = -(E_a - \Delta E_A) = -(29.3\,\text{kJ}\,\text{mol}^{-1} - 21\,\text{kJ}\,\text{mol}^{-1}) = -8.3\,\text{kJ}\,\text{mol}^{-1}$

5. The variation of the rate of hydrogenation of ethylene on solid alumina catalyst with pressure of hydrogen (P_{H_2}) and temperature (T) are given below. The mechanism proposed for this reaction is $H_2(g) + C_2H_4(s) \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} C_2H_5(s) + H(s) \overset{k_2}{\longrightarrow} C_2H_6(g)$, where X(s) indicates adsorbed X.

(a) With the help of the steady state approximation, verify that the mechanism explains the observed dependence of the rate on pressure of hydrogen.

(b) Sketch the predicted dependence of the rate on the pressure of ethylene.

(c) If all the rate constants show an Arrhenius temperature dependence, rationalize the temperature dependence of the rate.
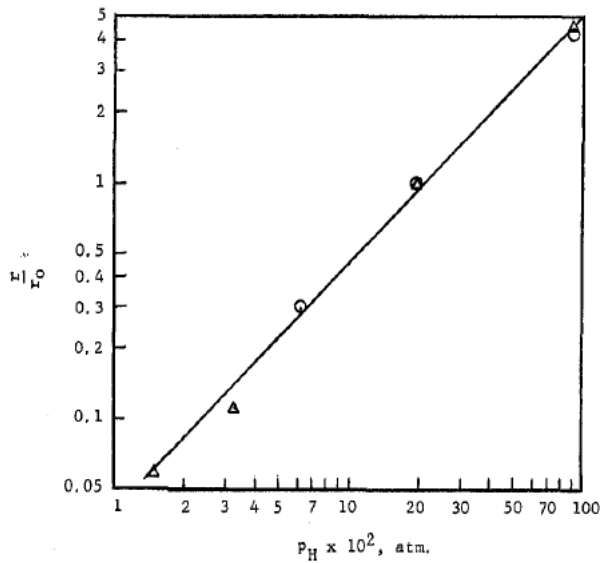
Answer: (a)

$$\text{Rate} = k_2\theta_{C_2H_5}\theta_H$$

$$\frac{d[C_2H_5(s)]}{dt} = k_1 P_{H_2}\theta_{C_2H_4} - k_{-1}\theta_{C_2H_5}\theta_H - k_2\theta_{C_2H_5}\theta_H = 0$$

Solving for $\theta_{C_2H_5}$

$$\text{Rate} = \frac{k_1 k_2}{k_{-1}+k_2}P_{H_2}\theta_{C_2H_4}.$$

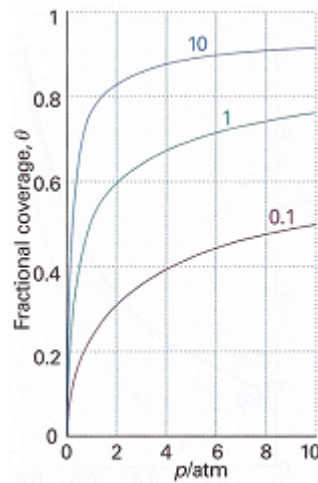This explains the linear dependence on the pressure.

(a) Rate dependence on P_{H_2}
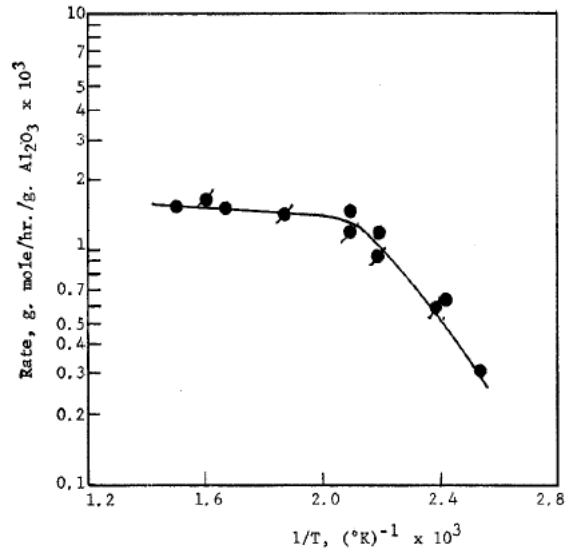
(b) Rate dependence on T

Figure 2: Hydogenation of ethylene on alumina

(b) Assuming a Langmuir-type adsorption for ethylene, we would get the typical pressure dependence of fractional coverage for ethylene.

(c) If k_2 is large compared to k_{-1} at low temperatures, the rate law reduces to the expression $k_1 p_{H_2} \theta_{C_2 H_4}$ and the over-all activation energy E is the temperature dependence of $k1$. However, if k_{-1} increases faster with temperature than k_2 (i. e., $E_{-1} > E_2$), then at a sufficiently high temperature k_{-1} will become large compared to k_2 and the rate law will reduce to $(k_1 k2/k_{-1} p_{H_2} \theta_{C_2 H_4})$. Under these conditions the over-all activation energy E becomes $E = E_1 + E_2 - E_{-1}$. By hypothesis E_{-1} is larger than E_2, and hence at high temperature, the activation energy E will be lower than E_1, resulting in a decreased dependence of the rate at high temperatures, as is observed.