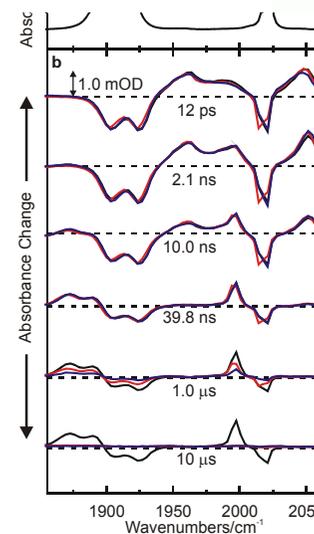
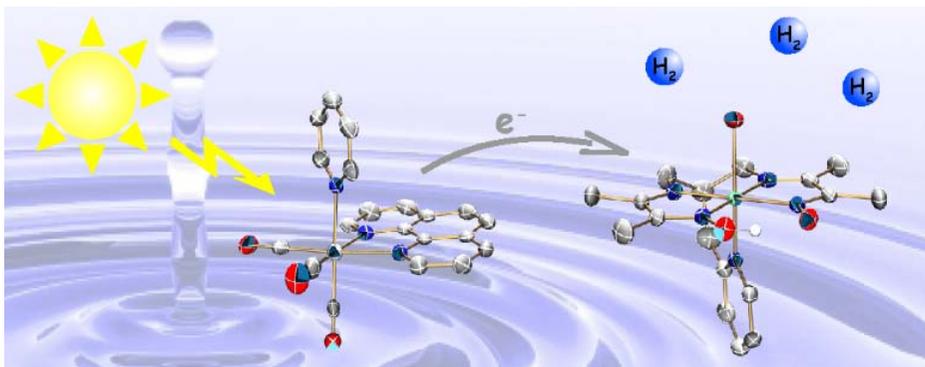
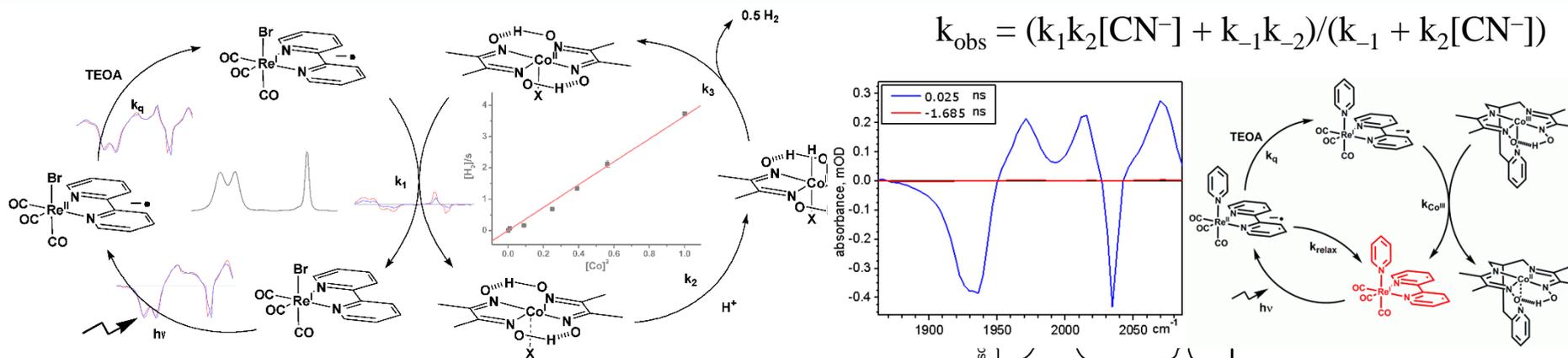


Reaction Mechanisms CHE 323

Christina Nevado, Roger Alberto FS 2018



$$\frac{dH_2}{dt} = k_3 K_2^2 \left(\frac{[HA^+]}{[A]} \right)^2 c^2 [Co]_{tot}^2 = k_{obs} [Co]_{tot}^2$$

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Literature

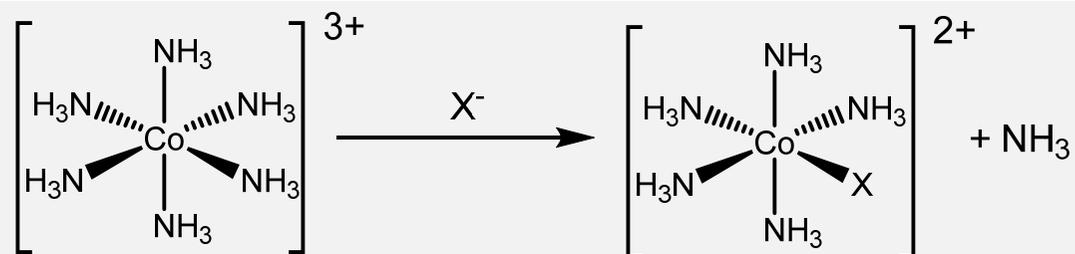
- **James H. Espenson**
Chemical Kinetics and Reaction Mechanisms, McGraw Hill 2002
or newer
- **M.L. Tobe, J. Burgess**
Inorganic Reaction Mechanisms, Addison Wesley Longman,
Essex UK, 2000
- **Robert B. Jordan**
Reaction Mechanisms of Inorganic and Organometallic
Systems, Oxford University Press, 2007 (Third Edition)
- **Jim D. Atwood**
Inorganic and Organometallic Reaction Mechanisms, VCH,
1997 (Second Edition)

Motivation

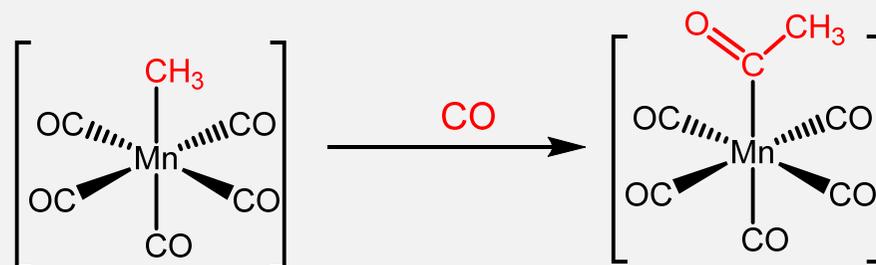
Why do we want to know a reaction mechanism ?

What can we do to deduce a mechanism ?

A few examples or "How does it work ?"



dissociative – associative – pH (in(dependent) X- dependent ?
does it matter at all ?

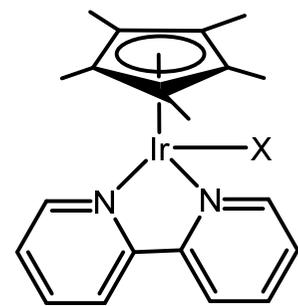


CO insertion or methyl migration
or does it matter at all ?

Motivation

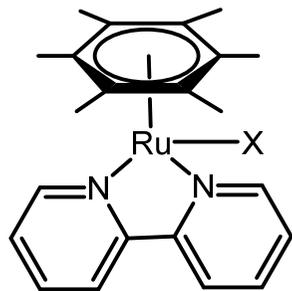
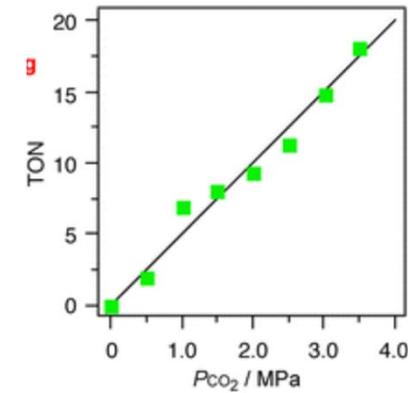
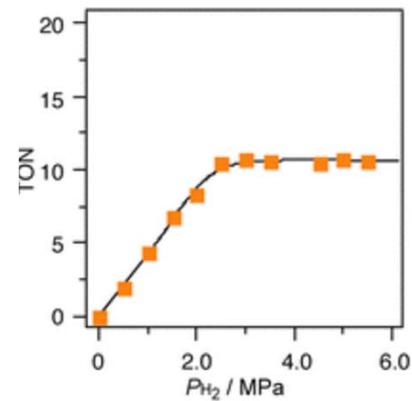
Why do we want to know a reaction mechanism ?

What can we do to deduce a mechanism ?

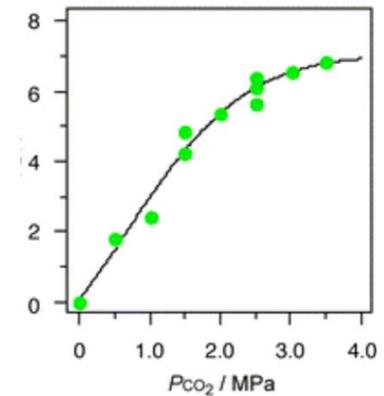
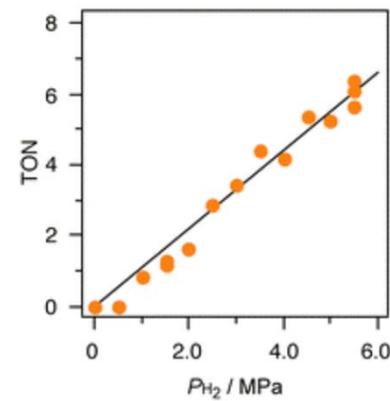


HCOOH

H₂ / CO₂



HCOOH



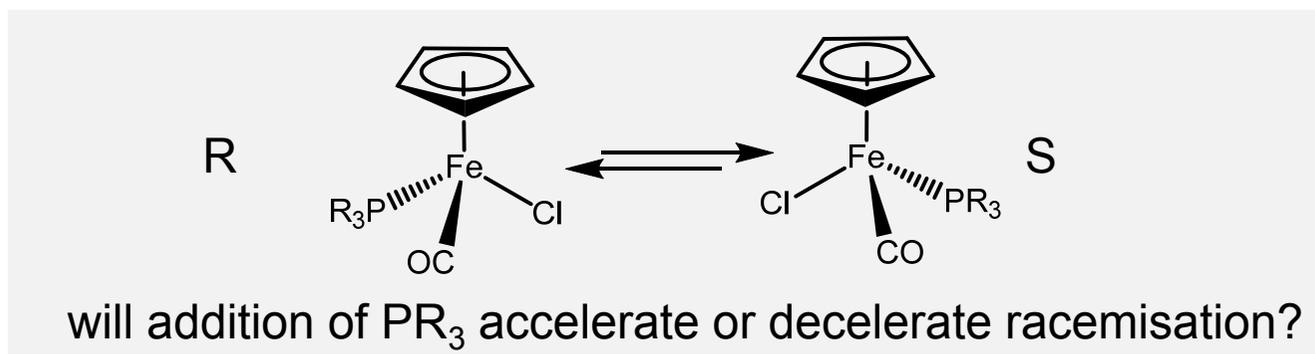
H₂ / CO₂: The more the better ?

Motivation

Why do we want to know a reaction mechanism ?

What can we do to deduce a mechanism ?

A few examples or "How does it work ?"



is an $\text{S}_{\text{N}}1$ reaction always 1st order in R-LG ?

why does Photosynthesis work with 10% sunlight at its maximum speed ?

are there reactions which become slower with increasing temperature?

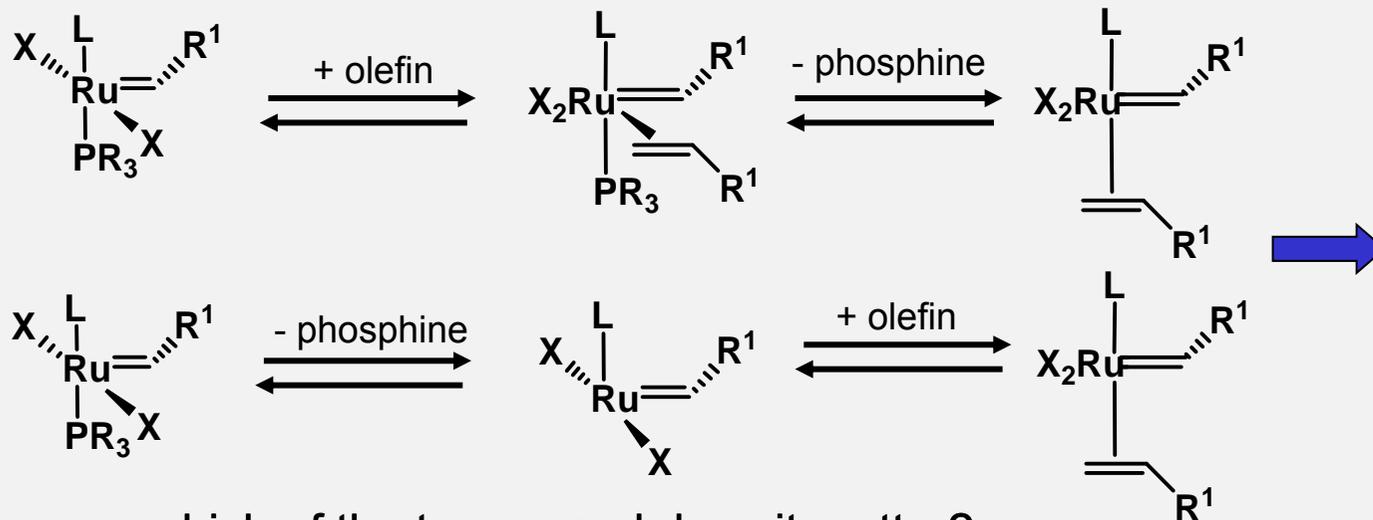
Motivation

The only way to understand and confirm a reaction mechanism is through deduction of its rate and the influence the different parameters have on it

$$\text{rate} = k \cdot [\text{A}]$$

the **rate "constant" k** in an overall reaction is often not constant, but
Comprises the information needed to understand a mechanism

It is easy to draw a mechanism but it is difficult to prove it



which of the two ... and does it matter?

Motivation

A study into the kinetics of a chemical reaction is usually carried out with one or both of two main goals in mind:

1. Analysis of the **sequence of elementary steps** giving rise to the overall reaction. i.e. the *reaction mechanism*.
2. Determination of the **absolute rate** of the reaction and/or its individual elementary steps.

The aim of the 1st part is to show you how these two goals may be achieved

1.1 Some Basic Concepts

Reaction Mechanism

A reaction is composed of several reaction steps (elementary reaction) the entity of which is called the reaction mechanism

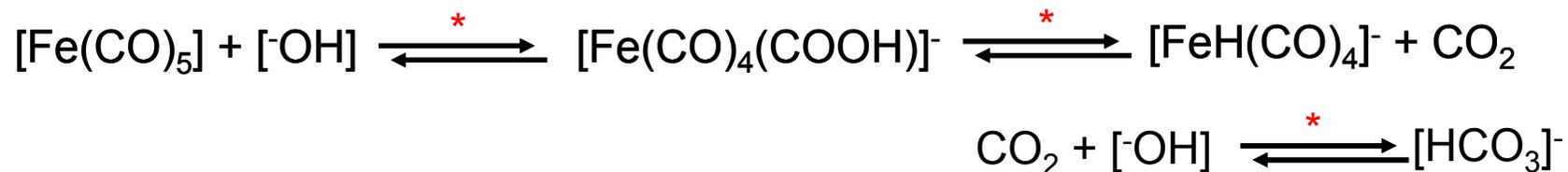


Stoichiometric equation, certainly not a single step reaction mechanism!

Elementary steps or reaction*

An elementary step in a reaction mechanism leads from A → B over one single transition state with A, B being intermediates, educt or product.

A reaction mechanism is composed of one to many elementary steps or reactions



1.1 Some Basic Concepts

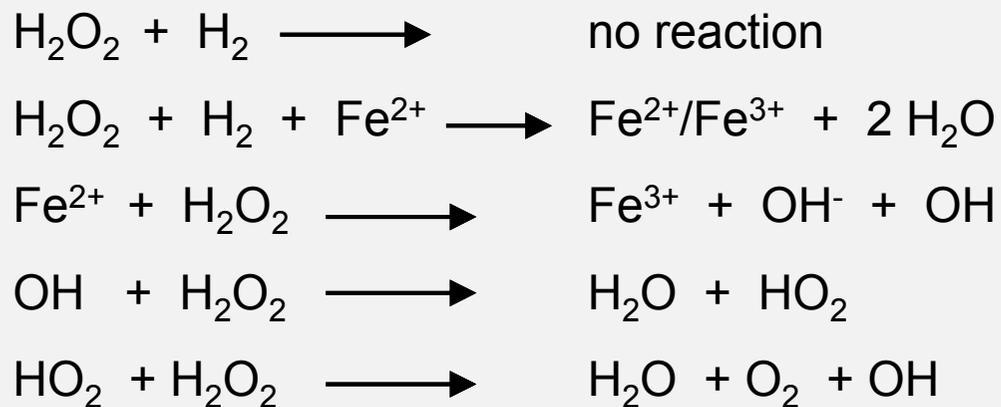
Elementary step: smallest unit of a chemical reaction on the molecular level

Reaction mechanism: Composition of elementary reactions

Relevance of elementary reactions

Kinetic schemes → rate laws

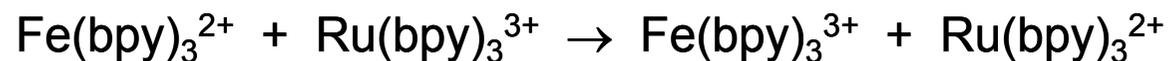
→ **Understanding** of mechanisms on a molecular level



Some Basic Concepts

Elementary step: smallest unit of a chemical reaction on the molecular level

1) Electron Transfer Reaction (without bond break or bond formation)



2) Bond formation of bond breaking: $\text{H}^+ + \text{H}^- / \text{H}-\text{H}$

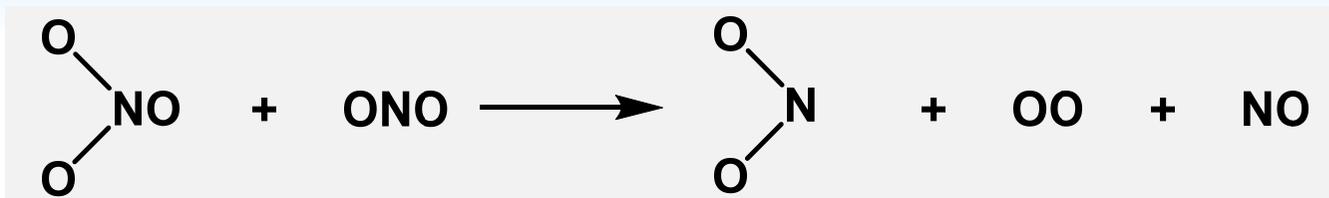


3) Simultaneous bond breaking and bond making



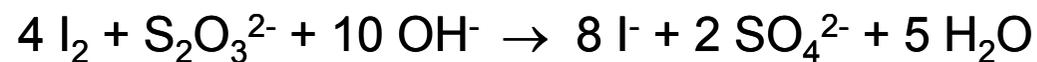
Some Basic Concepts

4) Simultaneous breaking and making of two bonds (rare)



Characteristics of elementary reactions:

Molecularity rarely exceeds 3 (see also below)



no an elementary reaction, multi-step

➔ **Structural and electronic changes** in an elementary step should be small

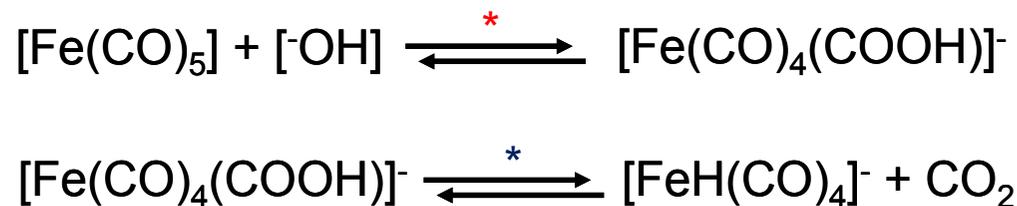
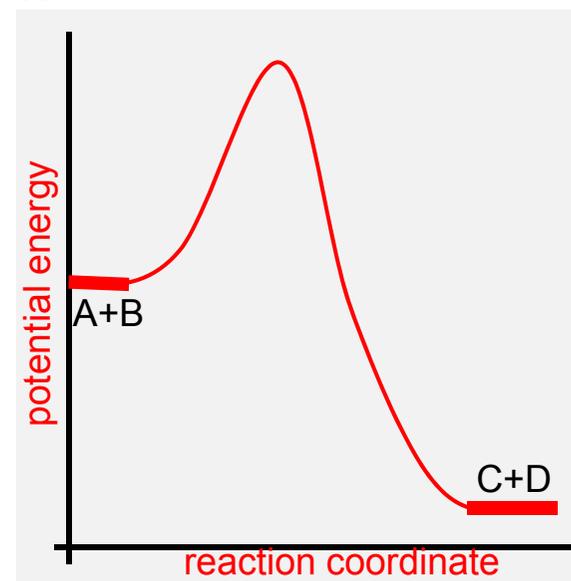
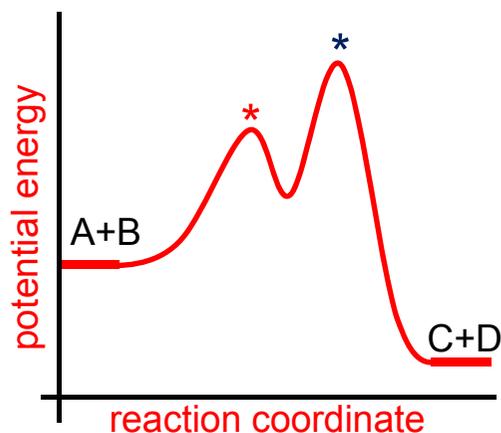
Some Basic Concepts

Reaction profile

The reaction profile describes the potential energy of a chemical system as a function of the reaction coordinate

Reaction coordinate

Describes the atomic movement during a reaction of reactants and products along the reaction profile

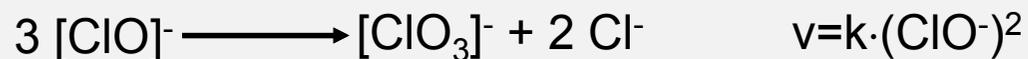


A reaction profile with two elementary steps

Some Basic Concepts

Molecularity

Describes the **number of molecules** which form the transition state of or participate in in an elementary step



the rate law generally **does not follow** the overall reaction equation

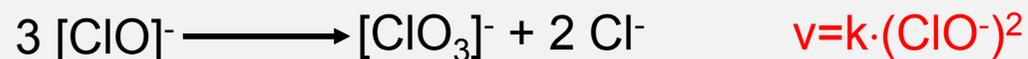
This makes sense since ...

... the overall reaction equation for a multi-step process is simply the net result of all of the **elementary reactions** in the mechanism

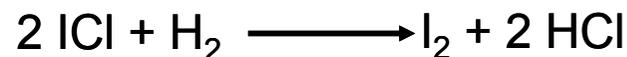
Some Basic Concepts

Reaction Order

Is expressed by the exponent in the rate law. It describes how many molecules of a reactant participate in an elementary step or in the overall reaction.



The reaction is **2nd order** in hypochlorite



implies a rate law which is 2nd order in ICl and 1st order in H₂

In fact, the rate law follows the equation



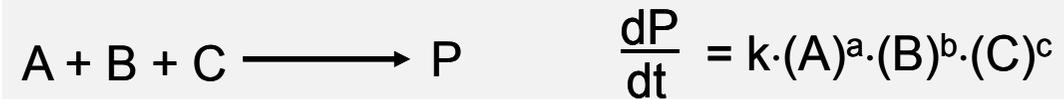
This reaction is 1st order in ICl and 1st order in H₂,
overall, it is a 2nd order (bimolecular) reaction

Some Basic Concepts

Rate Law

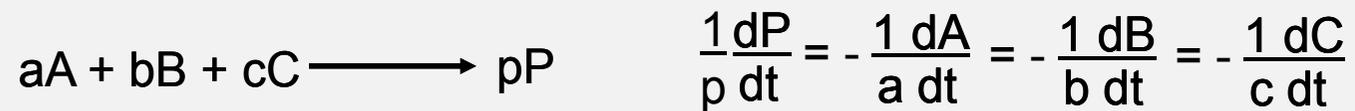
The rate law describes the **concentration change** of a reactant as a function of time

In its most general form it looks like:



a, b, c are the **reaction orders**,

they must not correspond to the stoichiometric coefficients



we find that $v = k \cdot (N_2) \cdot (H_2)$ but $-\frac{dN_2}{dt} = -\frac{1}{3} \frac{dH_2}{dt} = \frac{1}{2} \frac{dNH_3}{dt}$

Some Basic Concepts

Rate Law

The rate law describes the **concentration change** of a reactant as a function of time

In its most general form it looks like:

The rate law comprises all reactants influencing the course of a reaction, these **may or may not** show up in the stoichiometric equation.

A simple, bimolecular reaction with one elementary step



"k" is called the **Rate Constant** or the **Proportionality Factor**

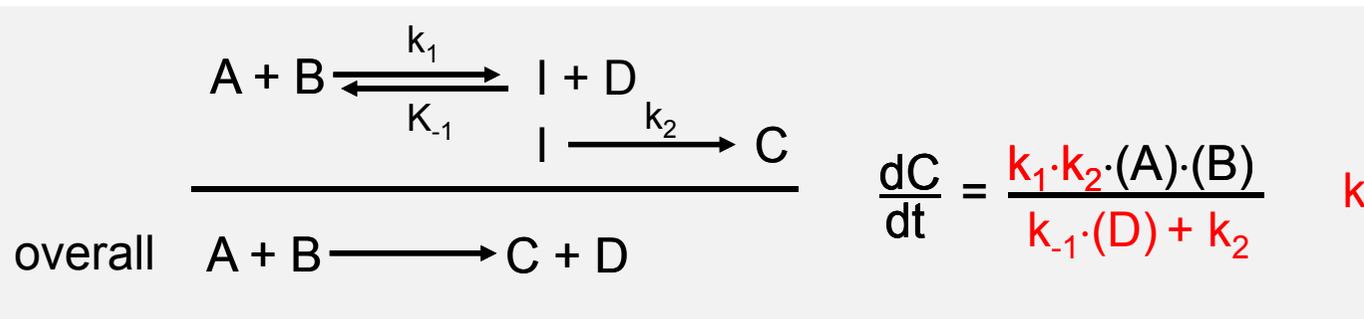
It is only "constant" if it describes one elementary step in a reaction scheme..

.. but may be **concentration dependent** in the rate law of an overall reaction

Some Basic Concepts

reaction equation $A + B \longrightarrow C + D$ implies $\frac{dC}{dt} = k \cdot (A) \cdot (B)$

from experiment: the more "D", the slower the reaction

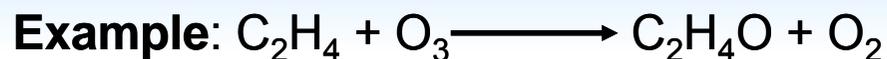


two elementary steps, k depends on D

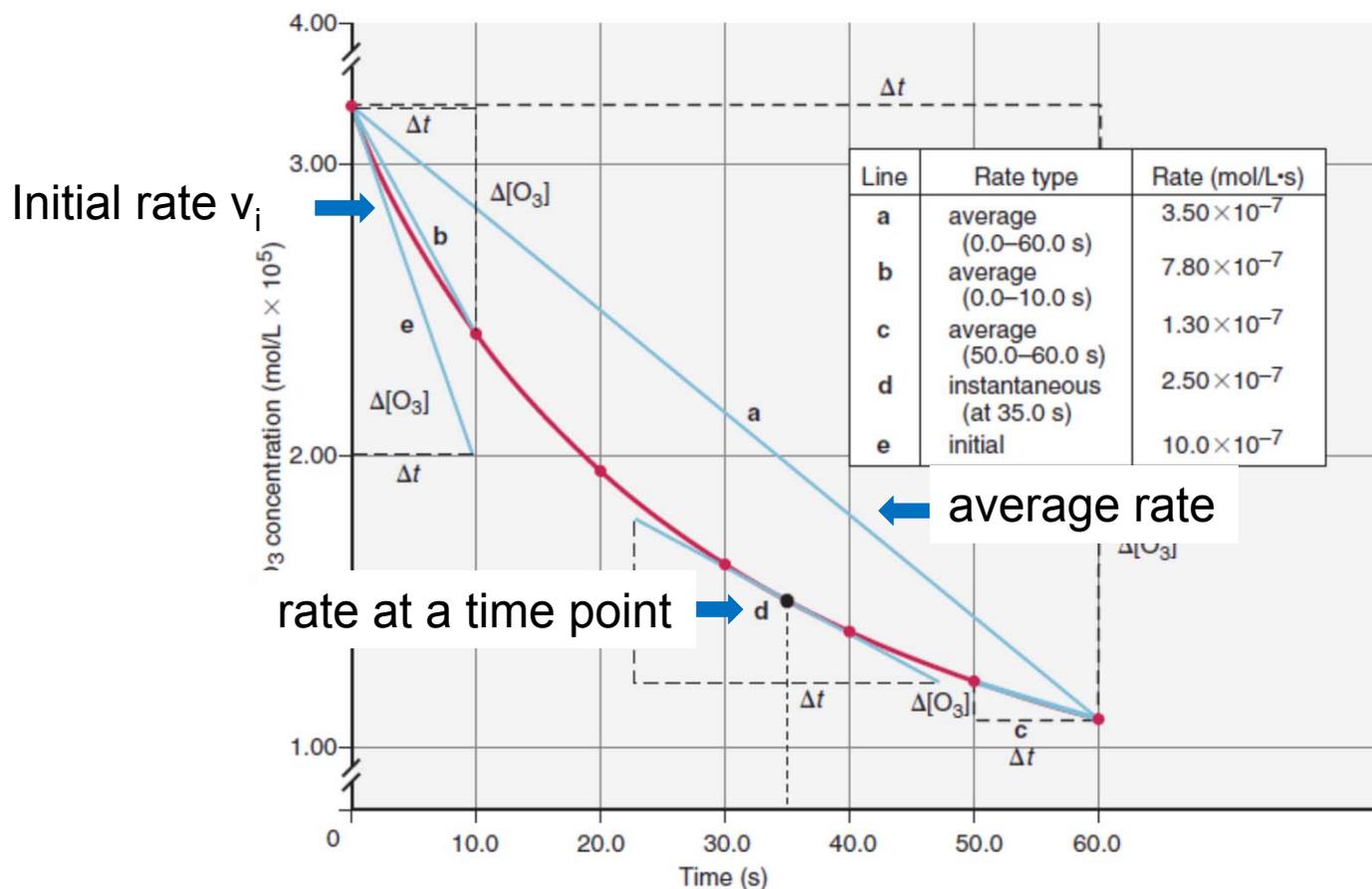
➔ "k" contains the information for a **mechanistic elucidation !**

Some Basic Concepts

Rates describes the concentration change of a reactant over a certain time



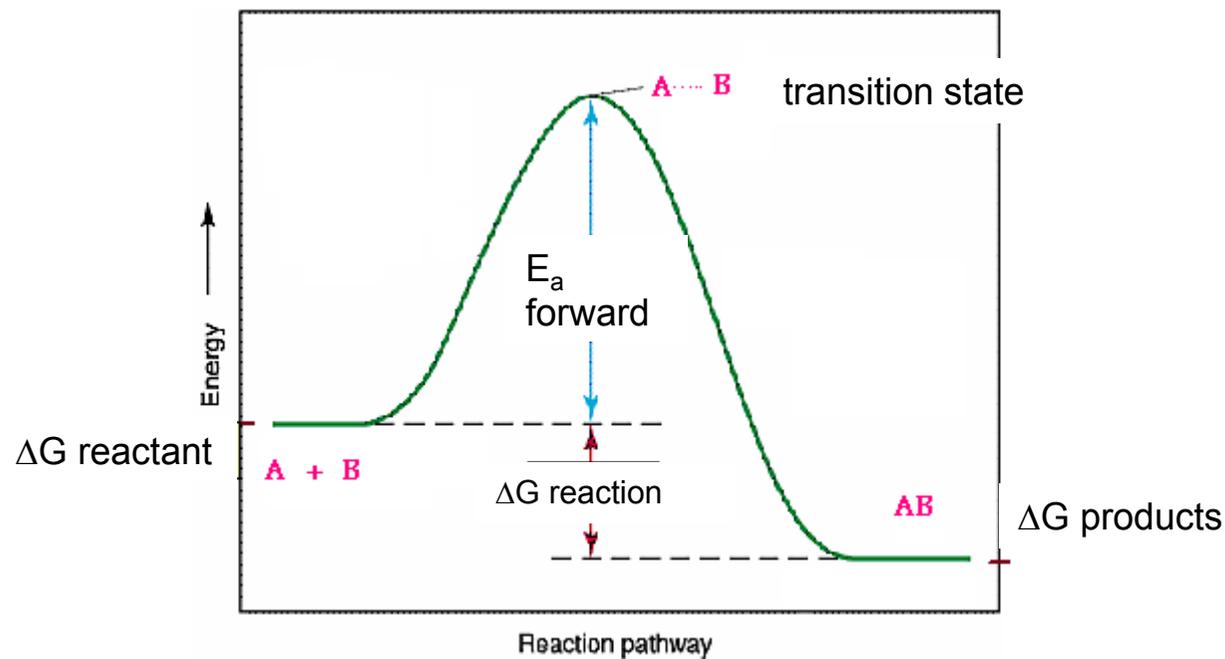
rates can **explicitly be calculated** at any time point, if the rate law is known



Some Basic Concepts

Thermodynamics and Kinetics

In a single elementary reaction, kinetic and thermodynamic are not coupled to each other e.g. the more exergonic a reaction the faster it runs.



however, the $\Delta\Delta G$ may translate into a $\Delta\Delta G^*$, then thermodynamics impacts kinetics

Some Basic Concepts

Activation Energies

Empirical: Arrhenius-equation

$$k = A \cdot \exp(-E_a/RT)$$

Arrhenius-Energy of activation

A: pre-exponential factor

Transition state theory: Eyring- equation

$$k = \frac{k_B \cdot T}{h} \cdot \exp(-\Delta G^\ddagger / RT) = \frac{k_B \cdot T}{h} \cdot \exp(\Delta S^\ddagger / R) \cdot \exp(-\Delta H^\ddagger / RT)$$

k_B : Boltzmann-constant, T: abs. temperature, h: Planck constant

$k_B \cdot T/h \approx 10^{13} \text{ sec}^{-1}$ bei 25°C

ΔG^\ddagger : free energy of activation

ΔH^\ddagger : enthalpy of activation

ΔS^\ddagger : entropy of activation

Mechanism and Rate Law

Example: Product formation via intermediate in equilibrium



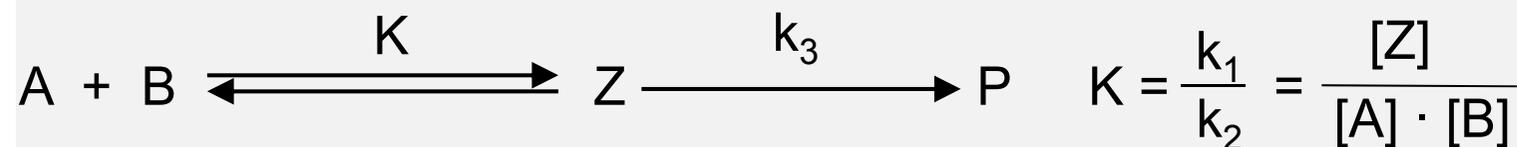
$$d[A]/dt = d[B]/dt = -k_1 \cdot [A] \cdot [B] + k_2 \cdot [Z]$$

Rate equations:

$$d[Z]/dt = k_1 \cdot [A] \cdot [B] - k_2 \cdot [Z] - k_3 \cdot [Z]$$

$$d[P]/dt = -d[Z]/dt = k_3 \cdot [Z] = f([A], [B])$$

Example: rapid prior equilibrium



$$d[P]/dt = k_3 \cdot [Z] = k_3 \cdot K \cdot [A] \cdot [B]; \quad k_3 \cdot K = \text{konst.}$$

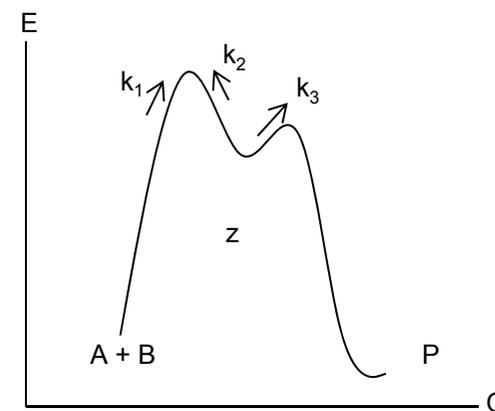
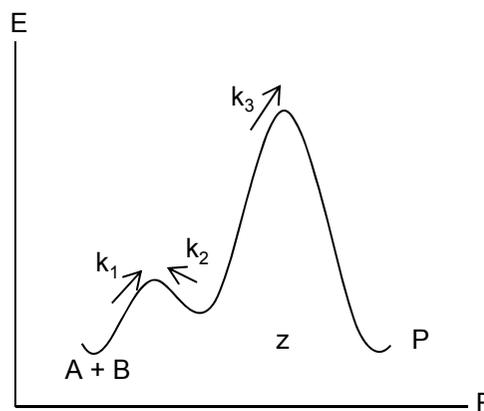
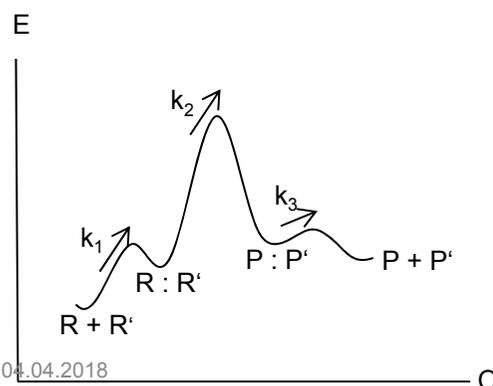
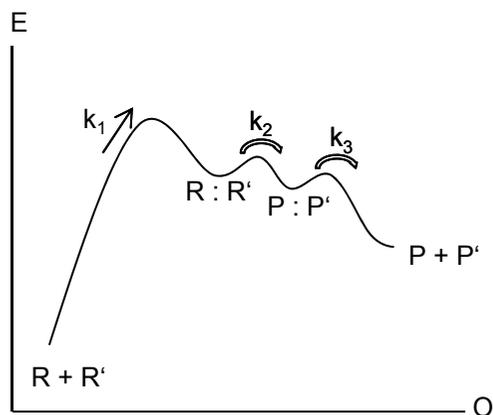
Mechanism and Rate Law

Example: Z is a short-lived intermediate

Principle of stationary: $d[Z]/dt = 0$

$$d[Z]/dt = 0 = k_1 \cdot [A] \cdot [B] - k_2 \cdot [Z] - k_3 \cdot [Z] \quad [Z] = \frac{k_1}{k_2 + k_3} \cdot [A] \cdot [B]$$

$$v = k_3 \cdot [Z] = \frac{k_3 \cdot k_1}{k_2 + k_3} \cdot [A] \cdot [B]; \quad \frac{k_3 \cdot k_1}{k_2 + k_3} = \text{konst.}$$



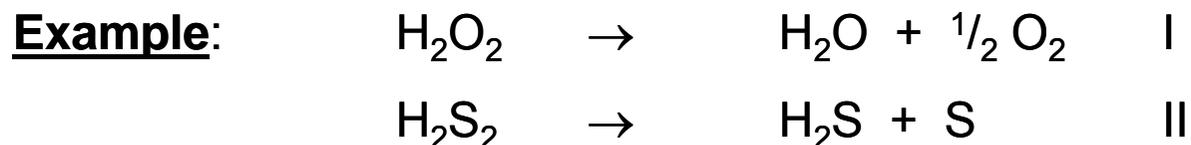
Mechanism and Structure

Geometry and Mechanism: $k = p \cdot Z \cdot e^{-E_a/RT}$

Steric factor p influences the respective geometric arrangements of colliding molecules (cross section).

Influence of: Size and geometry of atoms, molecules and ions

Example: Perchlorate und Periodate are strong oxidants.
Periodates are "larger" and react much faster.



I more exothermic than II, but II faster than I

Geometry of molecules

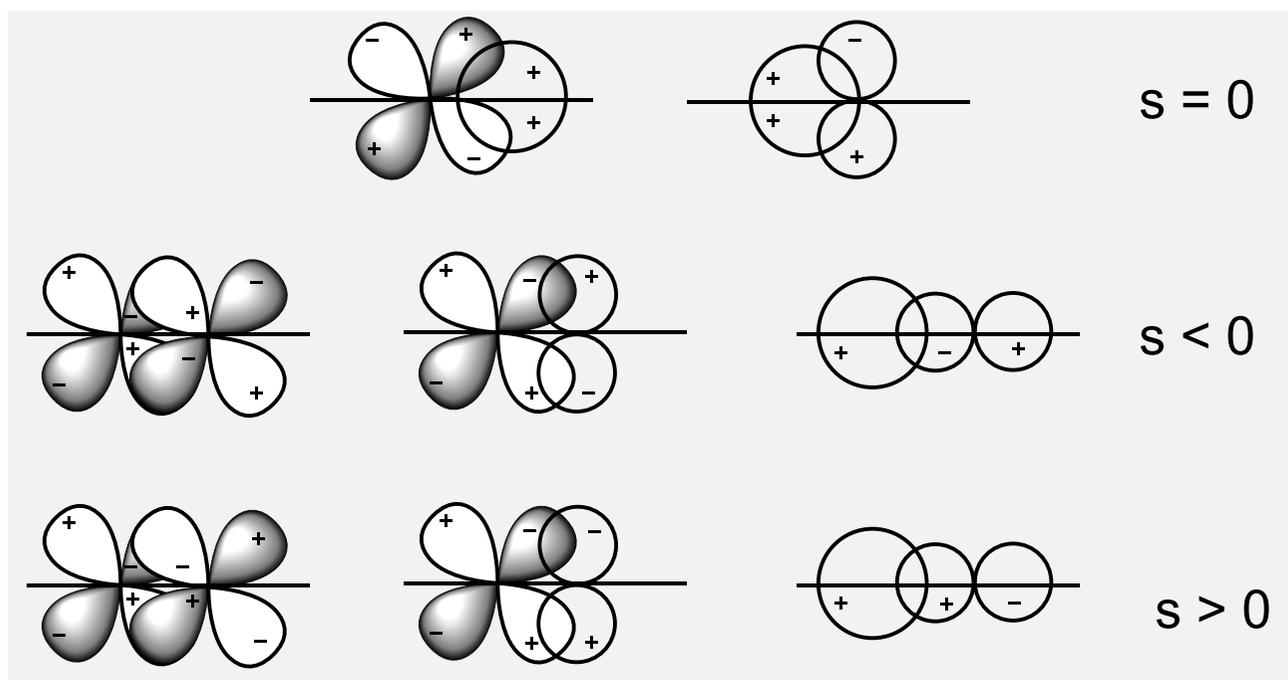
SF₄ hydrolyses very fast

SF₆ hydrolyses very slow if ever

Mechanism and Structure

Orbital control

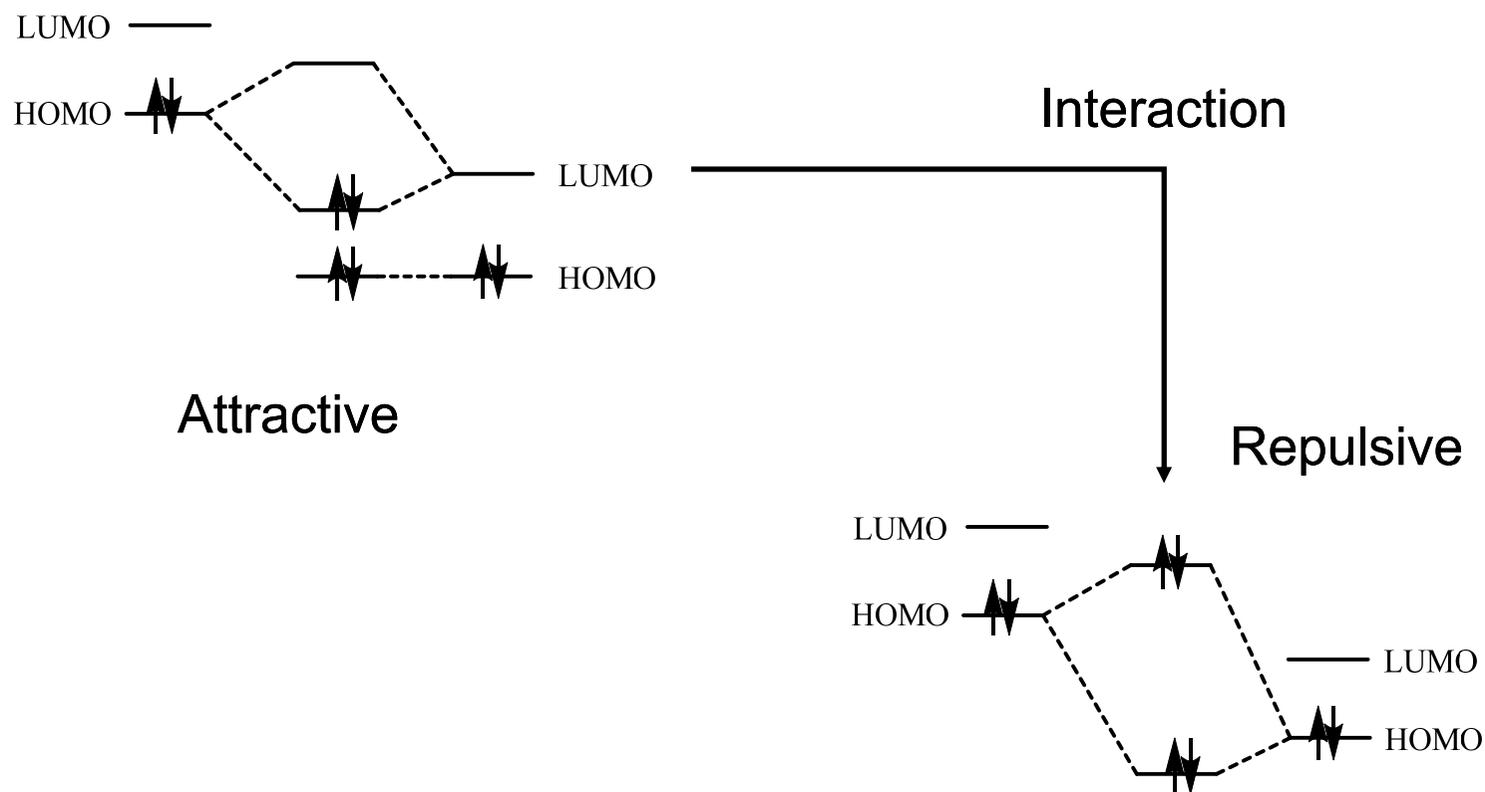
"least energy pathway" and the criteria of maximal overlap



„Least energy pathway“ controlled by HOMO – LUMO interaction

Mechanism and Structure

- **HOMO/LUMO-Interaction** of molecules as criteria for reaction pathway



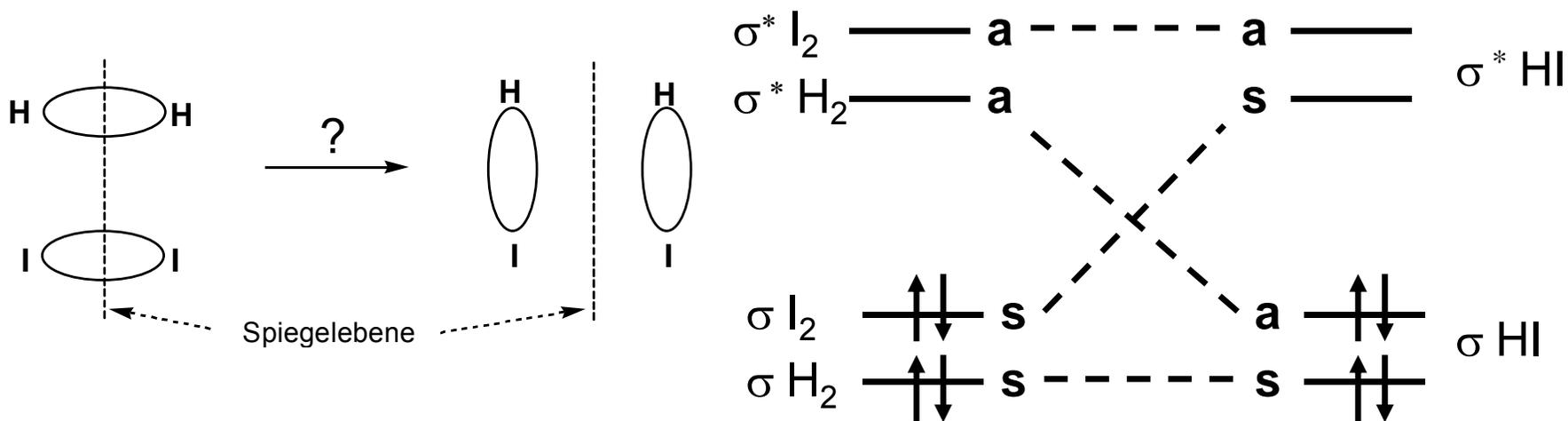
- Retention of orbital symmetry (Woodward-Hoffmann Rules)

Mechanism and Structure

Reactions of molecules with covalent bonds

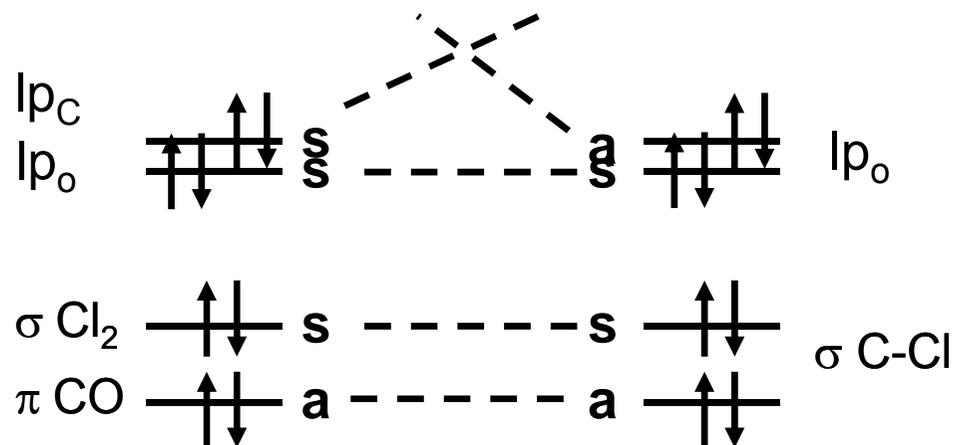
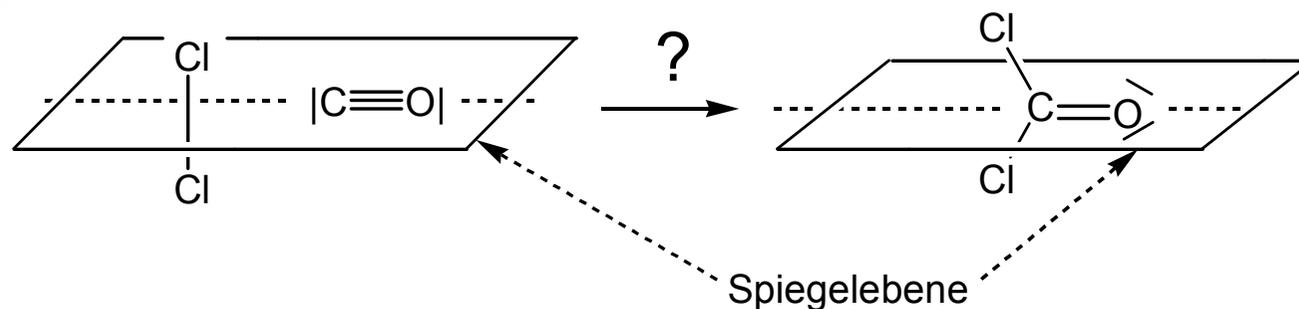
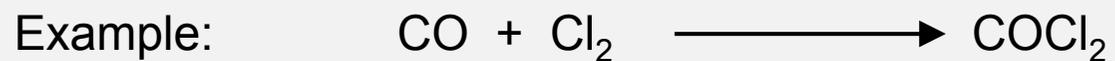
Retention of orbital symmetry as reaction path criteria

(Woodward-Hoffmann Rules)



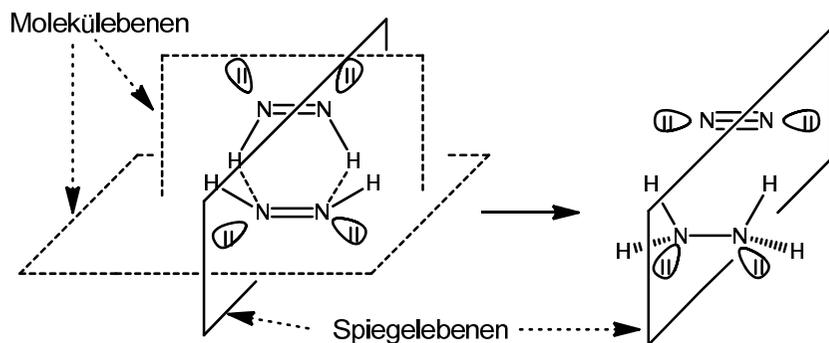
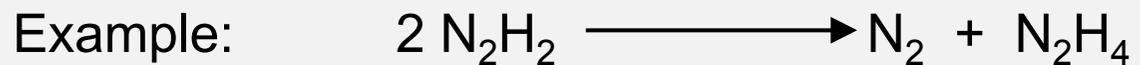
Symmetry forbidden \rightarrow different reaction pathway

Mechanism and Structure



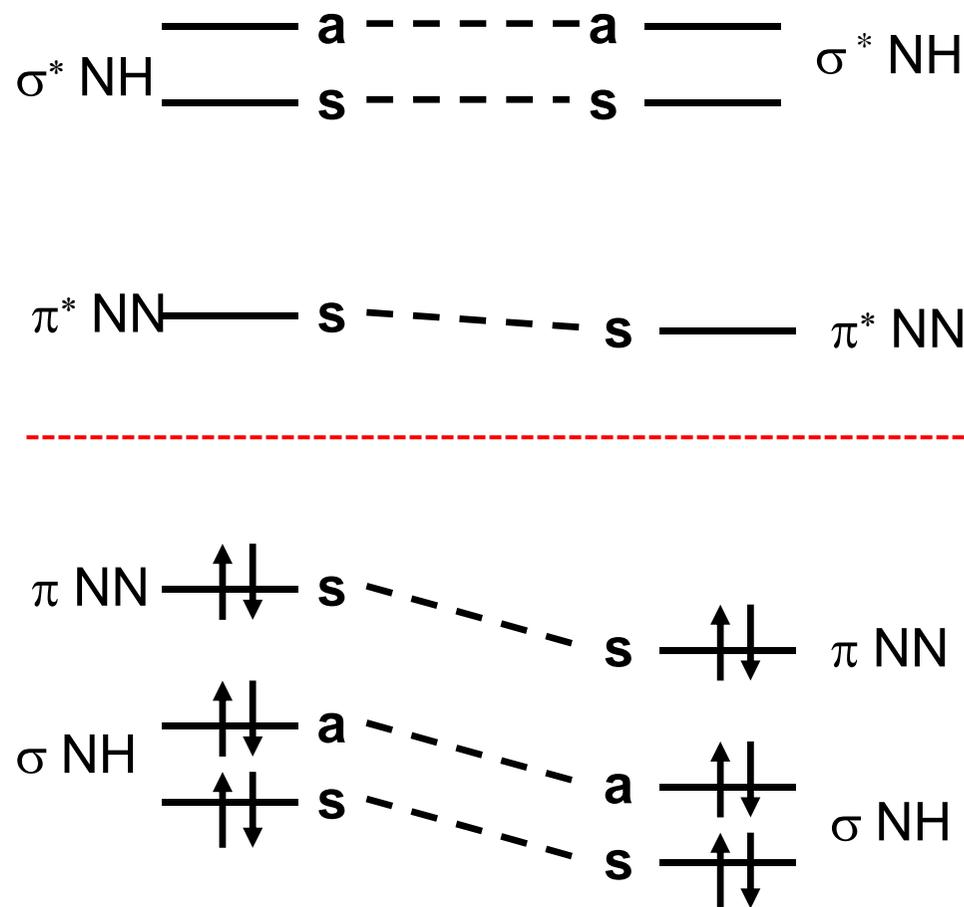
Symmetry forbidden \Rightarrow radical reaction

Mechanism and Structure



Symmetry allowed

➔ di-imine is very unstable



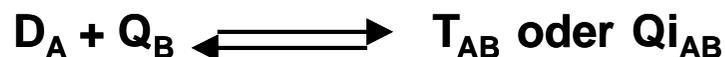
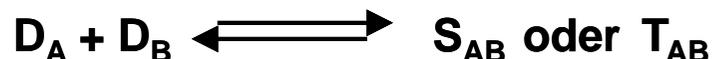
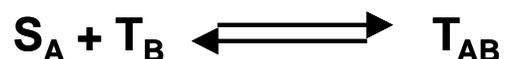
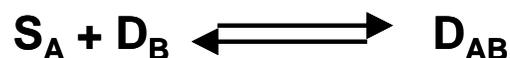
Mechanism and Structure

Radical reactions Spin conservation rules



AB = product, intermediate or activated complex

Spin allowed, if obeying the Wigner und Witmer rules:



unpaired electrons

S = Singulett	0
D = Doublet	1
T = Triplet	2
Q = Quartet	3
Qi = Quintet	4
Sx = Sextet	5
Sp = Septet	6

Mechanism and Structure

Spin Multiplicity Matrix for radical reactions

M_B	M_A							
	1	2	3	4	5	6	7	8
1	1	2	3	4	5	6	7	8
2		1,3	2,4	3,5	4,6	5,7	6,8	7,9
3			1,3,5	2,4,6	3,5,7	4,6,8	5,7,9	6,8,10
4				1,3,5,7	2,4,6,8	3,5,7,9	4,6,8,10	5,7,9,11

Multiplicities M_A und M_B : $2S+1$

here $M_A \geq M_B$

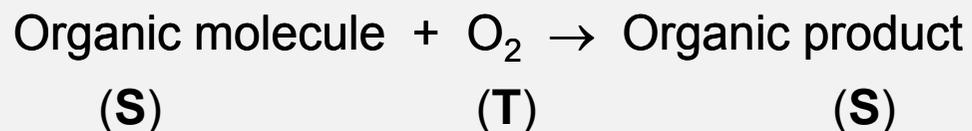
Maximum multiplicities for reactions M_A und $M_B = M_A + M_B - 1$

Minimum multiplicities for reactions M_A und $M_B = M_A - M_B + 1$

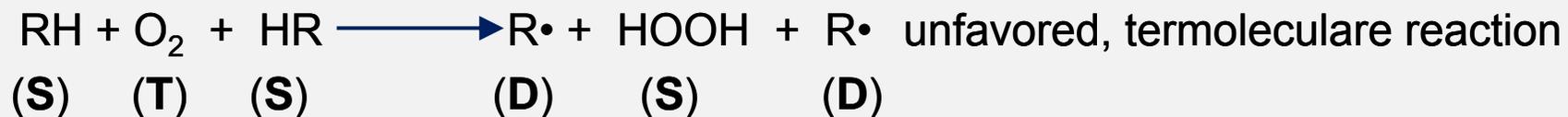
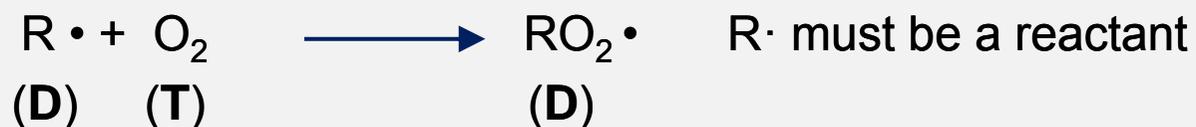
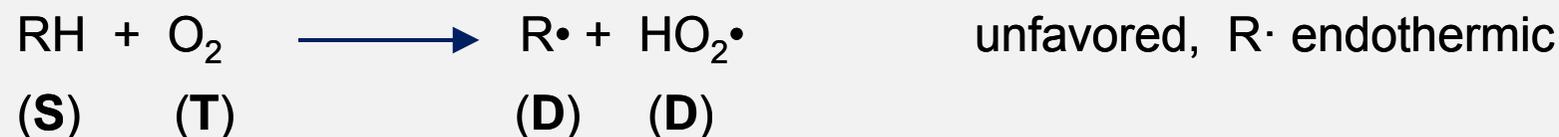
Mechanism and Structure

Radical Reactions

„Spin forbidden“



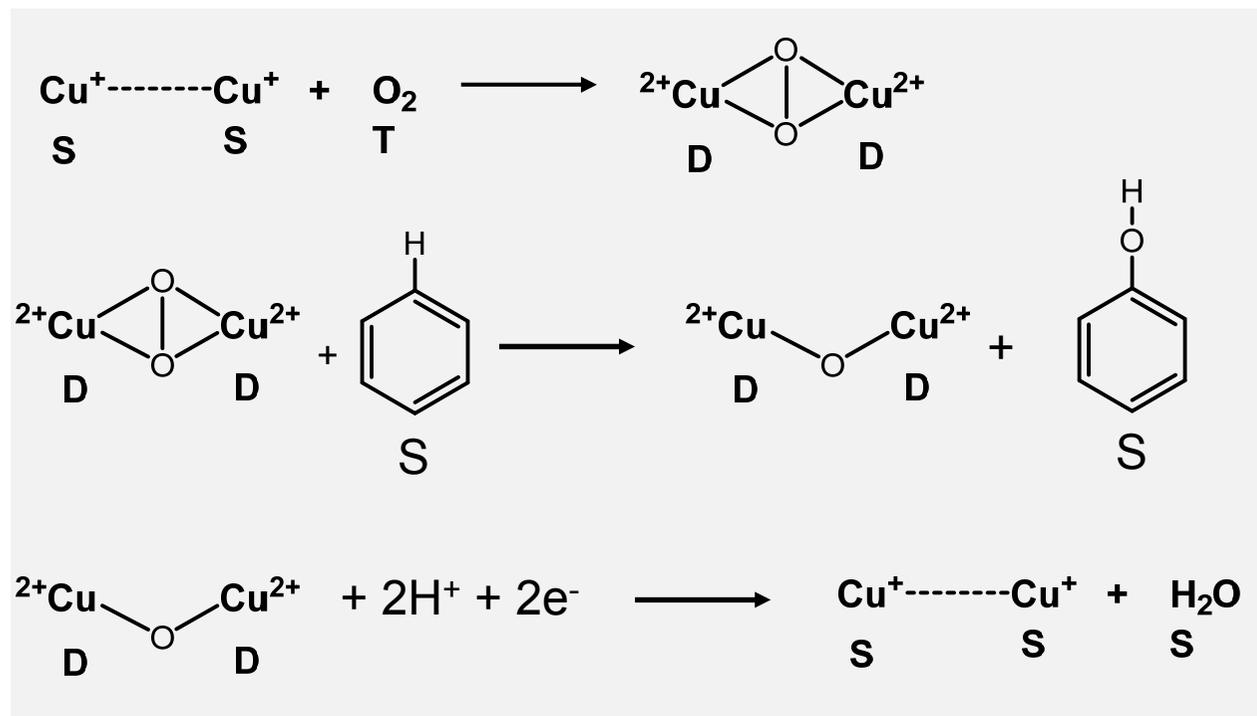
„Spin allowed“



Mechanism and Structure

Radical reactions: how does nature "activate e.g. $^3\text{O}_2$?

by spin transfer to metal centres

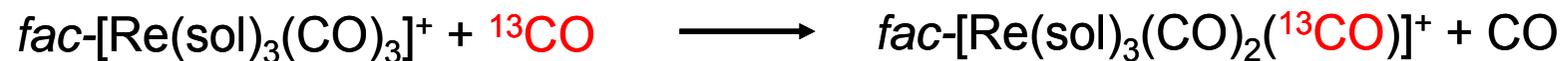


Hemocyanine acts as O₂ transporter and as O₂ activator

Similar reactivities are also found for Fe^{2+/3+} mono- or dinuclear enzymes

Exercises

Find the elementary steps in the following reactions



2. Rate Laws and Mechanisms

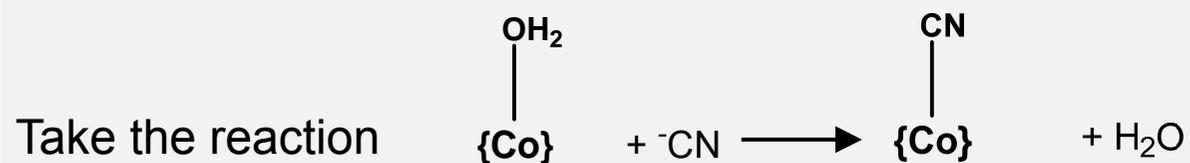
2.1 Simple kinetic rate laws



Reactions following a **1st order mechanism** are rare (do you know one?)

These are generally **intramolecular** reactions

Important: To deduce a mechanism from kinetics, even complicated schemes can be reduced to 1st (pseudo 1st) order rate laws



Follows the rate law $\frac{d\text{Co}(\text{CN})}{dt} = \frac{(k_1 k_2 [\text{CN}^-] \{\text{Co}\})}{(k_{-1} + k_2 [\text{CN}^-])}$ \Rightarrow 1st order?

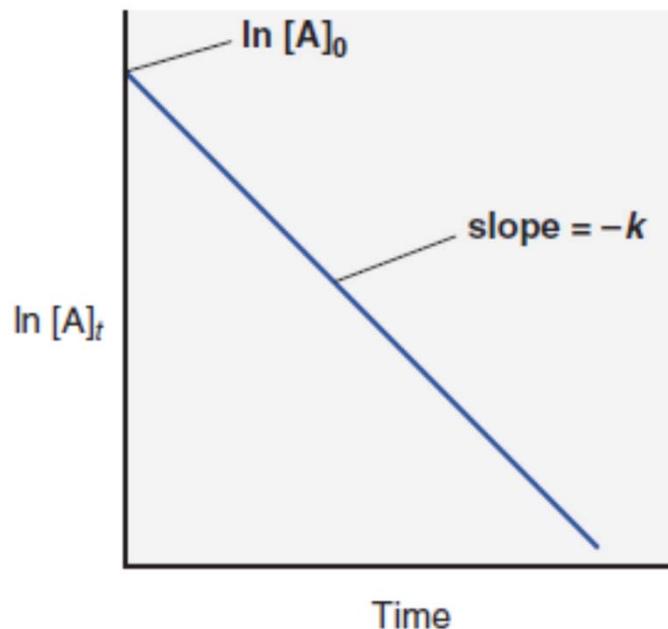
2. Rate Laws and Mechanisms

2.1 Simple kinetic rate laws



can easily be integrated

$$\ln[A] = \ln[A_0] - kt \quad \text{or} \quad A(t) = A_0 e^{-kt} \quad \text{and} \quad t_{1/2} = \frac{\ln 2}{k}$$



half life time independent of A_0

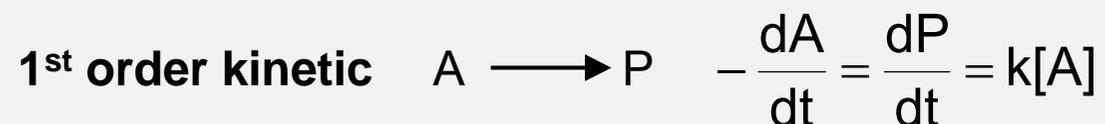
graphical representation $\ln[A]$ vs. t is linear

➔ **slope gives k (lin reg)**

variation of $[A_0]$ over a broad range
should yield **constant k**

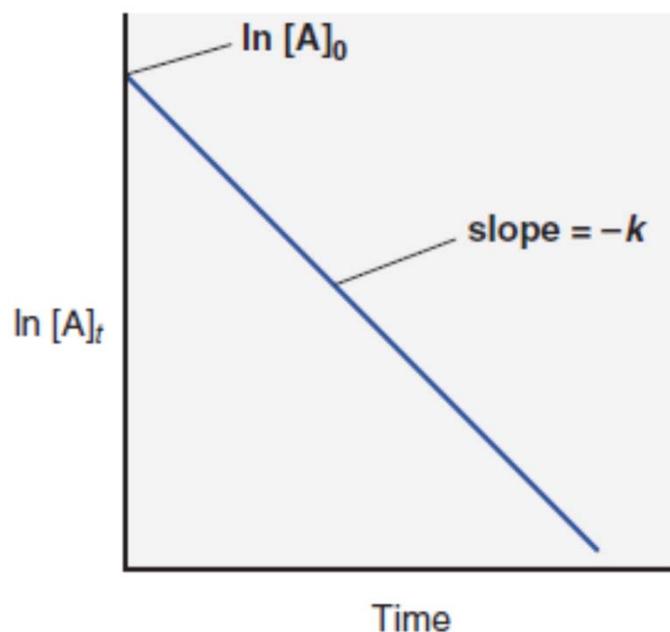
2. Rate Laws and Mechanisms

2.1 Simple kinetic rate laws



can easily be integrated

$$\ln[A] = \ln[A_0] - kt \quad \text{or} \quad A(t) = A_0 e^{-kt} \quad \text{and} \quad t_{1/2} = \frac{\ln 2}{k}$$



calculate k from each point $[A]_t$ relative to $[A_0]$, should yield a constant k , no trends

$$\Rightarrow k = \frac{1}{t} \cdot \ln \frac{[A]_0}{[A]_t}$$

Kinetic pseudo 1. order $k_1 \rightarrow k_{\text{obs}}$

2. Rate Laws and Mechanisms

2.1 Simple kinetic rate laws

Let's look at a **bimolecular, single step** reaction



We choose $[B]_0 \gg [A]_0$ so that $[B](t) = [B]_0 = \text{constant}$

$$v = k_{\text{obs}} [A] \quad 1. \text{ Ordnung mit } k_{\text{obs}} = k_1 [B]_0$$

graphical representation of k_{obs} vs $[B]_0$ gives k_1 \rightarrow **pseudo 1st order reaction**

$$\frac{d\text{Co(CN)}}{dt} = \frac{(k_1 k_2 [\text{CN}^-] \{\text{Co}\})}{(k_{-1} + k_2 [\text{CN}^-])}$$

can also be considered as pseudo 1st order
but k_{obs} vs. $[\text{CN}]_0$ is not linear

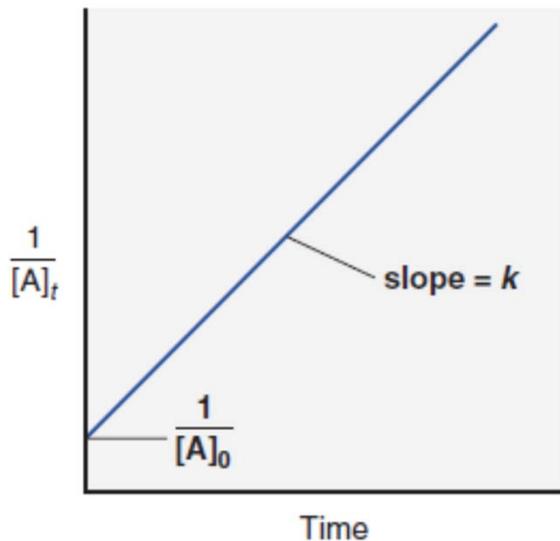
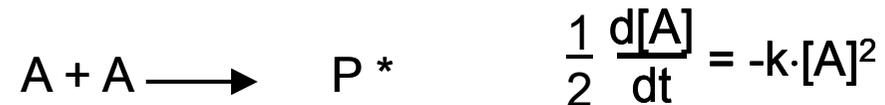
2. Rate Laws and Mechanisms

2.1 Simple kinetic rate laws

Let's look at a **bimolecular, single step** reaction



2. Order Kinetics



$$\frac{1}{[A]_t} = \frac{1}{[A]_0} + 2 \cdot k \cdot t \quad \rightarrow \quad [A]_t = [A]_0 - 2 \cdot k [A]_0 \cdot t \quad [A]_t$$

graphic representation

$[A]_t$ vs. $t \cdot [A]_0 \cdot [A]_t$ gives line with slope $-2k$

or

$[A]_t^{-1}$ vs. t

$$t_{1/2} = \frac{1}{k[A]_0}$$

2. Rate Laws and Mechanisms

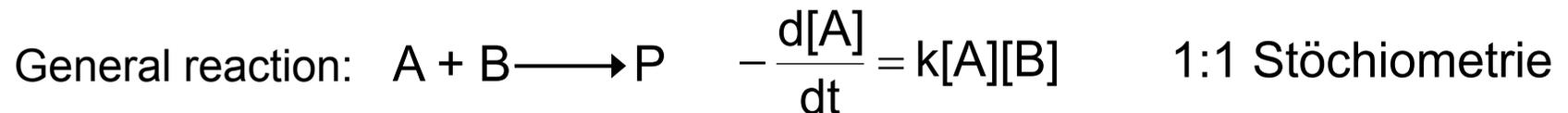
2.1 Simple kinetic rate laws

Let's look at a **bimolecular, single step** reaction



$$t_{1/2} = \frac{1}{k \cdot [A]_0}$$

depends on initial concentration



For $[A]_0 = [B]_0$, treatment as for [A] only since $[A]_t = [B]_t$

For slight excess of $[A]_0$ or $[B]_0$ over the other

Concentration dependencies:

$$[B]_t = \underbrace{[B]_0 - [A]_0}_{\Delta_0} + [A]_t$$

2. Rate Laws and Mechanisms

2nd order rate laws

Concentration dependencies

$$[B]_t = [B]_0 - \underbrace{[A]_0}_{\Delta_0} + [A]_t$$

rate law: $\frac{dA}{[A](\Delta_0 + [A])} = -k \cdot dt$

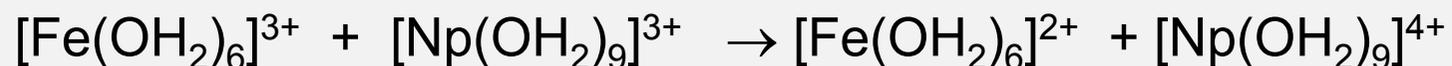
can be integrated: $\ln \frac{[B]_t}{[A]_t} = \ln \frac{[B]_0}{[A]_0} + k \cdot \Delta_0 \cdot t$

$[B]_t$ can be substituted by $[A]_t$ and $[A]_0$

If $[B]_0 \gg [A]_0$, $[B]_t = \text{const.} = [B]_0$ and equation **reduces to pseudo 1st law**

2. Rate Laws and Mechanisms

Example: outer sphere electron transfer reaction



Reaction is monitored by UV/vis spectroscopy

$$[\text{Np}^{3+}]_0 = 1.53 \cdot 10^{-4} \text{ M}, \quad [\text{Fe}^{3+}]_0 = 2.24 \cdot 10^{-4} \text{ M}$$

Absorption for $t_\infty = 0.351$

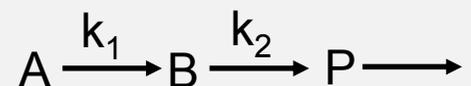
Zeit/ s:	0	2.5	3.0	4.0	5.0	7.0	10.0	15.0	20.0
Absorption*:	0.100	0.228	0.242	0.261	0.277	0.300	0.316	0.332	0.341

calculate the rate constant for $-k [\text{Np}^{3+}] [\text{Fe}^{3+}]$ along different approaches

* absorption at 723 nm, $\text{Fe}^{2+/3+}$ do not absorb at this wave length

2. Rate Laws and Mechanisms

2.2 Consecutive Reactions



typical scheme for radioactive decay in **actinide elements**

B can be a **true intermediate** which accumulates or a short lived intermediate

in chemistry: **no step of reversibility**

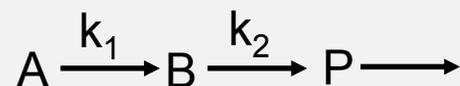
$$\frac{d[A]}{dt} = -k_1[A] \quad \frac{d[B]}{dt} = k_1[A] - k_2[B] \quad \frac{d[C]}{dt} = -k_2[B]$$

dissolve 1st equation: $[A] = [A]_0 \cdot e^{-k_1 \cdot t}$... and substitute in 2nd equation, integrate

$$[B]_t = \frac{k_1[A]_0}{k_2 - k_1} (\exp(-k_1 \cdot t) - \exp(-k_2 \cdot t))$$

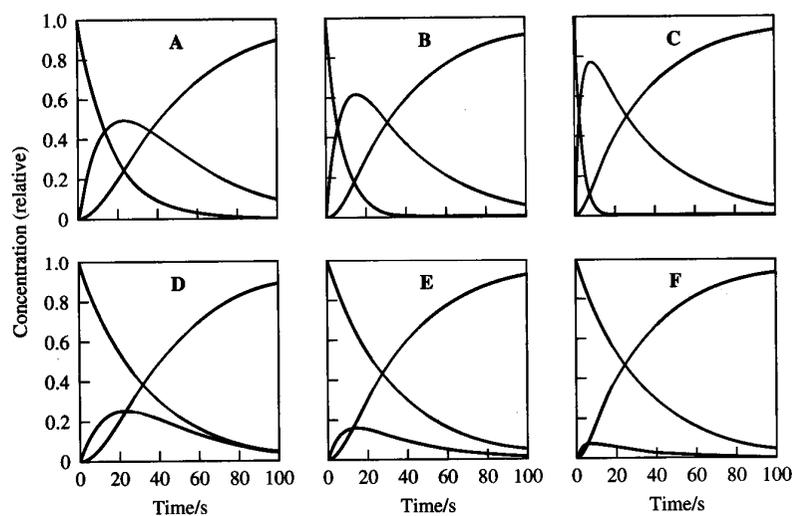
2. Rate Laws and Mechanisms

2.2 Consecutive Reactions



since $[C] = [A]_0 - [B] - [A]$ $[C]_t = \left(1 + \frac{k_1 \exp(-k_2 t) - k_2 \exp(-k_1 t)}{k_2 - k_1}\right) \cdot [A]_0$

$[B]_t$ will go through a maximum $\Rightarrow \frac{dB}{dt} = 0 \Rightarrow t_{\max} = \frac{\ln(k_2/k_1)}{(k_2 - k_1)}$



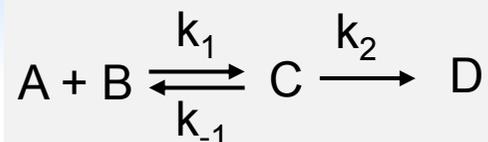
Panel	k_1/s^{-1}	k_2/s^{-1}	$t_{1/2}(A)/s$	t_{\max}/s	$[I]_{\max}$
A	0.06	0.03	11.55	23.10	0.500
B	0.12	0.03	5.78	15.40	0.630
C	0.30	0.03	2.31	8.60	0.774
D	0.03	0.06	23.10	23.10	0.250
E	0.03	0.12	23.10	15.40	0.158
F	0.03	0.30	23.10	8.60	0.0774

→ biexponential function in k_1 and k_2

→ or non-linear least square fitting

2. Rate Laws and Mechanisms

2.2 Consecutive Reactions with one reversible step



$$\frac{d[A]}{dt} = \frac{d[B]}{dt} = -k_1[A][B] + k_{-1}[C]$$

equations: $\frac{d[C]}{dt} = k_1[A][B] - k_2[C]$ \Rightarrow cannot be solved analytically
numerical integration

$$\frac{d[D]}{dt} = k_2[C]$$

assumption: if $k_{-1} \gg k_2$ then $k_1[A][B] = k_{-1}[C]$ and $K = \frac{k_1}{k_{-1}}$

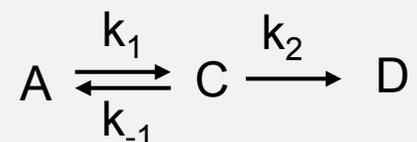
therefore $\frac{d[D]}{dt} = k_2[C] = k_2K \cdot [A] \cdot [B]$ and $k_{\text{obs}} = k_2 \cdot K$ 2nd order kinetics

This is also a simplified form of the **pre-equilibria** situation (see later)

2. Rate Laws and Mechanisms

2.2 Consecutive Reactions with one reversible step

simpler form



$$\text{Rate laws: } \frac{d[A]}{dt} = -k_1[A] + k_{-1}[C] \quad \frac{d[C]}{dt} = k_1[A] - k_{-1}[C] + k_2[C]$$

$$\frac{d[I]}{dt} = k_1[A] - k_{-1}[I] - k_2[I] \quad [A]_t = \frac{k_1[A]_0}{\lambda_2 - \lambda_3} \left\{ \frac{\lambda_2 - k_2}{\lambda_2} \exp(-\lambda_2 t) - \frac{\lambda_3 - k_2}{\lambda_3} \exp(-\lambda_3 t) \right\}$$

Can be solved explicitly but does not make sense for chemical reactions

➔ approximations

$$\text{therefore } \frac{d[D]}{dt} = k_2[C] = k_2 K \cdot [A] \cdot [B] \quad \text{and } k_{\text{obs}} = k_2 \cdot K \quad 2^{\text{nd}} \text{ order kinetics}$$

2. Rate Laws and Mechanisms

2.3 Reversible and concurrent reactions



$$\frac{dA}{dt} = -k_{+1}[A] + k_{-1}[P] \quad \text{at equilibrium} \quad \frac{dA}{dt} = 0$$

$$[A]_0 + [P]_0 = [A]_e + [P]_e = [A]_t + [P]_t$$

$$k_{+1} [A]_e = k_{-1} [P]_e \quad \frac{[P]_e}{[A]_e} = K = \frac{k_{+1}}{k_{-1}}$$

can be integrated after transformation

$$[A]_t = [A]_e + ([A]_0 - [A]_e)(\exp[-(k_{+1} + k_{-1})t])$$

$\Rightarrow \ln([A]_t - [A]_e)$ vs. t gives slope $k_{+1} + k_{-1}$

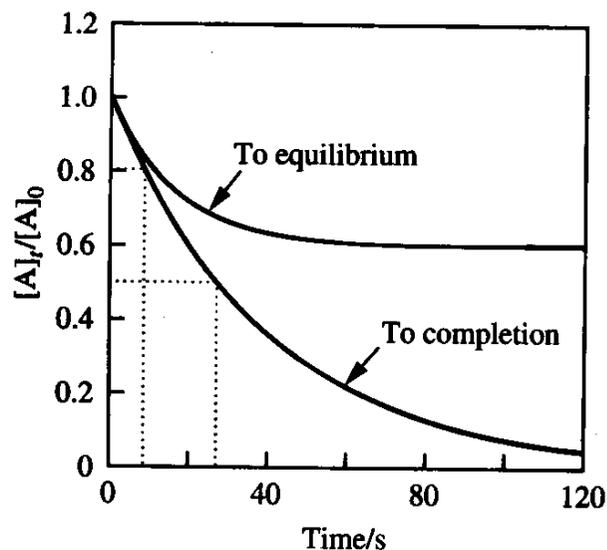
2. Rate Laws and Mechanisms

2.3 Reversible and concurrent reactions

→ $k_{+1} + k_{-1}$ is the apparent and **observed** k_e

$$k_e = k_{+1} + k_{-1} \quad \text{oder} \quad k_{+1} = \frac{k_e}{1+K^{-1}}, k_{-1} = \frac{k_e}{1+K}$$

$[A]_t$ decreases with k_e $t_{1/2} = \ln 2/k_e$ → k_e can be calculated

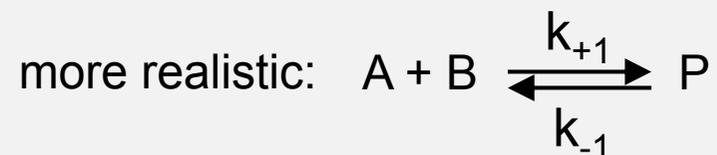


If P is involved in a consecutive equilibrium,
 $t_{1/2}$ becomes $\ln 2/k_1$

or if $k_{-1} \ll k_{+1}$, simple one step reaction

2. Rate Laws and Mechanisms

2.3 Reversible and concurrent reactions



such schemes reduce to the former if $[B_0] \gg [A_0]$

p.m. $\ln([A]_t - [A]_e)$ vs. t gives slope $k_e = k_1 + k_{-1}$

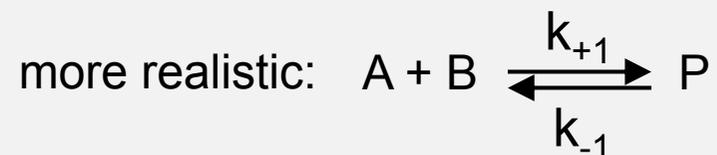
now " k_1 " ($=k_{obs1}$) becomes $k_1[B_0]$

which can be substituted to yield

$$k_e = k_1[B_0] + k_{-1} \quad \text{and} \quad k_1 = \frac{k_e}{[B_0] + K^{-1}} \quad k_{-1} = \frac{k_e}{1 + K[B_0]}$$

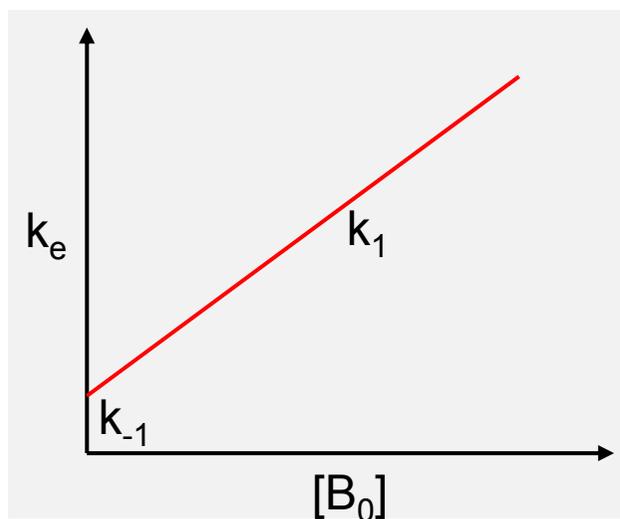
2. Rate Laws and Mechanisms

2.3 Reversible and concurrent reactions



$$k_e = k_1[B_0] + k_{-1} \quad \text{and} \quad k_1 = \frac{k_e}{[B_0] + K^{-1}} \quad k_{-1} = \frac{k_e}{1 + K[B_0]}$$

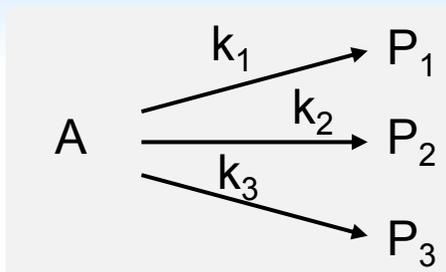
note: kinetic measurements allow the determination of equilibrium constants



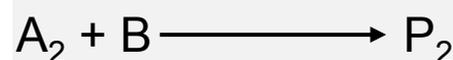
2. Rate Laws and Mechanisms

2.3 Reversible and concurrent reactions

parallel reactions:



concurrent reactions:



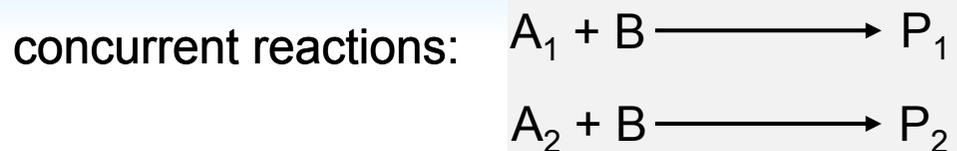
may be a side reaction

product ratio $P_1/P_2 = k_1/k_2$ if A_1 and A_2 are in large excess over B

if not the case: $\frac{k_1}{k_2} = \frac{\ln \left[\frac{[A_1]_0}{[A_1]_0 - [P_1]_f} \right]}{\ln \left[\frac{[A_2]_0}{[A_2]_0 - [P_2]_f} \right]}$ f=final

2. Rate Laws and Mechanisms

2.3 Reversible and concurrent reactions



often, only one or the other product is observable analytically

B is limiting since not in excess, thus e.g. $[P_1]_f = [B]_0 - [P_2]_f$

if we observe P_2 we know $k_2 \Rightarrow k_1 = k_2 \cdot \frac{[A_2]_0([B]_0 - [P_2]_f)}{[A_1]_0[P_2]_f}$

if not the case: $\frac{k_1}{k_2} = \frac{\ln \left\{ \frac{[A_1]_0}{[A_1]_0 - [P_1]_f} \right\}}{\ln \left\{ \frac{[A_2]_0}{[A_2]_0 - [P_2]_f} \right\}}$

3. Overall Reaction and Reaction Rate

k_{obs} contains the mechanistic information

complicated mechanisms should be reduced to pseudo 1st order kinetics

e.g. for $v = -k_1 [A] [B]$

chose $[B]_0 \gg [A]_0$

k_{obs} is then in the simplest case $k_1 [B]_0$

$$k_1 = \frac{k_{obs}}{[B]_0}$$

graphical evaluation

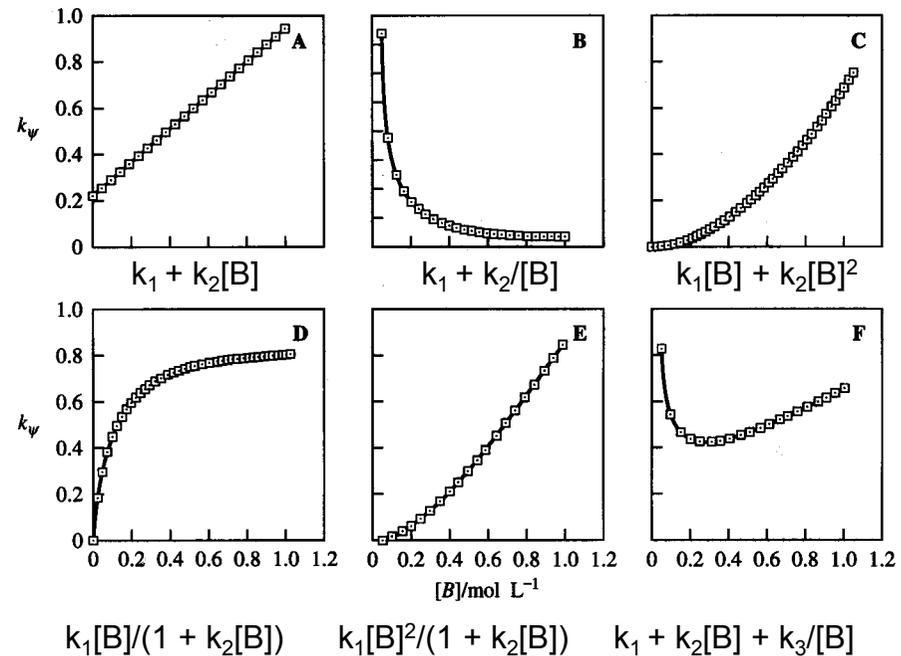


line goes through origin

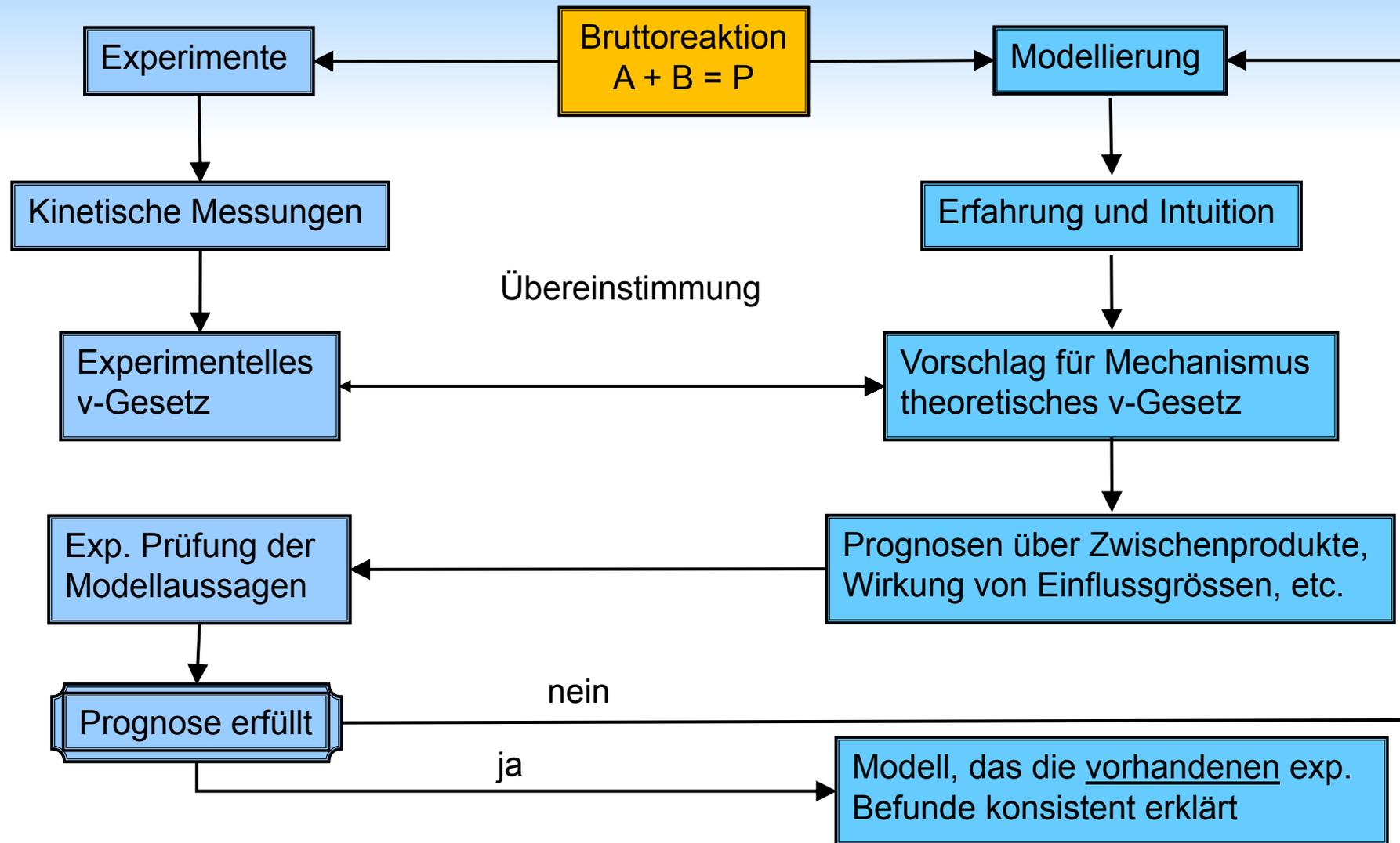
but: we often find non-linear behaviour for k_{obs} vs. $[B]_0$

possible dependencies

mechanism is more complex!



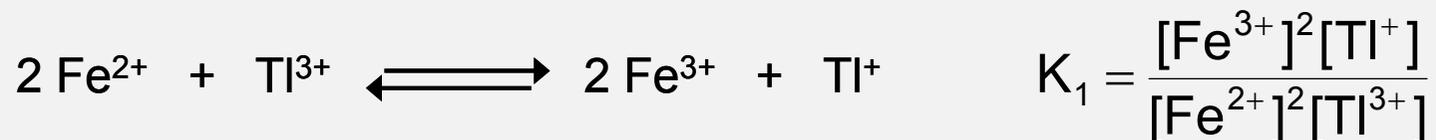
3. Overall Reaction and Reaction Rate



Mechanismus $\xleftrightarrow[\text{mehrdeutig}]{\text{eindeutig}}$ Geschwindigkeitsgesetz

Overall Reaction and Reaction Rate

The overall reaction shows the number of participating particles



does not mean that $v = k [\text{Fe}^{2+}]^2 [\text{Tl}^{3+}]$

the rate law is proportional to concentrations but must be deduced experimentally

in fact: $v_1 = k [\text{Fe}^{2+}] [\text{Tl}^{3+}]$

p.m. stoichiometry defines the time dependence of the particles

$$-\frac{1}{2} \frac{d[\text{Fe}^{2+}]}{dt} = -\frac{d[\text{Tl}^{3+}]}{dt} = \frac{1}{2} \frac{d[\text{Fe}^{3+}]}{dt} = \frac{d[\text{Tl}^{+}]}{dt}$$

and, if no intermediates show up in significant concentrations

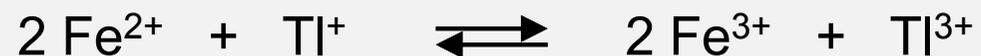
$$\frac{1}{2}([\text{Fe}^{2+}]_0 - [\text{Fe}^{2+}]_t) = ([\text{Tl}^{3+}]_0 - [\text{Tl}^{3+}]_t) = \frac{1}{2}([\text{Fe}^{3+}]_t - [\text{Fe}^{3+}]_0) = ([\text{Tl}^{+}]_t - [\text{Tl}^{+}]_0)$$

Overall Reaction and Reaction Rate

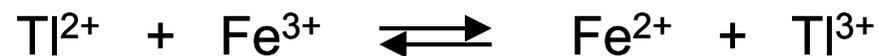
a mechanism is **the sum of all elementary steps**

reactions looking termolecular are often bimolecular

for instance this one



in fact, a reaction with two elementary steps



how to deduce the **mechanism**, how to formulate the rate law ?

how to explain the **reaction order = 1** in $[\text{Fe}^{2+}]$

Overall Reaction and Reaction Rate

reaction mechanism often depend on particles which **do not show up** in the equation

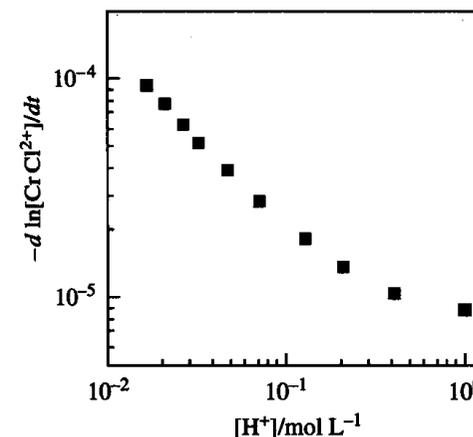


$$v_1 = k_1 [(\text{H}_2\text{O})_5\text{CrCl}][\text{H}_3\text{O}^+]$$

$$-\frac{d[(\text{H}_2\text{O})_5\text{CrCl}^{2+}]}{dt} = \left(k_1 + \frac{k_2}{[\text{H}^+]} \right) [(\text{H}_2\text{O})_5\text{CrCl}^{2+}]$$

$$k_{\text{obs}} = k_1 + \frac{k_2}{[\text{H}^+]}$$

the more alkaline, the faster, why?



Overall Reaction and Reaction Rate

reaction mechanism often **do not depend** on particles which show up in the equation

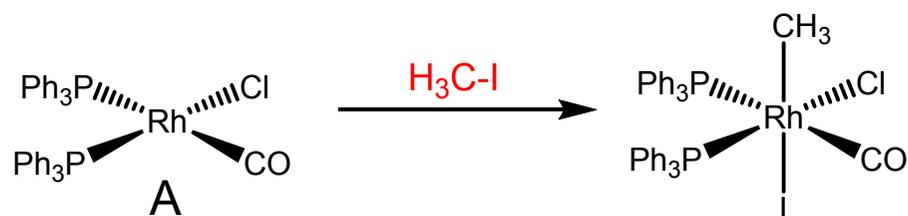


$$v = k [(\text{NH}_3)_5\text{CoBr}]$$

→ composed of multi elementary steps



unpredictable based on the reaction equation



$$v = k_{\text{obs}} \cdot [\text{A}] [\text{CH}_3\text{I}] \cdot [\text{I}^-]$$

Overall Reaction and Reaction Rate

Parameters which influence the reaction rate

Concentration: at least one concentration will accelerate the reaction

Products: can accelerate or decelerate the reaction

Impurities: can have a (positive or negative) catalytic effect

pressure / temperature / radiation

Solvent: polarity, viscosity, donor-acceptor properties, electrolytes buffer etc.

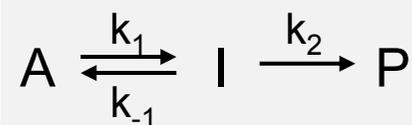
Concentration dependencies

determine the rate if all other parameters are kept constant

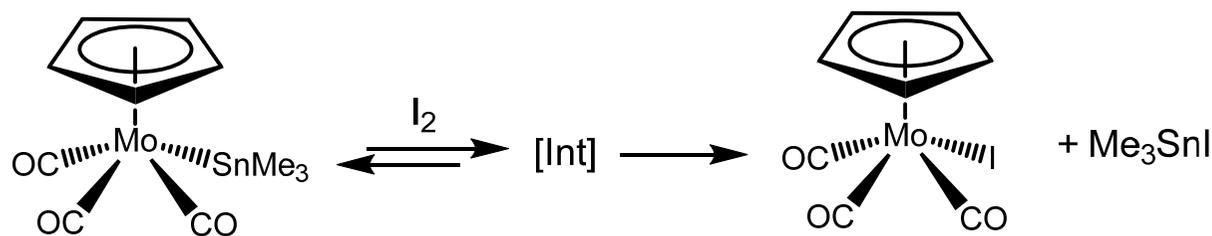
Overall Reaction and Reaction Rate

3.1 Steady state approximation

consecutive two step reaction with one reversible step



this (simplified) reaction is often decisive in chemical processes



approximation: the intermediate does not change during the reaction

$$\text{steady state: } \frac{dI}{dt} = 0 \quad [Int] \text{ is small but not } 0$$

Overall Reaction and Reaction Rate

3.1 Steady state approximation

steady state: $\frac{dI}{dt} = 0$ or $k_1[A] - k_{-1}[Int] - k_2[Int] = 0$

mit $[I]^* = \frac{k_1[A]}{k_{-1} + k_2}$ in steady state

$[I]^*$ can now be substituted in the rate law to yield

$$-\frac{dA}{dt} = \underbrace{\frac{k_1 \cdot k_2}{k_{-1} + k_2}}_{k_{\text{obs}}} [A] = \frac{dP}{dt}$$

approximation! if $\frac{dI}{dt} = 0$ then $-\frac{dA}{dt} = 0$

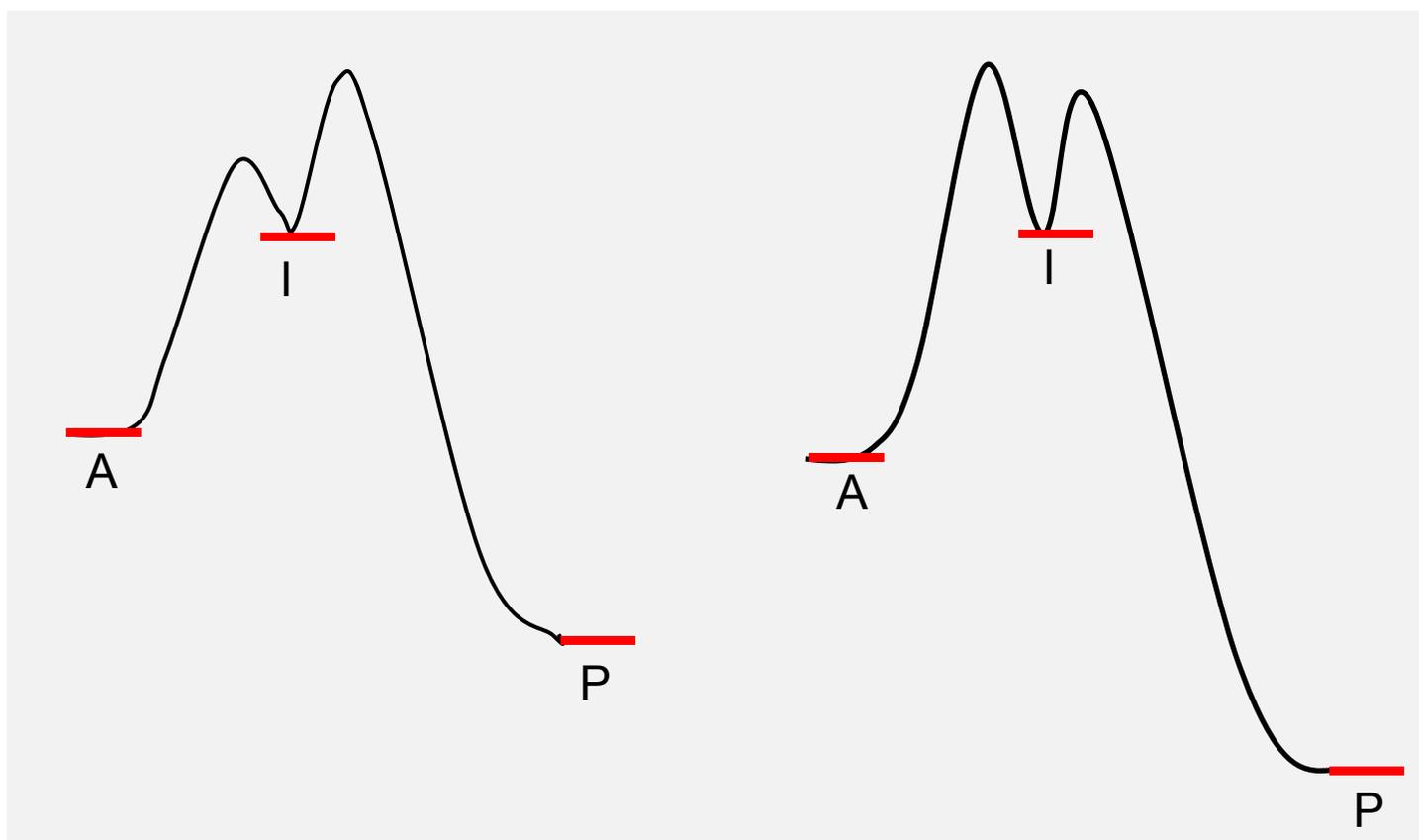
$$[I]_{\text{ss}} = \frac{k_1}{k_{-1} + k_2} [A]_0 \cdot e^{-k_{\text{ss}} \cdot t}$$

very small since $k_{-1} + k_2 \gg k_1$

Overall Reaction and Reaction Rate

3.1 Steady state approximation

Reaction profiles for a steady state approximation

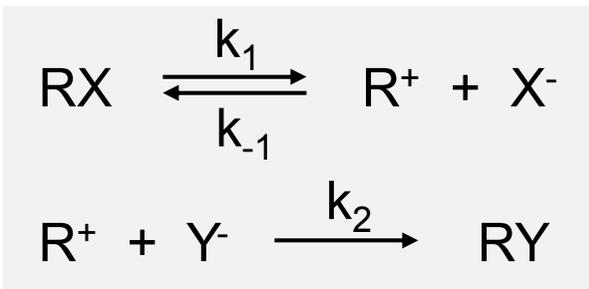


Overall Reaction and Reaction Rate

3.1 Steady state approximation

example: S_{N1} type reaction, frequent in chemistry

in inorganic chemistry, this is called a D (Dissociation) type reaction



$$\frac{d\text{R}^+}{dt} = k_1[\text{RX}] - k_{-1}[\text{R}^+][\text{X}^-] - k_2[\text{R}^+][\text{Y}^-] = 0 \quad [\text{R}^+]_{\text{ss}} = \frac{k_1[\text{RX}]}{k_{-1}[\text{X}^-] + k_2[\text{Y}^-]}$$

$$\text{rate law: } -\frac{d[\text{RX}]}{dt} = \frac{d[\text{RY}]}{dt} = \frac{k_1 \cdot k_2 [\text{RX}][\text{Y}^-]}{k_{-1}[\text{X}^-] + k_2[\text{Y}^-]}$$

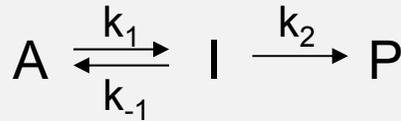
bei altering [X] or [Y], we can differentiate cases for k_1 , k_{-1} , k_2

Overall Reaction and Reaction Rate

3.1 Steady state approximation

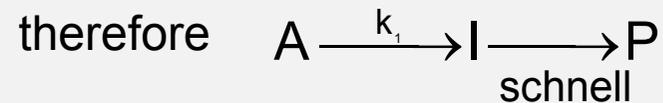
different cases

1. $k_2 \gg k_{-1}$



$$\rightarrow k_{\text{obs}} = \frac{k_1 \cdot k_2}{k_{-1} + k_2} \quad \rightarrow k_{\text{obs}} = k_1$$

k_1 becomes rate determining



if k_2 is slow in comparison to k_{-1} , we cannot use the steady state approximation but a consecutive scheme with irreversible steps

$\frac{1}{k_{\text{obs}}}$ is a measure for the life time of the reaction

1. reaction $(k_1 + k_{-1})^{-1}$
2. reaction k_2^{-1}

Overall Reaction and Reaction Rate

3.1 Steady state approximation

different cases

2. $k_2 \ll k_{-1}$



$$k_{\text{obs}} = \frac{k_1 \cdot k_2}{k_{-1} + k_2} \quad k_{\text{obs}} = \frac{k_1 \cdot k_2}{k_{-1}} = K \cdot k_2$$

k_2 becomes rate determining

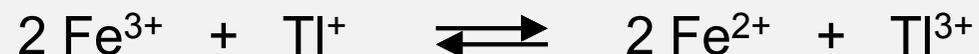
this situation is also called **rapid pre-equilibrium**

The first equilibrium is achieved at all time points during the reaction

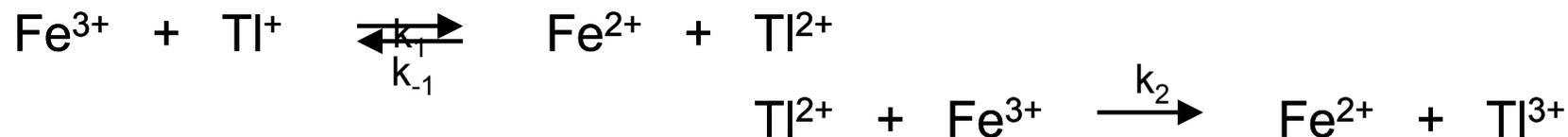
does not mean that k_1 is necessarily fast, it can be slow either but much faster than k_2

Overall Reaction and Reaction Rate

3.1 Steady state approximation



in fact, a reaction with two elementary steps



the correct **rate law** under steady state conditions is:

$$\frac{d\text{Ti}^{3+}}{dt} = \frac{k_1 \cdot k_2 \cdot [\text{Fe}^{3+}]^2 \cdot [\text{Ti}^{3+}]}{k_{-1} \cdot [\text{Fe}^{2+}] + k_2 \cdot [\text{Fe}^{3+}]}$$

since we find $v = k \cdot [\text{Fe}^{3+}] \cdot [\text{Ti}^{3+}]$

we conclude that **the second step is very fast** i.e. $k_2 \gg k_{-1}$

and $k = k_1$ is the **rate determining step**

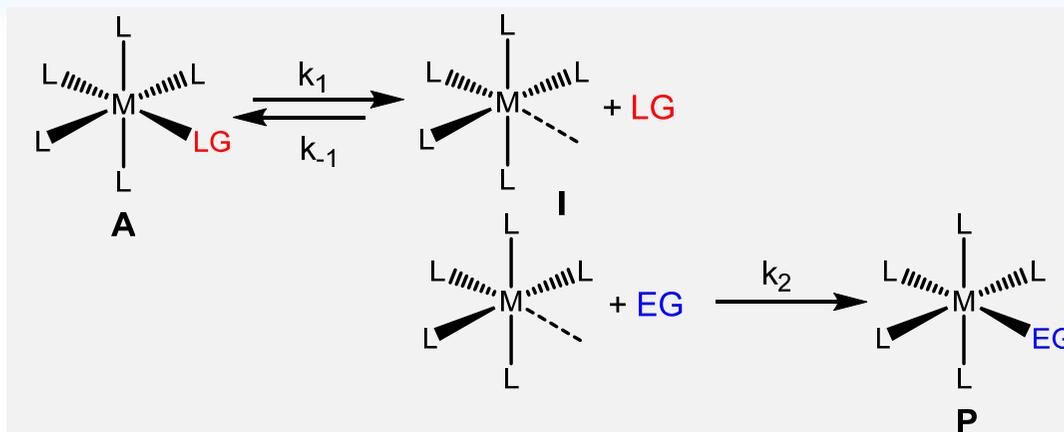
we don't "see" the steps beyond the rate determining step !!

Overall Reaction and Reaction Rate

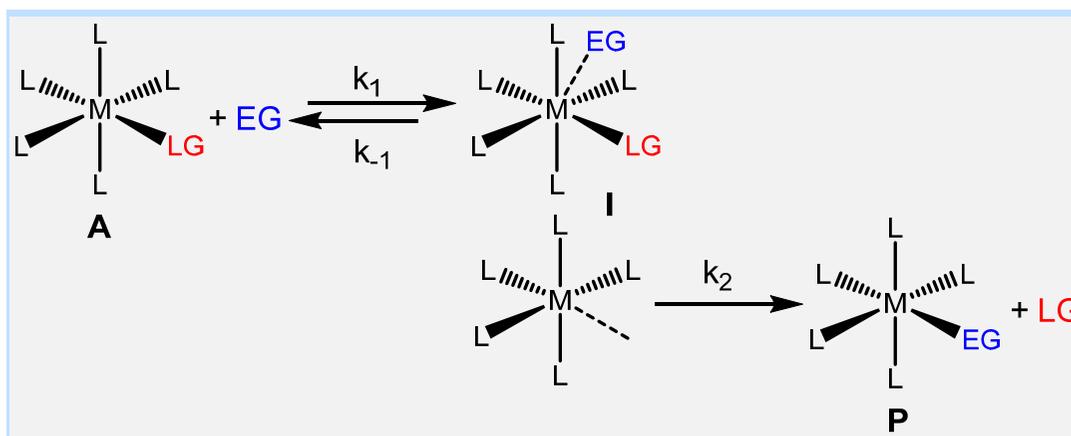
3.1 Steady state approximation

substitution mechanisms

A Associative



or D Dissociative



I intermediates and not transition states, distinguishable ?

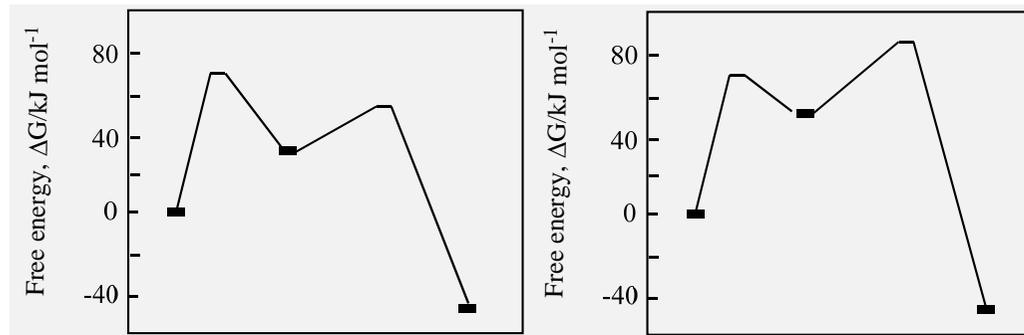
Overall Reaction and Reaction Rate

3.1 Steady state approximation

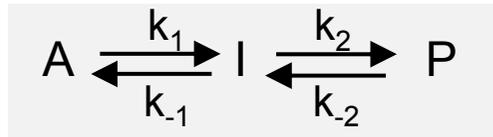
if $k_{-1} [\text{Fe}^{3+}] \ll k_2 [\text{Fe}^{2+}]$ then $\frac{d[\text{Ti}^+]}{dt} \cong k_1 [\text{Fe}^{2+}] [\text{Ti}^{3+}]$

if $k_{-1} [\text{Fe}^{3+}] \gg k_2 [\text{Fe}^{2+}]$ then $\frac{d[\text{Ti}^+]}{dt} \cong K_1 k_2 \frac{[\text{Fe}^{2+}]^2 [\text{Ti}^{3+}]}{[\text{Fe}^{3+}]}$

Darstellung als Reaktionsprofil



reaction profiles for two cases of rate constants



a: $k_1=1$; $k_{-1}=10^{-8}$; $k_2=10^{10}$; $k_{-2}=10^{-4} \text{ s}^{-1}$

b: $k_1=1$; $k_{-1}=10^{10}$; $k_2=10^8$; $k_{-2}=10^{-8} \text{ s}^{-1}$

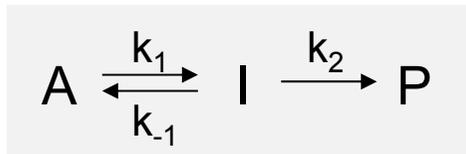
Overall Reaction and Reaction Rate

3.1 Steady state approximation

which is now the **rate determining step**?

It is **not necessarily the one with the highest transition state**

introduce a **control factor** for each elementary step



an excess rate factor is introduced for each step

$$E_1 = k_1[A] - \frac{k_1 k_2 [A]}{k_{-1} + k_2} = \frac{k_1 k_{-1} [A]}{k_{-1} + k_2} \quad \text{und} \quad E_2 = \frac{k_2 k_{-2} [P]}{k_{-1} + k_2}$$

if a step produces a **lot of excess**, it is not rate determining

Overall Reaction and Reaction Rate

3.1 Steady state approximation

which is now the **rate determining step**?

if a step produces a **lot of excess**, it is not rate determining

$$E_1 = k_1[A] + \frac{k_1k_2[A]}{k_{-1} + k_2} = \frac{k_1k_{-1}[A]}{k_{-1} + k_2} \quad \text{und} \quad E_2 = \frac{k_2k_{-2}[P]}{k_{-1} + k_2}$$

forward rate for the 1st step - overall rate: if zero or small, then rate determining

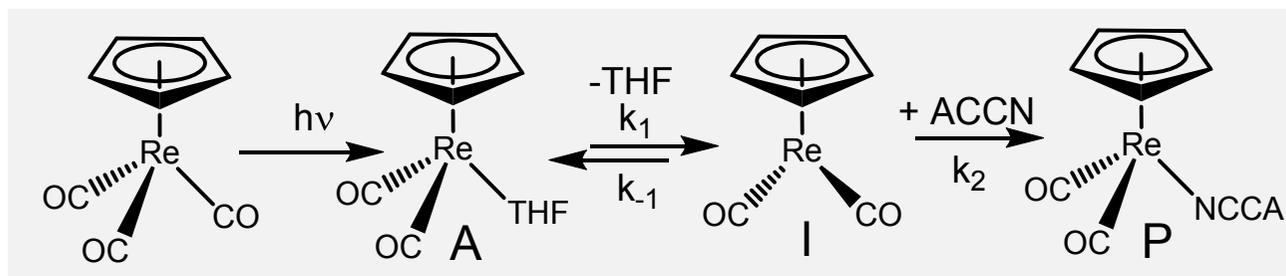
Rate Constants, s⁻¹

k ₁	k ₋₁	k ₂	Case	RCS
1	10 ⁻²	10 ²	k ₂ >> k ₁ >> k ₋₁	1
10 ⁴	10 ⁶	10 ²	k ₋₁ >> k ₁ >> k ₂	2
10 ²	10 ⁶	10 ⁴	k ₋₁ >> k ₂ >> k ₁	2
1	10 ²	10 ⁴	k ₂ >> k ₋₁ >> k ₁	1
10 ²	1	10 ⁻²	k ₁ >> k ₋₁ >> k ₂	2
1	10 ⁻⁴	10 ⁻²	k ₁ >> k ₂ >> k ₋₁	2
1	10 ²	10 ²	k ₂ = k ₋₁ >> k ₁	None
2	1	1	k ₁ > k ₋₁ = k ₂	None

Overall Reaction and Reaction Rate

Example: Typical reaction under steady state conditions

Photolytic cleavage of CO in the presence of a coordinating solvent



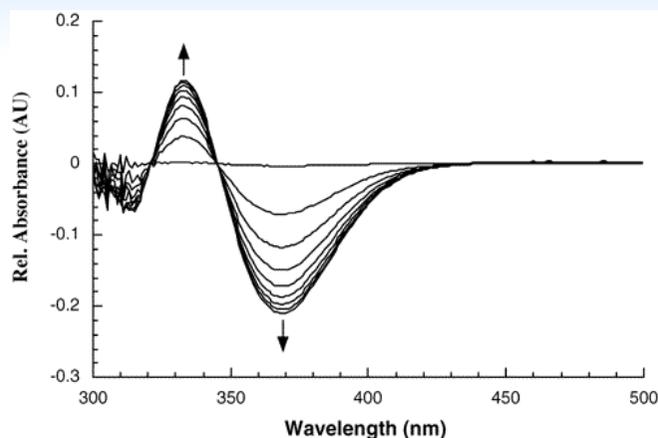
Since A and P are 18 e⁻ species, "I" should be very reactive and only present in minute concentrations

accordingly:

$$k_{\text{obs}} = \frac{k_1 \cdot k_2 [\text{CH}_3\text{CN}]}{k_{-1} [\text{THF}] + k_2 [\text{CH}_3\text{CN}]}$$

Overall Reaction and Reaction Rate

The product formation is followed by UV/vis spectroscopy



$[\text{CH}_3\text{CN}]$ and $[\text{THF}]$ are **constant** and $\gg [\text{Re}]$

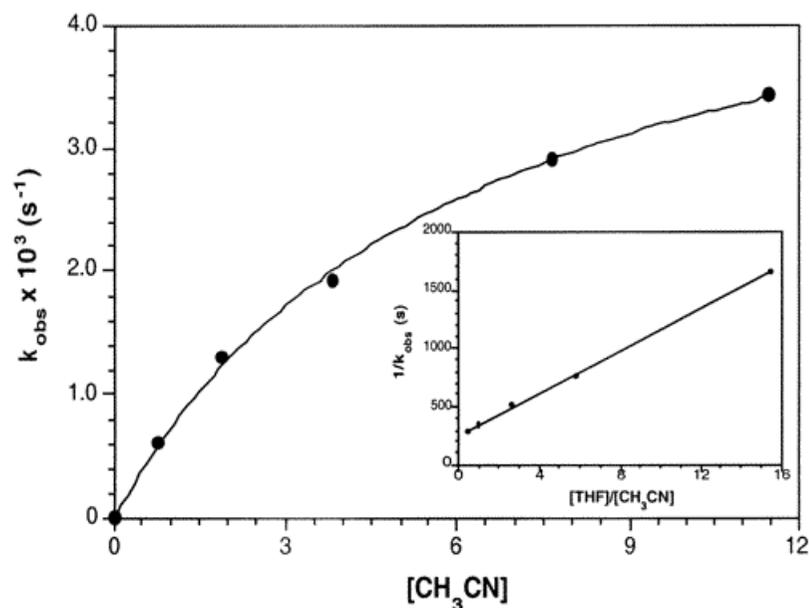
The equation can be modified to account for the solvent/ligand ratio

$$[\text{THF}] = [\text{THF}]_0 - [\text{ACCN}] \quad \text{und} \quad a = \frac{[\text{THF}]}{[\text{CH}_3\text{CN}]}$$

$$k_{\text{obs}} = \frac{k_1 \cdot k^* [\text{CH}_3\text{CN}]}{12.3 + (k^* - a)[\text{CH}_3\text{CN}]}$$

Overall Reaction and Reaction Rate

The graphical relationship k_{obs} vs. $[\text{CH}_3\text{CN}]$ allows the calculation of k_1 and k^* with the least square method.



Calculated rate constants k_1 [sec^{-1}]

$$k_1[\text{Re}]: 9.24 \pm 0.5$$

$$k_2[\text{Re}]: 2.85 \pm 0.13$$

$$k_3[\text{Re}]: 0.29 \pm 0.01$$

$$\Delta H^\ddagger = 24 \pm 0.9 \text{ kJ/mol}$$

$$\Delta S^\ddagger = 9.2 \pm 2.7 \text{ e.u.}$$

Interpretation: k_{obs} increases with $[\text{ACCN}]$, but becomes constant for large $[\text{ACCN}]$

➔ k_1 becomes rate determining

Overall Reaction and Reaction Rate

3.2 The prior equilibrium

an approximation often found in inorganic and pH dependent reactions

again a dependence of the different k's

if k_1 and or $k_{-1} \gg k_2$, we consider the first step as an equilibrium

we assume that the **equilibrium is rapidly established** and maintained over the whole reaction



$$\text{then: } I_t = \frac{k_1}{k_{-1}} [A]_t \quad \text{für } [k_{+1} + k_{-1}] \gg k_2$$

important: $[I]_t$ is not necessarily small and $\frac{dI}{dt} \neq 0$

Overall Reaction and Reaction Rate

3.2 The prior equilibrium

Dependencies: $-\frac{d[A]}{dt} - \frac{d[I]}{dt} = \frac{d[P]}{dt} = k_2[I]$ $k_{\text{obs}} = \frac{k_1 \cdot k_2}{k_{-1}} = K \cdot k_2$

ergibt: $[I]_t = \frac{k_1[A]_0}{k_1 + k_{-1}} e^{-k_{\text{pe}} t}$ mit $k_{\text{pe}} = \frac{k_1 \cdot k_2}{k_1 + k_{-1}}$

substituted and integrated: $[A] = \frac{k_{-1}[A]_0}{k_1 + k_{-1}} e^{-k_{\text{pe}} t}$ $[P]_t = [A]_0 \left\{ 1 - e^{-k_{\text{pe}} t} \right\}$

if we consider **the equilibria and the rate constants**, we can write

$$k_{\text{pe}} (= k_{\text{obs}}) = \frac{K \cdot k_2}{(1 + K)}$$

is frequently found in ligand exchange reactions

Overall Reaction and Reaction Rate

3.2 The prior equilibrium

Dissociative and associative complex formation reactions

- ➔ Complexation / decomplexation
- ➔ Ion Pair formations,
- ➔ charge transfer complex formations

Complex formation in a prior-equilibrium situation



$$\frac{dP}{dt} = k_2[AL]$$

we consider the **distribution coefficients** as a function of K_L

Overall Reaction and Reaction Rate

3.2 The prior equilibrium



$$f_{AL} = \frac{[AL]}{[AL] + [A]} \quad K_L = \frac{[AL]}{[A] \cdot [L]} \quad [A] = \frac{[AL]}{[K] \cdot [L]} \quad \text{and} \quad f_{AL} = \frac{K \cdot [L]}{1 + K \cdot [L]}$$

$$[AL]_t = f_{AL}([A] + [AL]) = f_{AL}[A]_{\text{tot}} = f_{AL}([A_0] - [P]_t)$$

$$\frac{dP}{dt} = k_2[AL] = \frac{k_2 \cdot K \cdot [L] \cdot ([A_0] - [P]_t)}{1 + K \cdot [L]}$$

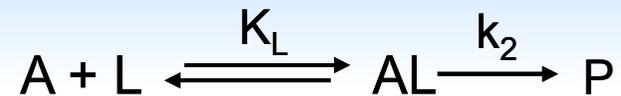
if $[L]$ is constant, the distribution coefficient f_{AL} and f_A is constant over the reaction

if we keep $[L]$ constant and follow $[P]$ we can get K and k_2

K: a thermodynamic parameter from a kinetic measurement !

Overall Reaction and Reaction Rate

3.2 The prior equilibrium



$$f_A = \frac{[A]}{[A]_T} = \frac{1}{1+K \cdot [L]} \quad \text{resp.} \quad f_{AL} = \frac{K \cdot [L]}{1+K \cdot [L]}$$

the same principle can be followed if A is the reactive compound



kinetic observations

reaction is **pseudo first order** at constant $[\text{SO}_4]^{2-}$ concentration $\gg [\text{Co}]$

reaction rate **remains constant** at high $[\text{SO}_4]^{2-}$ concentration

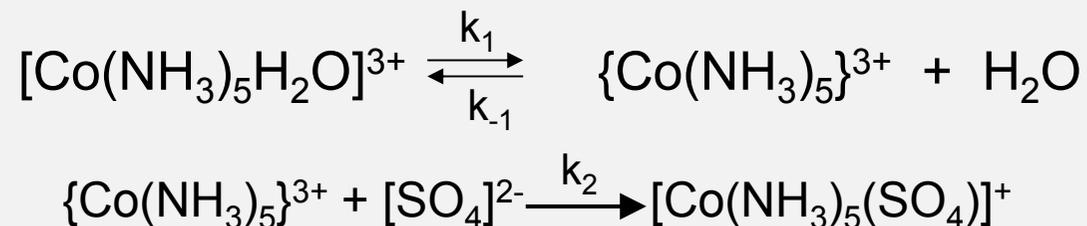
Overall Reaction and Reaction Rate

3.2 The prior equilibrium



rate constant:
$$k_{\text{obs}} = \frac{k[\text{SO}_4^{2-}]}{1 + k'[\text{SO}_4^{2-}]}$$

looks like a **D mechanism** (or $\text{S}_{\text{N}1}$ if you wish)



k becomes $\frac{k_1 \cdot k_2}{k_{-1}}$ and $k' = \frac{k_2}{k_{-1}}$

Overall Reaction and Reaction Rate

3.2 The prior equilibrium

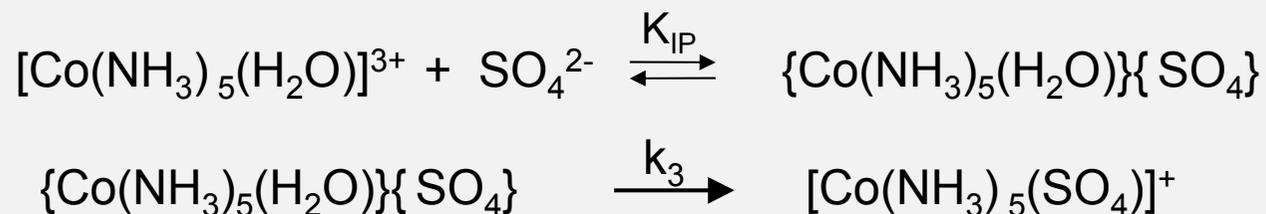


k becomes $\frac{k_1 \cdot k_2}{k_{-1}}$ and $k' = \frac{k_2}{k_{-1}}$

→ the plateau correspond to k_1 which is then rate determining

this is a typical **D mechanism** (Langford-Gray)

or



then, rate constant:

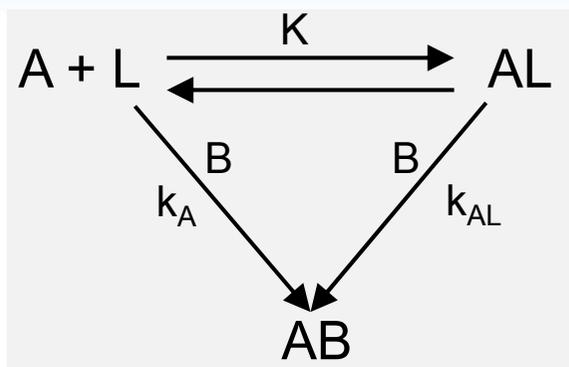
$$k_{\text{obs}} = \frac{k_3 K_{\text{IP}} [\text{SO}_4^{2-}]}{1 + K_{\text{IP}} [\text{SO}_4^{2-}]}$$

k_3 is now rate limiting (plateau)

Overall Reaction and Reaction Rate

3.2 The prior equilibrium

Typical reaction scheme for substitution reactions (see also later)



$$f_A = \frac{[A]}{[A]_{\text{tot}}} = \frac{1}{1+K \cdot [L]} \quad f_{AL} = \frac{[AL]}{[A]_{\text{tot}}} = \frac{K \cdot [L]}{1+K \cdot [L]}$$

reaction goes **directly** with k_a or with k_{AL} via a coupled intermediate (solvent complex)

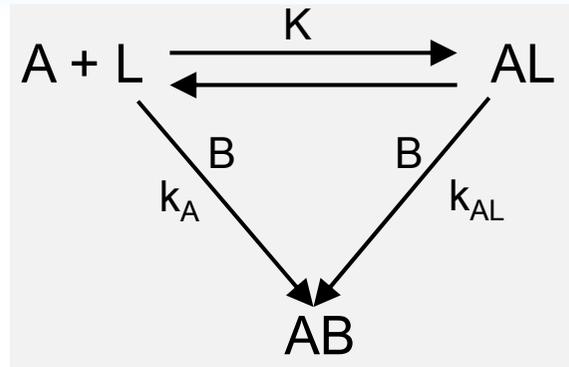
$$\frac{d[AB]}{d[t]} = k_a \cdot [A] \cdot [B] + k_{AL}[AL] \cdot [B] = \left(\frac{k_A + k_{AL}K[L]}{1 + K[L]} \right) [A]_T [B]$$

[A] and [AL] can be expressed with f_A and f_{AL}

Overall Reaction and Reaction Rate

3.2 The prior equilibrium

Typical reaction scheme for substitution reactions (see also later)



$$\frac{d[AB]}{d[t]} = k_a \cdot [A] \cdot [B] + k_{AL}[AL] \cdot [B] = \left(\frac{k_A + k_{AL}K[L]}{1 + K[L]} \right) [A]_T [B]$$

1. case: $1 \gg K[L]$ then $(k_A + k_{AL} \cdot K \cdot [L]) \cdot [A]_{tot} \cdot [B]$ (AL very small)

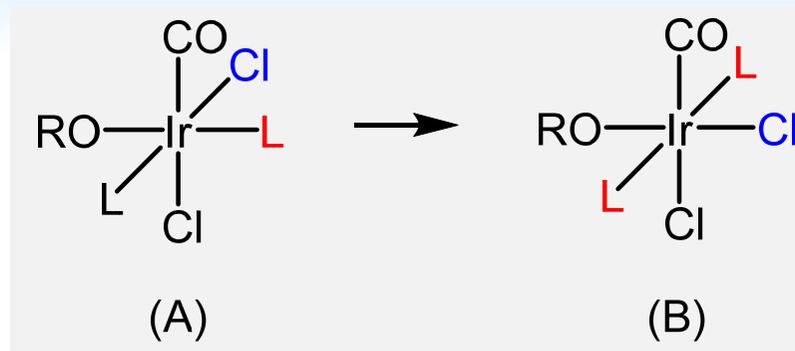
2. case: equilibrium but only AL reactive

$$v = \frac{k_{AL}K[L]}{1 + K[L]} [A]_T [B]$$

3. case:

Overall Reaction and Reaction Rate

Interesting example at the end: **an isomerization**



we observe that isomerization becomes slower upon the addition of excess "L"

$$\text{rate law: } \frac{dB}{dt} = \frac{k[A]}{1 + k'[L]}$$

generally, isomerisation's become faster upon addition of one of the components

which mechanism explains this rate law?

Overall Reaction and Reaction Rate

Practical kinetic example: The Monsanto Acetic Acid Process

The synthesis of acetic acid from methanol is probably the most important industrial process in homogenous catalysis

worldwide requirements are about $5.5 \cdot 10^6$ t/a

Monsanto process invented 1966, before

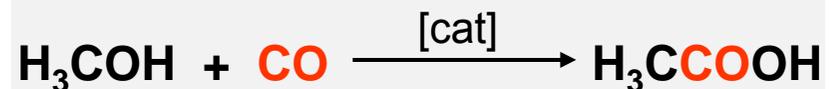
⇒ oxidation of ethanol

⇒ oxidation of acetaldehyde

⇒ naphta oxidation

⇒ Carbonylation of methanol (Co-catalyst)

⇒ Carbonylation of methanol (Rh-catalyst)

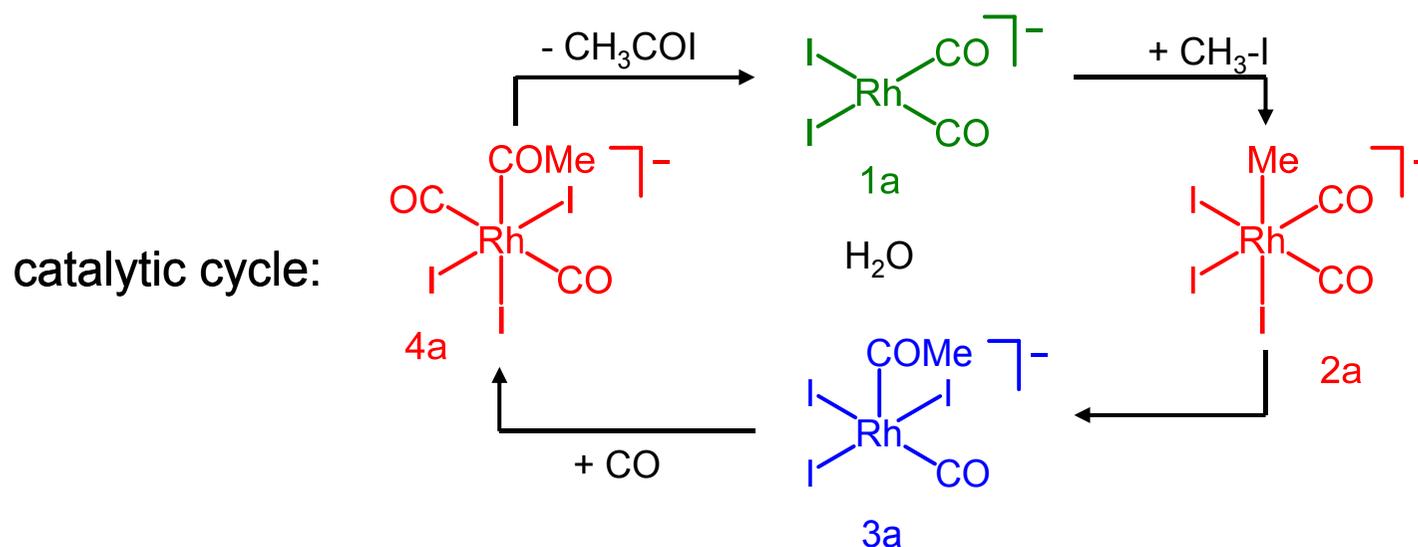


Overall Reaction and Reaction Rate

Practical kinetic example: The Monsanto Acetic Acid Process

Comparison of the processes

	selectivity	T/°C	p/atm
Naphta oxidation BP	65-70	185	48
MeOH / CO (Co cat)	90	230	600
MeOH / CO (Rh cat)	99	150-200	30-60
MeOH / MeOAc carbonylation (Rh cat)	high	150-200	30-50



6 (!) reactions are interlinked to form one closed working catalytic system

Overall Reaction and Reaction Rate

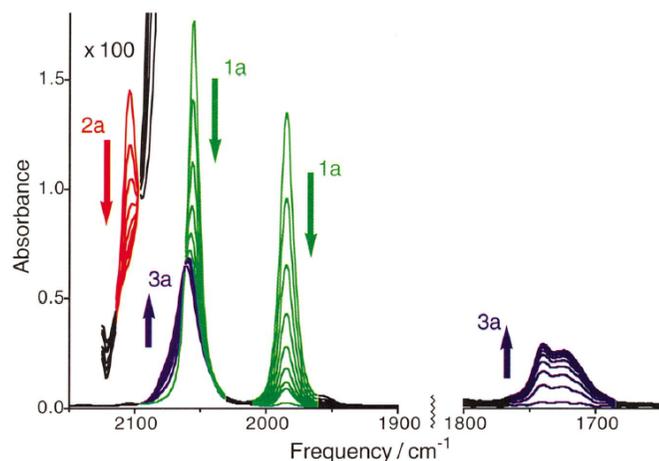
Practical kinetic example: The Monsanto Acetic Acid Process

Many studies focused on the rds of the cycle, namely

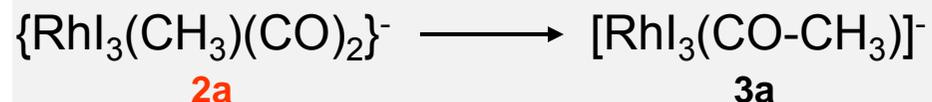


→ polar solvents and addition of I⁻ accelerate the reaction

→ {RhI₃(CH₃)(CO)₂}⁻ (2a) difficult to detect spectroscopically



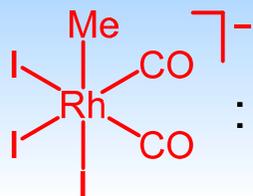
The next step, CO insertion, is fast



very careful analysis of the i.r. spectra

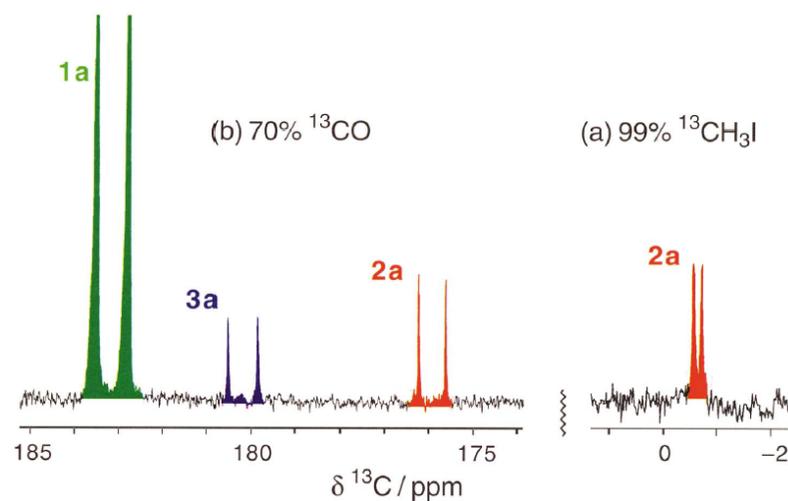
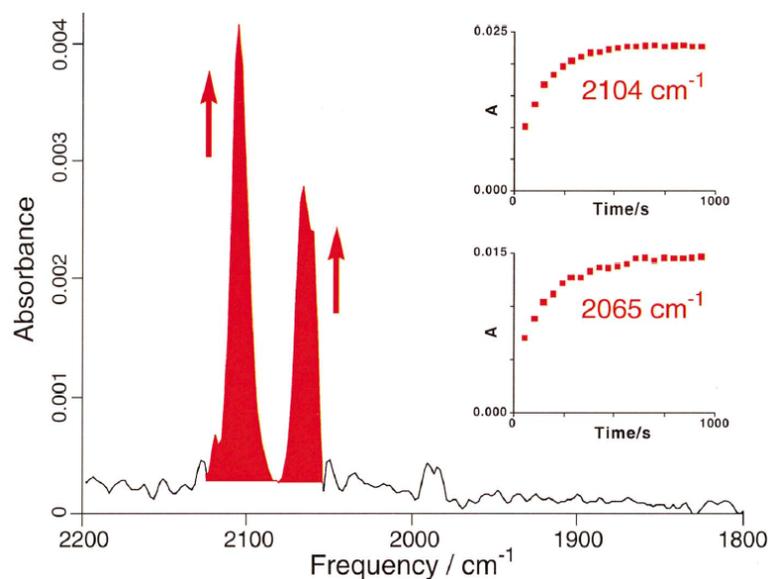
→ 2a and 3a become „visible“

Overall Reaction and Reaction Rate



: increases and decreases with **1a**

2a



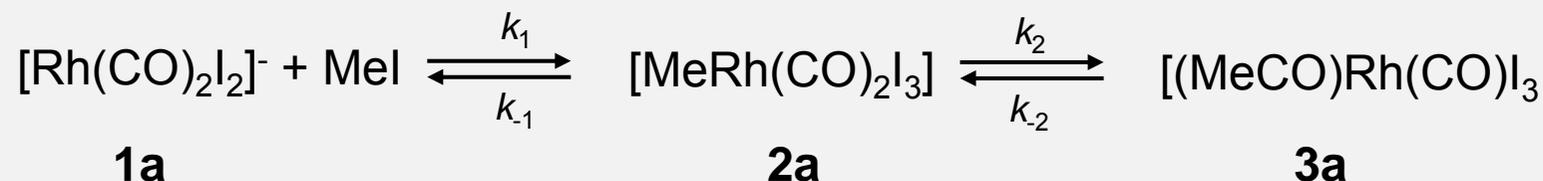
^1H and ^{13}C NMR in deep cold solution of $\text{CH}_3\text{-I}$ confirm the finding from i.r.

i.r. pattern discloses a *fac,cis*-dicarbonyl structure

The detection of **2a** allows for determination of the individual steps

Overall Reaction and Reaction Rate

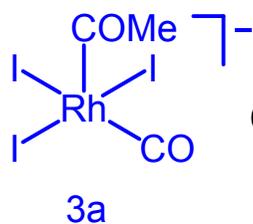
Practical kinetic example: The Monsanto Acetic Acid Process



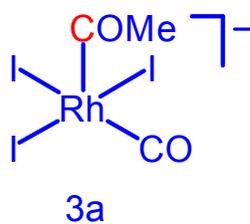
rate constant: $\frac{k_1 \cdot k_2 \cdot [\text{MeI}]}{k_1 + k_2}$

ratio of concentrations: $\frac{[\mathbf{2a}]}{[\mathbf{1a}]} = \frac{k_1 \cdot [\text{MeI}]}{k_1 + k_2} \quad \Rightarrow \quad k_2 = k_{\text{obs}} \cdot \frac{[\mathbf{1a}]}{[\mathbf{2a}]}$

$\frac{[\mathbf{1a}]}{[\mathbf{2a}]}$ can be estimated from i.r. intensities \Rightarrow which allows for the **determination of k_2**



can be isolated

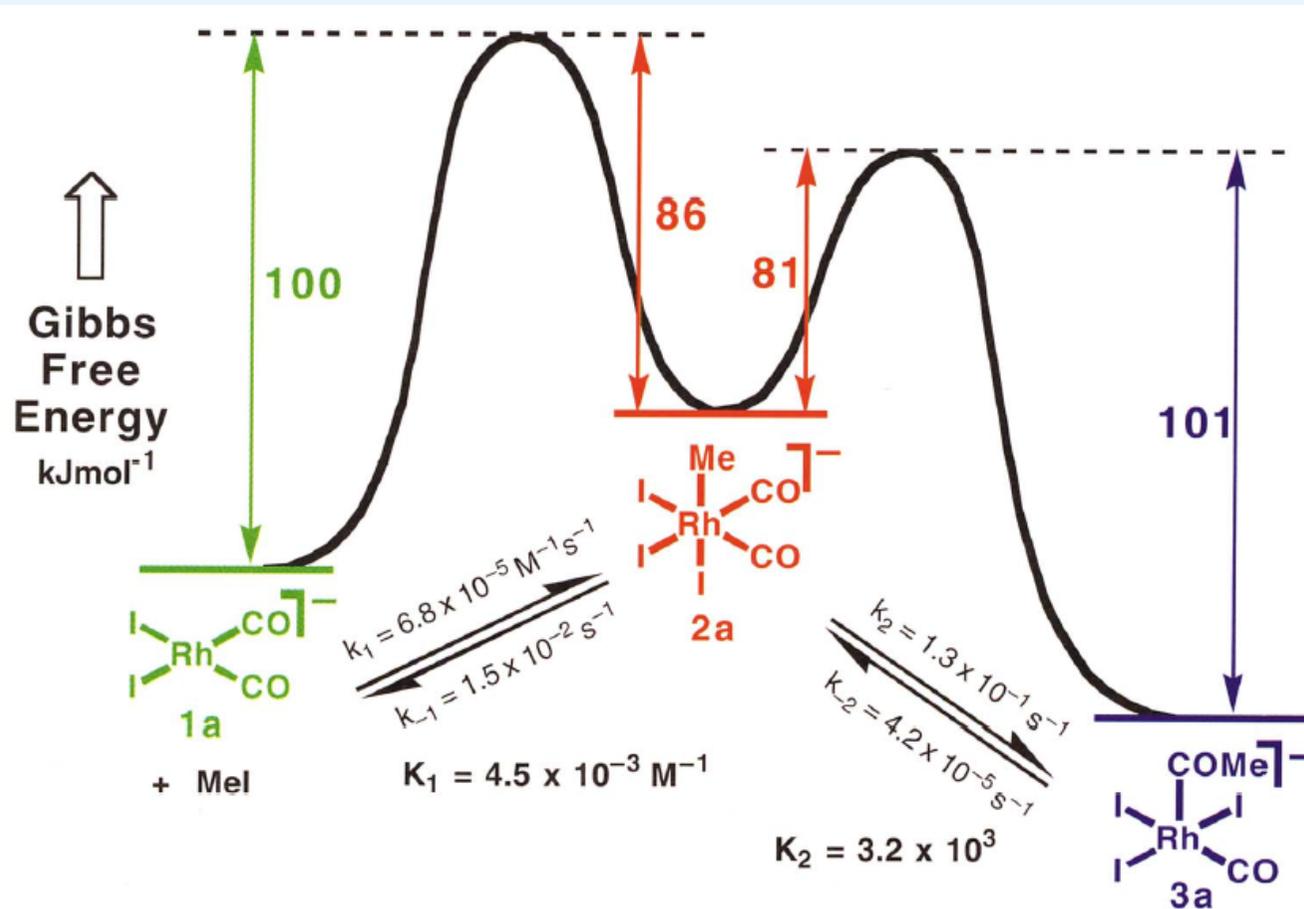


heating in the presence of exc. $\text{H}_3\text{C-I}$ allows for the determination of k_2

label exchanges between COCH_3 and CO

Overall Reaction and Reaction Rate

Practical kinetic example: The Monsanto Acetic Acid Process

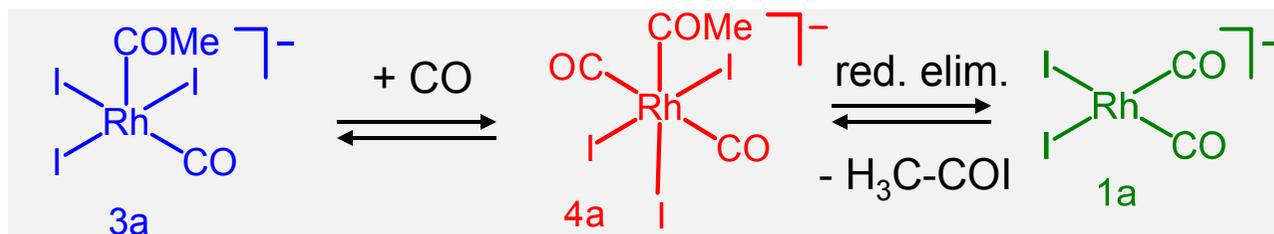


from these experiments, a **complete set of data** could be deduced

Overall Reaction and Reaction Rate

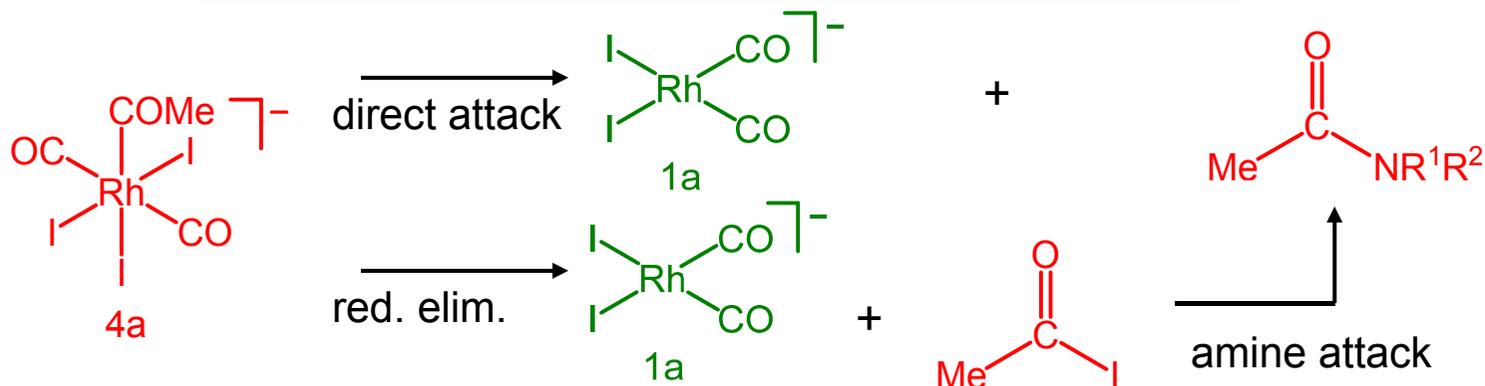
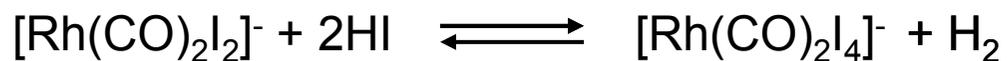
Practical kinetic example: The Monsanto Acetic Acid Process

what about the remaining processes leading to acetic acid ?



⇒ 3a is immediately trapped by CO

⇒ 4a eliminates H₃C-COI rather slowly ($t_{1/2} = 12\text{h}$ at r.t.)
 difficulty to measure accurate kinetics (due to traces of water)



Overall Reaction and Reaction Rate

3.3 Enzyme catalysed reactions and catalysis

a catalyst does not necessarily reduce E_a

➔ but a catalyst induces a different reaction pathway, profile

➔ a catalyst exists at least in two different forms

as free catalyst and ..

.. in interaction with **substrate(s)**

➔ a catalyst circulates between inactive and active form

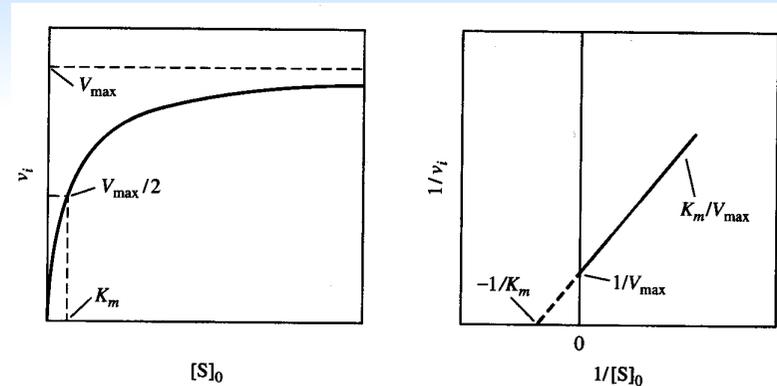
➔ a **good catalyst** is present in very low concentration and ..

.. intermediates are not observable

➔ if $[S] \gg [cat]$, the **reaction is 0th order in [cat]** since it is fully active (saturation)

Overall Reaction and Reaction Rate

3.3 Enzyme catalysed reactions and catalysis



enzymes: concentration vs. rate of an enzyme

at low [S], reaction is 1st order in [S]: **Michaelis Menten Kinetics**

$$v = \frac{V_{\max} [S]}{K_m + [S]} \quad \text{M-M Gleichung}$$

K_m : [S] at which $v = v_{\max}/2$, i.e. the smaller K_m the more efficient the catalyst (enzyme)

K_m is called M-M constant [dim=M]

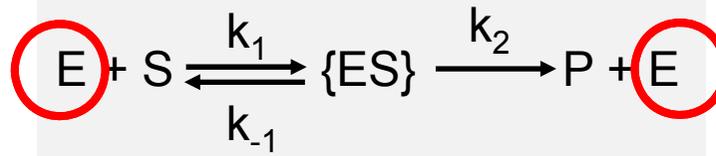
Overall Reaction and Reaction Rate

3.3 Enzyme catalysed reactions and catalysis

linearization:
$$\frac{1}{v_i} = \frac{1}{v_{\max}} + \frac{K_m}{v_{\max}[S_0]}$$

v_i = initial rate as a function of $[S_0]$ \Rightarrow allows now calculation of v_{\max} and K_m

how to get rate constants from these numbers, mechanism behind?

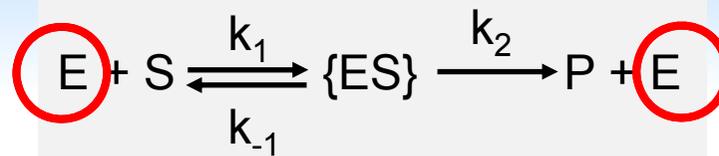


known scheme but this one is catalytic!

$$\text{initial rat } v_i = k_2 \cdot \{\text{ES}\}$$

Overall Reaction and Reaction Rate

3.3 Enzyme catalysed reactions and catalysis



steady state approximation: $\frac{d[\text{ES}]}{dt} = 0 = -k[\text{ES}] - k_{-1}[\text{ES}] + k_1[\text{S}]_0 \cdot [\text{E}]$

$$\Rightarrow [\text{ES}] = \frac{k_1[\text{E}][\text{S}]_0}{k_{-1} + k_2}$$

catalytic condition: mit $[\text{E}] = [\text{E}]_0 - [\text{ES}] = \text{constant}$

$$[\text{ES}]_{\text{ss}} = \frac{[\text{E}]_0 \cdot [\text{S}]_0}{\frac{k_{-1} + k_2}{k_1} + [\text{S}]_0} \quad \text{gives an initial rate} \quad v_i = \frac{k_2 \cdot [\text{E}]_0 \cdot [\text{S}]_0}{\frac{k_{-1} + k_2}{k_1} + [\text{S}]_0} = \frac{d\text{P}}{dt}$$

Overall Reaction and Reaction Rate

3.3 Enzyme catalysed reactions and catalysis

$$v_i = \frac{k_2 \cdot [E_0] \cdot [S_0]}{\frac{k_{-1} + k_2}{k_1} + [S_0]}$$

comparison with M-M equation shows

$$v_{\max} = k_2[E]_0$$

and

$$K_m = \frac{k_{-1} + k_2}{k_1}$$

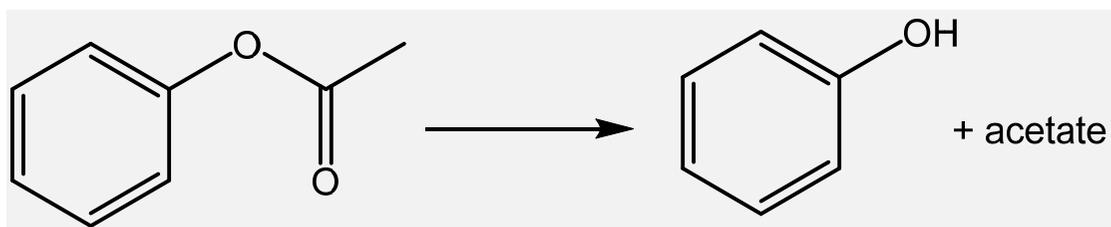
→ k_2 is also assigned as k_{cat}

the system is reminiscent to chemical reaction with
the difference that a catalyst **is not used up!**

Overall Reaction and Reaction Rate

3.3 Enzyme catalysed reactions and catalysis

Example: Hydrolysis of phenyl-acetate by acetyl-cholin esterase



the following v_i 's have been determined

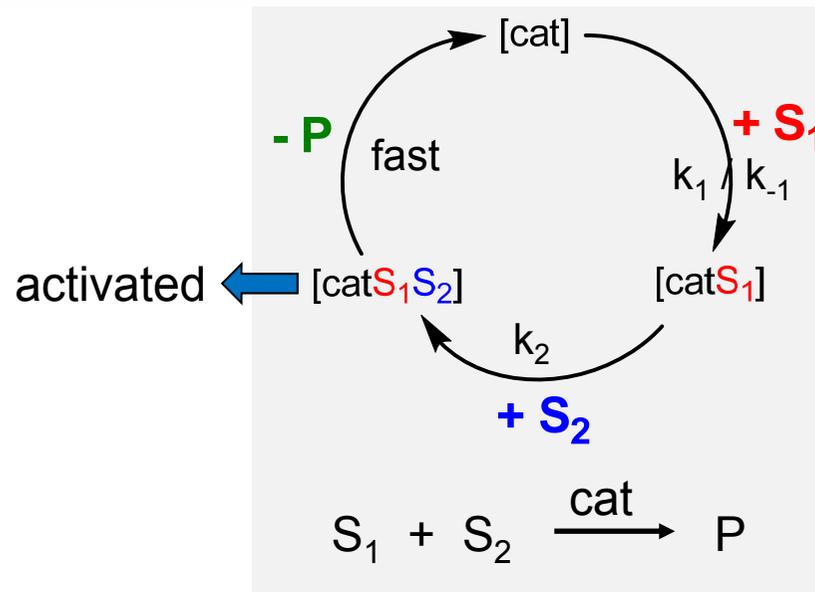
[PhOAc], 10^{-4} M:	22.5	16.3	12.6	9.73	8.45	7.18	6.21	5.55
v_i , $\cdot 10^{-7}$ l \cdot mol $^{-1}$ s $^{-1}$:	7.24	6.35	5.32	4.61	4.53	4.02	3.38	3.26

calculate K_m and k_{cat} !

Overall Reaction and Reaction Rate

3.3 Enzyme catalysed reactions and catalysis

in chemical catalysis: catalysis happens generally with two (or more) compounds



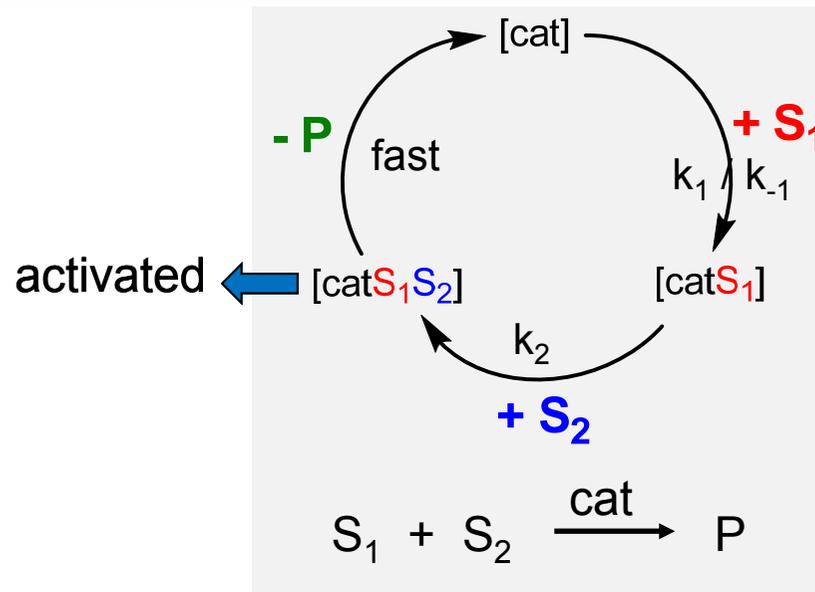
apply the steady state approach as in M-M kinetics

$$[catS_1]_{ss} = \frac{k_1 \cdot [cat] \cdot [S_1]}{k_{-1} + k_2[S_2]} \quad [cat]_0 = [cat] + [catS_1]$$

Overall Reaction and Reaction Rate

3.3 Enzyme catalysed reactions and catalysis

in chemical catalysis: catalysis happens generally with two (or more) compounds



replacement etc. yields:

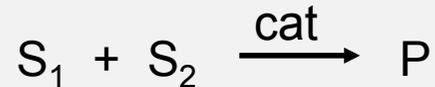
$$v_i = \frac{k_2 \cdot [E_0] \cdot [S_1][S_2]}{\frac{k_{-1}}{k_1} + \frac{k_2[S_2]}{k_1} + [S_1]}$$

$$\frac{k_{-1}}{k_1} + \frac{k_2[S_2]}{k_1} = K'_m$$

Overall Reaction and Reaction Rate

3.3 Enzyme catalysed reactions and catalysis

in chemical catalysis: catalysis happens generally with two (or more) compounds



replacement etc. yields:

$$v_i = \frac{k_2 \cdot [E_0] \cdot [S_1][S_2]}{\frac{k_{-1}}{k_1} + \frac{k_2[S_2]}{k_1} + [S_1]}$$

$$\frac{k_{-1}}{k_1} + \frac{k_2[S_2]}{k_1} = K'_m$$

initial rate approach (S_1 and S_2 constant) allows **determining k_2**

K'_m is now not constant but S_2 dependent

determination of K'_m as a function of S_2 allows determination of k_1 and k_{-1}

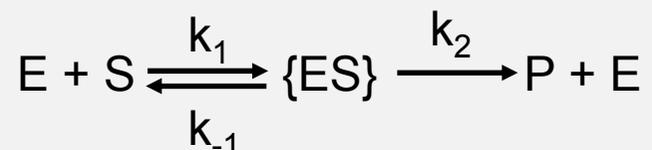
→ all kinetic parameters are know!!

what if S_2 enters first, distinguishable?

Overall Reaction and Reaction Rate

3.4 Competitive Inhibitors

Compounds may **inhibit the catalyst**: Important tool in pharmaceutical development



extended mass balance: $[E] = [E]_0 - [ES] - [EI]$ $v_i = \text{initial rate}$

$$v_i = \frac{k_2 \cdot [E_0] \cdot [S_0]}{K_m \cdot \left[1 + \frac{[I]}{K_i} \right] + [S_0]}$$

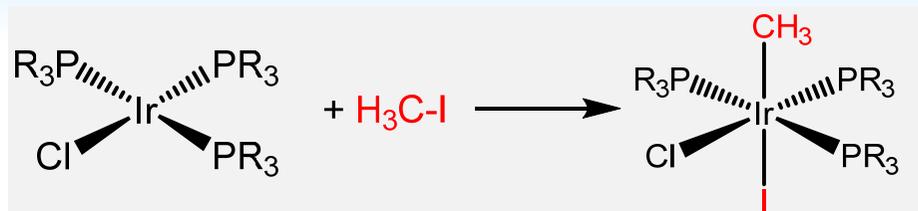
alteration of $[I]$ gives K_i , how is that done?

systematic alteration of catalyst allows determination of electronic / steric influence

Overall Reaction and Reaction Rate

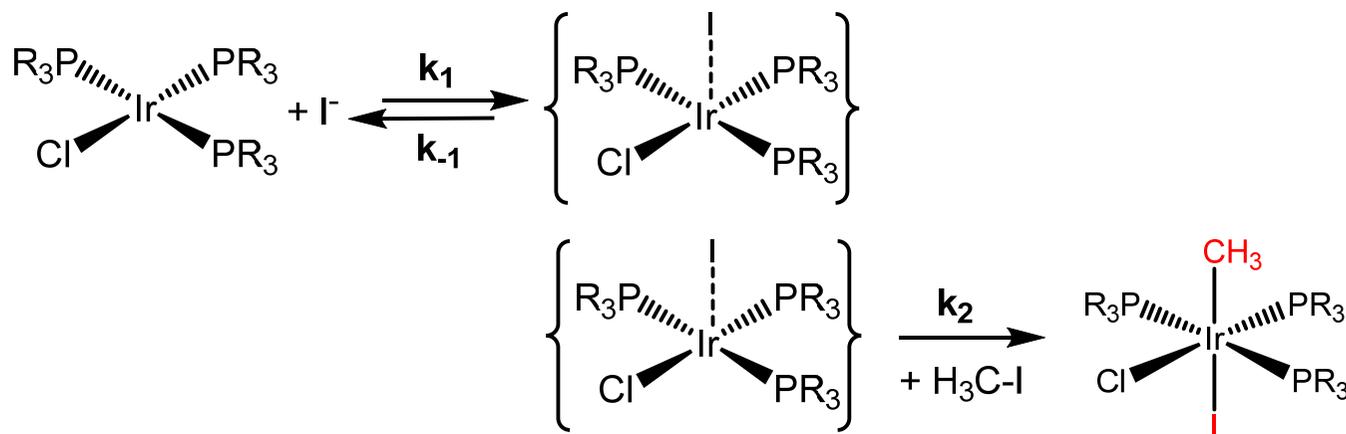
graphical evaluation in the steady state (and pe) schemes

example:



additional iodide accelerates the reaction

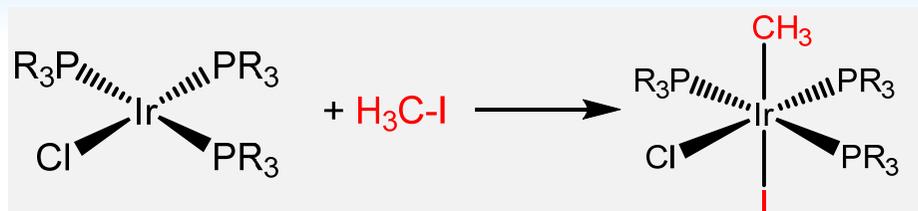
hypothesis 1:



Overall Reaction and Reaction Rate

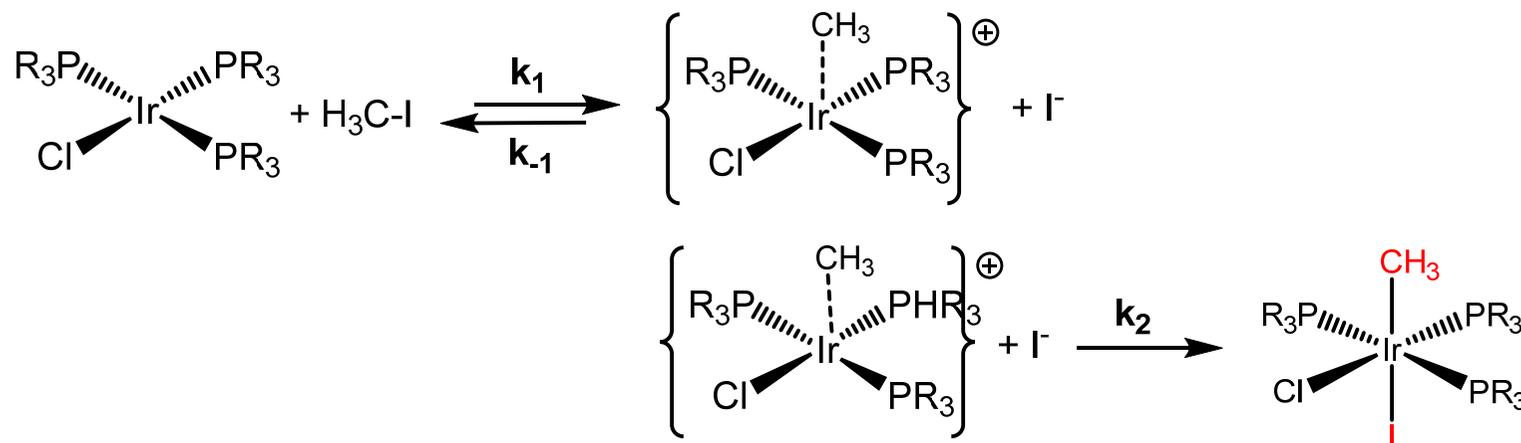
graphical evaluation in the steady state (and pe) schemes

example:



additional iodide accelerates the reaction

hypothesis 2:



➡ which one is correct, can they be distinguished ?

Overall Reaction and Reaction Rate

graphical evaluation in the steady state (and pe) schemes

"magic" formula in a steady state reaction scheme

$$[\text{intermediate}]_{\text{ss}} = \frac{\sum \text{all reactions producing [intermediate]}}{\sum \text{all pseudo 1}^{\text{st}} \text{ order reactions consuming [intermediate]}}$$

graphical analysis often by Lineweaver-Burk plots

many reaction rates in these schemes

go to saturation

under pseudo 1st order conditions

4. Deduction of Mechanisms

The path from rate laws to mechanisms is not routine (see p55)

- ➡ experimental results lead to **concentration dependencies**
- ➡ this leads to a reaction scheme which must **coincide with the rate law**
- ➡ experiments must be design which **exclude** equal possibilities
- ➡ concentration dependencies allow deduction of **intermediates**



dependencies: $k \cdot \frac{[\text{Tl}^{3+}][\text{Hg}_2^{2+}]}{[\text{Hg}^{2+}]}$

intermediate for ss: $[\text{Tl}]^{3+} + [\text{Hg}_2^{2+}] - [\text{Hg}^{2+}]$ oder $[\text{TlHg}^{3+}]$

➡ Mechanism ?

4. Deduction of Mechanisms



dependency: $[\text{S}_2\text{O}_4^{2-}]^{1/2}$

intermediate: $[\text{Mn}(\text{Mb})(\text{SO}_2)] \rightarrow$ cleavage of $\text{S}_2\text{O}_4^{2-}$

rate determining step: reaction of $[\text{SO}_2]^-$ with $[\text{Mn}^{\text{III}}(\text{Mb})]$

positive terms in rate law: forward reaction steps

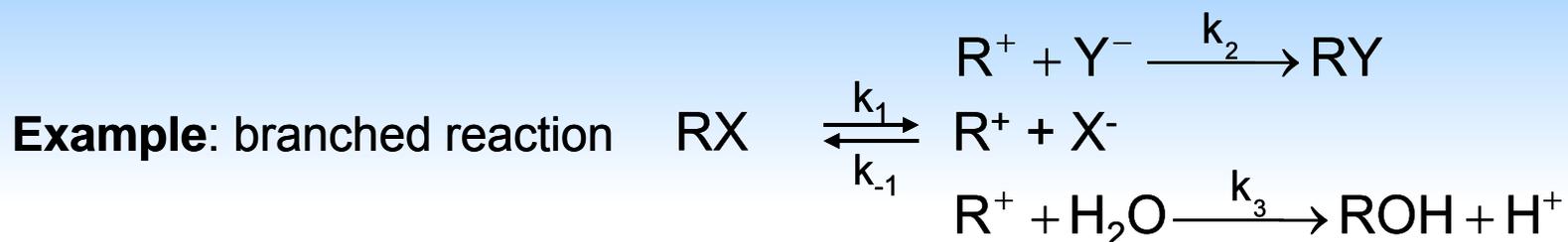
negative terms: reverse steps



$$v = k_1[\text{I}^-][\text{H}_2\text{O}_2] + k_2[\text{I}^-][\text{H}_2\text{O}_2][\text{H}^+]$$

two parallel reaction path with $[(\text{H}_2\text{O}_2)(\text{I}^-)]^\ddagger$ und $[(\text{H}_3\text{O}_2^+)(\text{I}^-)]^\ddagger$

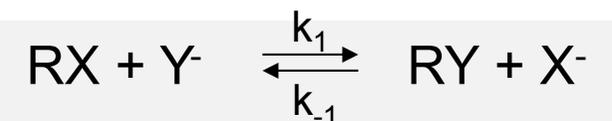
4. Deduction of Mechanisms



$$-\frac{d[\text{RX}]}{dt} = k_1[\text{RX}] \cdot \frac{k_2[\text{Y}^-] + k_3}{k_{-1}[\text{X}^-] + k_2[\text{Y}^-] + k_3}$$

The rate law shows what happens in **the rate determining step**

The **S_N2** comprises **shows reversibility**



$$v = k_1[\text{RX}][\text{Y}^-] - k_{-1}[\text{RY}][\text{X}^-]$$

➔ thus, a **negative term** arises

How does the S_N2 rate law looks for silicon ?

4. Deduction of Mechanisms

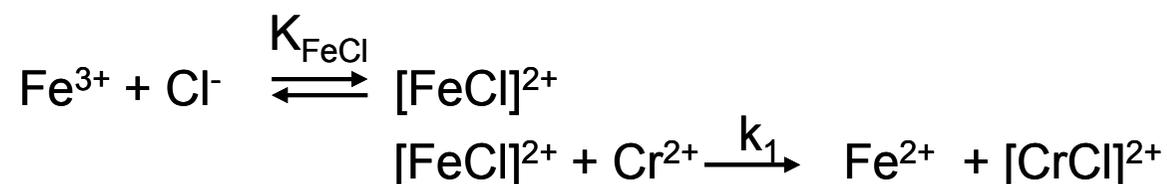
Example: The reaction of Cr^{2+} with Fe^{3+} in hydrochloric acid

$$v = \frac{k_1[\text{Cr}^{2+}][\text{Fe}^{3+}]}{[\text{H}^+]} + k_2[\text{Cr}^{2+}][\text{Fe}^{3+}][\text{Cl}^-]$$

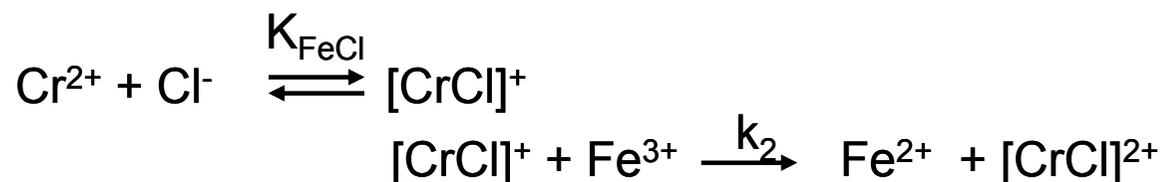
1st term: direct electron transfer reaction between Cr^{2+} and Fe^{3+} , but pH dependent

2nd term: chloride mediated reaction but no pH dependency

The mechanism of the 2nd term is **not unambiguously** clear since



or



→ rate laws ?

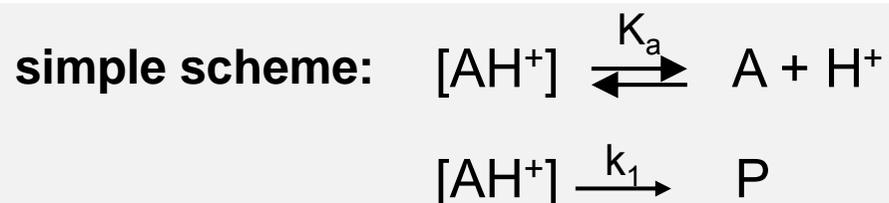
→ how to differentiate ?

Deduction of Mechanisms

4.1 pH dependencies

Many reactions **are pH dependent**, although $[H^+]$ or $[OH^-]$ do not show up in the equation

➔ only protonated or deprotonated form are active (see previous page)



here, only **the protonated form is active**

the amount of active species is defined by the pK_a

We think in distribution coefficients as in the case of complex formation

$$f_{AH} = \frac{[AH^+]}{[AH^+] + [A]} = \frac{[H^+]}{K_a + [H^+]} \quad \text{buffers !!}$$

if $[H^+]$ is large, $f_{HA} = 1$ and **everything is present as HA**

Deduction of Mechanisms

4.1 pH dependencies

$$f_{\text{AH}} = \frac{[\text{AH}^+]}{[\text{AH}^+] + [\text{A}]} = \frac{[\text{H}^+]}{K_a + [\text{H}^+]}$$

"HA" active

$$f_{\text{A}} = \frac{[\text{A}]}{[\text{AH}^+] + [\text{A}]} = \frac{K_a}{K_a + [\text{H}^+]}$$

deprotonated form "A" is active:

respective rate law: $\Rightarrow \frac{d[\text{P}]}{dt} = k_1[\text{HA}^+] = k_1 \cdot [\text{HA}]_{\text{T}}^* \cdot f_{\text{HA}}$ for HA active

gives for k_{obs} in the simplest case $\Rightarrow k_{\text{obs}} = \frac{k_1[\text{H}^+]}{K_a + [\text{H}^+]}$ $[\text{H}^+] = \text{konst.}$

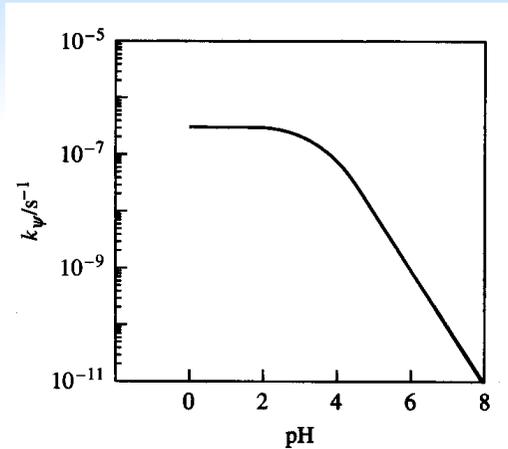
reaction **pseudo 1st order** if we can follow A or HA

k_1 und K_a can be determined experimentally

* $[\text{A}]_{\text{T}}$ = total concentration of particles with A (including HA) at time t

Deduction of Mechanisms

4.1 pH dependencies

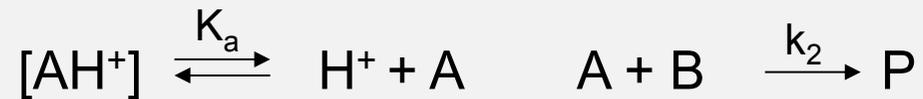


← if only HA is active

how does the curve look like if only "A" is active ?

determination of $\text{p}K_a$?

rate law if the conjugated base is the active form



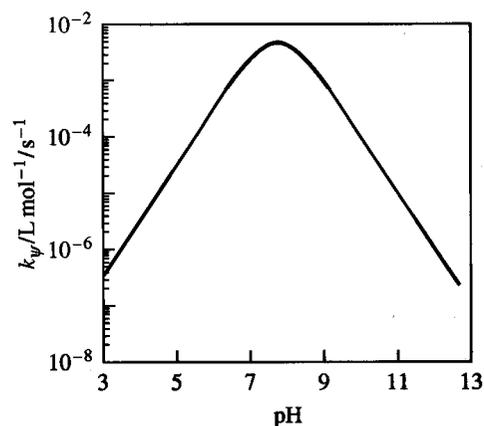
then $\rightarrow k_{\text{obs}} = \frac{k_2 K_a [\text{H}^+]}{K_a + [\text{H}^+]}$

Deduction of Mechanisms

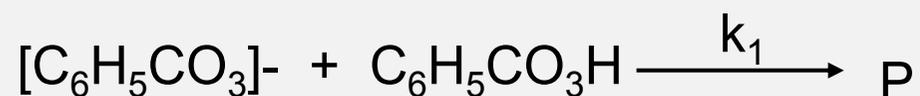
4.1 pH dependencies

if both "A" and "HA" are active as frequently found in real systems

→ pH vs. k_{obs} gives a curve with a maximum



example: decomposition of perbenzoic acid



With $[\text{C}] = [\text{HA}] + [\text{A}]$ we find

$$\rightarrow \frac{d[\text{O}_2]}{dt} = \frac{k_1 K_a [\text{H}^+][\text{C}]^2}{(K_a + [\text{H}^+])^2}$$

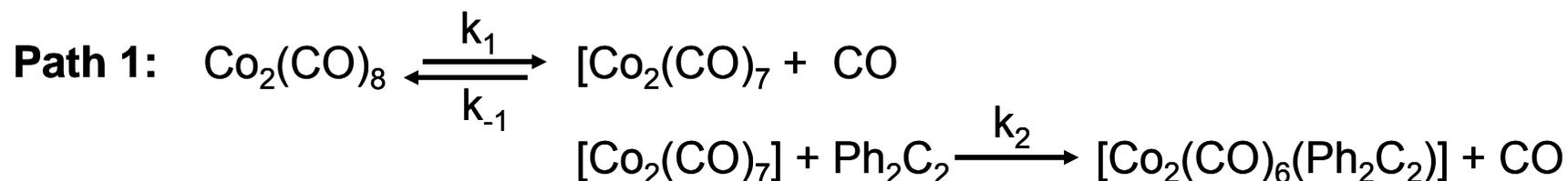
The reaction is **2nd order** and implies a mechanism with "A" and "HA"

Deduction of Mechanisms

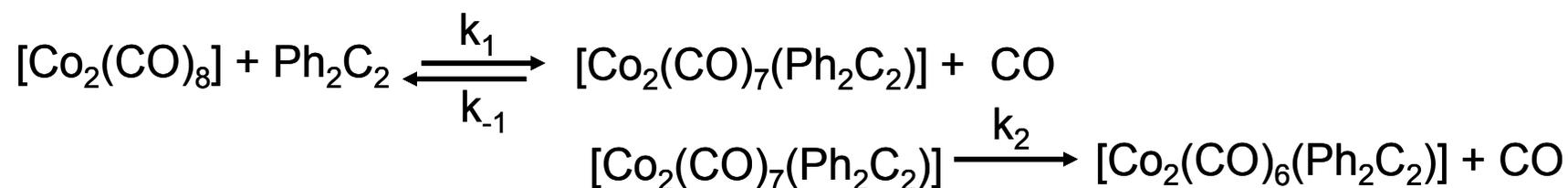
nice example from organometallic chemistry for mechanism deduction



these are 18 e⁻ species, associative mechanism is unlikely



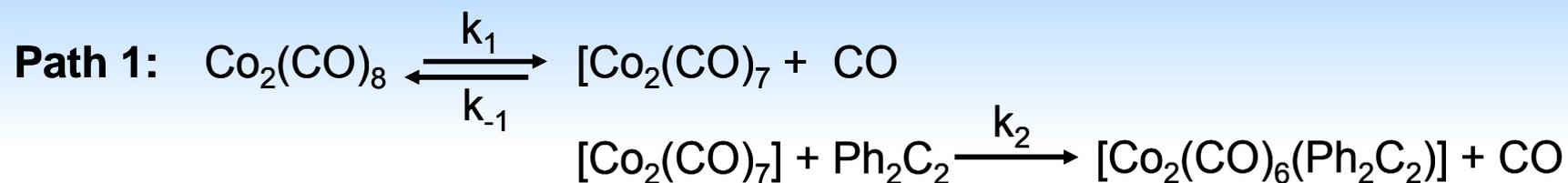
Path 2:



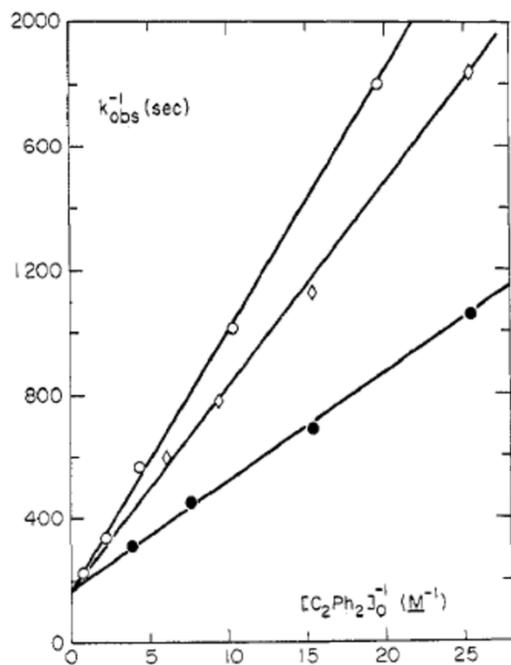
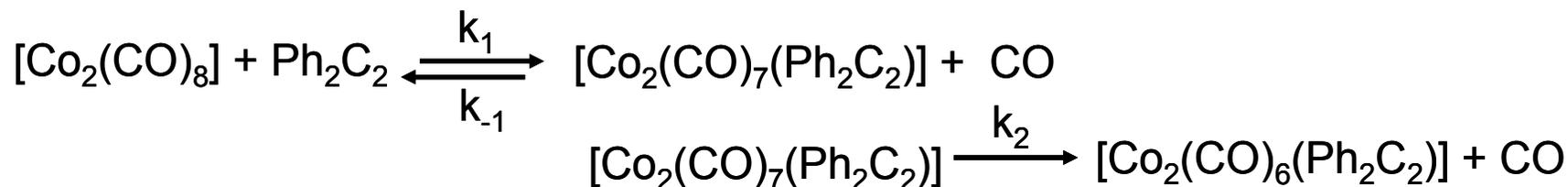
can we kinetically differentiate the two possibilities?

rate laws ?

Deduction of Mechanisms



Path 2:



experimental observations:

