











Power Law and Elementary Rate Laws In general $-r_{A} = k_{A} C_{A}{}^{\alpha} C_{B}{}^{\beta}$ $\alpha : order in A$ $\beta : order in B$ $n = \alpha + \beta = overall rxn order$ The unit of $-r_{A}$ is always = concentration / time For a rxn with "n" order: {k} = (concentration)^{1-n} / time Therefore for a zero-, first-, second-, and third-order rxn $zero - order \quad (n = 0) \qquad -r_{A} = k_{A}; \quad \{k\} = \frac{mol}{dm^{3}s}$ $1^{st} order \quad (n = 1) \qquad -r_{A} = k_{A}C_{A}; \quad \{k\} = \frac{1}{s}$ $2^{nd} order \quad (n = 2) \qquad -r_{A} = k_{A}C_{A}^{2}; \quad \{k\} = \frac{dm^{3}}{mol \cdot s}$ $3^{rd} order \quad (n = 3) \qquad -r_{A} = k_{A}C_{A}^{2}C_{B}; \quad \{k\} = \left(\frac{dm^{3}}{mol}\right)^{2} \cdot s^{-1}$

Elementary Reactions

A reaction follows an elementary rate law if and only if the stoichiometric coefficients are the same as the individual reaction order of each species. For the reaction in the previous example (A + B \rightarrow C + D), the rate law would be: -r_A = k C_A C_B These rate laws can be derived from Collision Theory.

if 2NO + $O_2 \rightarrow 2NO_2$ then $-r_{NO} = k_{NO} (C_{NO})^2 C_{O2}$ if elementary!!!

Question

What is the reaction rate law for the reaction A + $\frac{1}{2} B \rightarrow C$ if the reaction is elementary? What is r_B ? What is r_C ? Calculate the rates of A, B, and C in a CSTR where the concentrations are $C_A = 1.5 \text{ mol/dm}^3$, $C_B = 9 \text{ mol/dm}^3$ and $k_A = 2 \text{ (dm}^3/\text{mol})^{(\frac{1}{2})}$ (1/s)



Then,

$$-r_{A} = \left(2 \left(\frac{\mathrm{dm}^{3}}{\mathrm{mol}} \right)^{1/2} \frac{1}{\mathrm{s}} \right) \left(1.5 \frac{\mathrm{mol}}{\mathrm{dm}^{3}} \right) \left(9 \frac{\mathrm{mol}}{\mathrm{dm}^{3}} \right)^{1/2}$$

$$-r_{A} = 9 \frac{\mathrm{mol}}{\mathrm{dm}^{3} \cdot \mathrm{s}}$$

$$r_{B} = \frac{1}{2} \left(-k_{A} C_{A} C_{B}^{1/2} \right)$$

$$r_{B} = -4.5 \frac{\mathrm{mol}}{\mathrm{dm}^{3} \mathrm{s}}$$

$$r_{C} = -r_{A}$$

$$r_{C} = 9 \frac{\mathrm{mol}}{\mathrm{dm}^{3} \mathrm{s}}$$
For A + B \leftrightarrow C + D

$$-r_{A} = k \cdot \left[C_{A} \cdot C_{B} - \frac{C_{C} \cdot C_{D}}{K_{C}} \right]$$
Elementary

Non Elementary Rate Laws:

A large number of homogeneous or heterogeneous rxns do not follow simple rate laws.

If the rate law for the non-elementary reaction A + B \rightarrow 2C + D is found to be $-r_A = k C_A C_B^2$, then the rxn is said to be 2nd order in A and 1st order in B, and 3rd order overall.

For the homogeneous rxn

lf

$$\begin{split} & CO + Cl_2 \rightarrow COCl_2 \quad -r_{CO} = k \cdot C_{CO} \cdot C_{Cl_2}^{3/2} \\ & 2N_2O \rightarrow 2N_2 + O_2 \quad -r_{N_2O} = \frac{k \cdot C_{N_2O}}{1 + k' \cdot C_{O_2}} & \text{k and k' strongly} \\ & \text{temperature} \\ & \text{dependent} \end{split}$$

In this case we can not state overall rxn order.

Here, we can speak of reaction orders under certain limiting conditions as at very low conc'n of O_2 (1 >> k' C_{O_2})

$$-r_{N_2O} = k_{N_2O} \cdot C_{N_2O}$$
 1st order (apparent rxn order)

$$1 << k' C_{O_2} - r_{N_2O} - r_{N_2O} = \frac{k_{N_2O} \cdot C_{N_2O}}{k' \cdot C_{O_2}} \qquad 0^{\text{th} \text{ order (apparently)}}$$

These types are very common for liquid & gas phase rxns on solid catalysts



Example: Gas Phase Catalytic Reactions

When studying gas phase catalytic reactions the rate law is developed in terms of partial pressure,

$$-r'_{A} = \frac{k P_{A} P_{B}}{\left[1 + K_{A} P_{A}\right]}$$

To rewrite the rate law, just use ideal gas law to relate to concentrations C_{A} and C_{B}

$$C_A = \frac{P_A}{RT} \text{ or } P_A = C_A RT$$

and then write concentration in terms of conversion

$$-r'_{A} = \frac{k(RT)^{2}C_{A}C_{B}}{1 + K_{A}RTC_{A}}$$

The net rate of formation of any species is equal to its rate of formation in the forward reaction plus its rate of formation in the reverse reaction: rate_{net} = rate_{forward} + rate_{reverse} At equilibrium, rate_{net} \approx 0 and the rate law must reduce to an equation that is thermodynamically consistent with the equilibrium constant for the

reaction.

Example: Consider the exothermic, heterogeneous reaction; $A + B \rightarrow C$

At low temperature, the rate law for the disappearance of A is

$$-r_{A} = \frac{k_{A}P_{A}P_{B}}{1 + K_{A}P_{A}} \qquad \text{Recall } P_{A} = C_{A} R T$$

At high temperature, the exothermic reaction is significantly reversible:

 $A + B \Leftrightarrow C$

What is the corresponding rate law?

If the rate of formation of A for the forward reaction (A + B \rightarrow C) is

$$r_{A_{for}} = \frac{-k_A P_A P_B}{1 + K_A P_A}$$

then we need to assume a form of the rate law for the reverse reaction that satisfies the equilibrium condition. If we assume the rate law for the reverse reaction $(C \rightarrow A + B)$ is

$$r_{A_{rev}} = \frac{k_{-A}P_{C}}{1 + K_{A}P_{A}}$$

Then: $r_A = r_{A_{net}} = r_{A_{for}} + r_{A_{rev}}$ and: $-r_A = \frac{k_A}{1 + k_A P_A} \left(P_A P_B - \frac{P_C}{k_P} \right)$

Deriving –r_A:

The forward rate is: $r_{A_{for}} = \frac{-k_{A}P_{A}P_{B}}{1 + K_{A}P_{A}}$

And the reverse rate law is:

$$r_{A_{rev}} = \frac{k_{-A}P_{c}}{1+K_{A}P_{A}}$$

The net rate for species A is the sum of the forward and reverse rate laws:

$$r_A = r_{A_{net}} = r_{A_{for}} + r_{A_{rev}}$$

Substituting for r_{for} and r_{rev}:

g for
$$r_{for}$$
 and r_{rev} :

$$r_{A} = \frac{-k_{A}P_{A}P_{B}}{1+K_{A}P_{A}} + \frac{k_{-A}P_{C}}{1+K_{A}P_{A}}$$

$$-r_{A} = \frac{k_{A}P_{A}P_{B}}{1+K_{A}P_{A}} - \frac{k_{-A}P_{C}}{1+K_{A}P_{A}} = \frac{k_{A}P_{A}P_{B} - k_{-A}P_{C}}{1+K_{A}P_{A}}$$

$$k_{A} = \begin{pmatrix} -R_{C} & -R_{C} \end{pmatrix}$$

$$K_{P} = \frac{P_{Ce}}{P_{Ae}P_{Be}} = \frac{k_{A}}{k_{-A}} \qquad -r_{A} = \frac{k_{A}}{1 + K_{A}P_{A}} \left(P_{A}P_{B} - \frac{P_{C}}{K_{P}}\right)$$

Solving for Kp:Does this rate law satisfy our requirement at equilibrium.For a rxn at equilibrium $K_P = \frac{P_{Ce}}{P_{Ae}P_{Be}}$ At equilibrium, $r_{net} \approx 0$, so; $-r_A \approx 0 = \frac{K_A}{1+K_AP_{Ae}} \left(P_{Ae}P_{Be} - \frac{P_{Ce}}{K_P} \right)$ Solving for K_p : $0 = P_{Ae}P_{Be} - \frac{P_{Ce}}{K_P}$ $P_{Ae}P_{Be} = \frac{P_{Ce}}{K_P}$ $K_P = \frac{P_{Ce}}{P_{Ae}P_{Be}}$ The conditions are satisfied.

Rate Law for Reversible Reactions

Example: Write the rate law for the elementary reaction

$$A \xrightarrow{\underline{k_{fA}}} 2B$$

Here k_{fA} and k_{rA} are the forward and reverse specific reaction rates both defined with respect to A.

(1)
$$A \xrightarrow{k_{fA}} 2B$$
 $-r_{Af} = k_{fA}C_A$
(2) $2B \xrightarrow{k_{7A}} A$ $r_{Ar} = k_{rA}C_B^2$

rate net = rate forward + rate revense











Why is there an Activation Energy?

(1) the molecules need energy to distort or stretch their bonds in order to break them and to thus form new bonds

(2) as the reacting molecules come close together they must overcome both steric and electron repulsion forces in order to react

Energy Distribution of reacting molecules

In our development of collision theory we assumed all molecules had the same average energy. However, all the molecules don't have the same energy, rather there is distribution of energies where some molecules have more energy than others. The distraction function f(E,T) describes this distribution of the energies of the molecules. The distribution function is read in conjunction with dE

f(E, T) dE = fraction of molecules with energies between E and E + dE







Activation Energy

The activation energy is a measure of how temperature sensitive the reaction is. Reactions with large activation energies are very temperature sensitive.





Specific Reaction Rate Derivation

$$k = Ae^{-B/RT}$$
$$k_o = Ae^{-B/RT_o}$$

Taking the ratio:

$$k = k_o \exp\left[\frac{E}{R}\left(\frac{1}{T_o} - \frac{1}{T}\right)\right]$$

This says if we know the specific rxn rate $k_o(T_o)$ at a temperature T_o , and we know E, then we can find reaction rate k(T) at any other temperature T for that rxn.

STOICHIOMETRY

We shall set up <u>Stoichiometric Tables</u> using A as our basis of calculation in the following reaction. We will use the stoichiometric tables to express the concentration as a function of conversion. We will combine $C_i = f(X)$ with the appropriate rate law to obtain $-r_A = f(X)$.

$$A + \frac{b}{a}B \rightarrow \frac{c}{a}C + \frac{d}{a}D$$

<u>Species</u>	<u>Symbol</u>	Initial	<u>Change</u>	<u>Remaining</u>
А	А	N _{A0}	$-N_{A0} \times$	$N_A = N_{A0}(1-X)$
В	В	$N_{B0} = N_{A0} \odot_B$	$-\frac{b}{a}N_{A0}X$	$N_{B} = N_{A0} \left(\Theta_{B} - \frac{b}{a} X \right)$
с	с	$N_{C0} = N_{A0} \Theta_C$	$+\frac{c}{a}N_{A0}X$	$N_{\rm C} = N_{\rm A0} \left(\odot_{\rm C} + \frac{c}{a} \right)$
D	D	$N_{D0} = N_{A0} \Theta_{D}$	$+\frac{d}{a}N_{A0}X$	$N_{D} = N_{A0} \left(\Theta_{D} + \frac{d}{a} \right)$
Inert	Ι	$N_{I} = N_{A0}\Theta_{I}$		$N_{I} = N_{A0} \odot_{I}$
		N _{TO}		$N_T = N_{T0} + \delta N_{A0} X$

Concentration -- Batch System: $C_A = \frac{N_A}{V}$ **Constant Volume Batch:** Note: if the reaction occurs in the liquid phase or if a gas phase reaction occurs in a rigid (e.g., steel) batch reactor Then $V = V_0$ $C_A = \frac{N_A}{V} = \frac{N_{A0}(1-X)}{V_0} = C_{A0}(1-X)$ $C_B = \frac{N_B}{V} = \frac{N_{A0}}{V_0} (\Theta_B - \frac{b}{a}X) = C_{A0} (\Theta_B - \frac{b}{a}X)$ if $-r_a = k_a C_a^2 C_b$ then $-r_a = C_{a0}^3 (1-X)^2 (\Theta_b - \frac{b}{a}X)$ and we have $-r_a = f(x)$



What is $\Theta_{\rm B}$?

$$\theta_{\rm B} = \frac{{\rm N}_{\rm B0}}{{\rm N}_{\rm A0}} = \frac{1}{1} = 1$$

Species A is the limiting reactant because the feed is equal molar in A and B, and two moles of B consumes 3 moles of A.

$$A + \frac{2}{3}B \longrightarrow \frac{4}{3}C$$

$$\frac{b}{a} = \frac{2/3}{1} = \frac{2}{3}$$

$$C_{B} = C_{A0} \left(1 - \frac{2}{3}X\right)$$

$$-r_{A} = kC_{A0}^{5} (1 - X)^{3} \left(1 - \frac{2}{3}X\right)^{2}$$

$$= (0.01) \left(\frac{dm^{3}}{mol}\right)^{4} \frac{1}{s} \left(2\frac{mol}{dm^{3}}\right)^{5} (1 - X)^{3} \left(1 - \frac{2}{3}X\right)^{2}$$

$$= 0.32 (1 - X)^{3} \left(1 - \frac{2}{3}X\right)^{2} \frac{mol}{dm^{3} * s}$$

We now have $-r_A = f(X)$ and can size reactors or determine batch reaction times.



$$\begin{array}{ll} \text{Where:} \quad \oplus_i = \frac{F_{i0}}{F_{A0}} = \frac{C_{i0} \upsilon_0}{C_{A0} \upsilon_0} = \frac{C_{i0}}{C_{A0}} = \frac{y_{i0}}{y_{A0}} \quad \text{and} \quad \delta = \frac{d}{a} + \frac{c}{a} - \frac{b}{a} - 1 \\ \text{Concentration -- Flow System:} \quad C_A = \frac{F_A}{\upsilon} \\ \text{Liquid Phase Flow System:} \quad \upsilon = \upsilon_0 \\ \\ C_A = \frac{F_A}{\upsilon} = \frac{F_{A0}(1-X)}{\upsilon_0} = C_{A0}(1-X) \\ \\ C_B = \frac{F_B}{\upsilon} = \frac{F_{A0}}{\upsilon_0} \left(\bigoplus_B - \frac{b}{a} X \right) = C_{A0} \left(\bigoplus_B - \frac{b}{a} X \right) \qquad \text{etc.} \\ \\ \text{If the rate of reaction were -r_A = kC_AC_B then we would have} \\ -r_A = kC_{A0}^2(1-X)(\theta_B - \frac{b}{a} X) \\ \\ \\ \text{This gives us -r_A = f(X). Consequently, we can use the methods discussed in Chapter 2 to size a large number of reactors, either alone or in series. \\ \end{array}$$





Gas Phase Flow Systems

$$\begin{split} \upsilon &= \upsilon_0 \left(1 + \varepsilon X \right) \frac{\mathsf{T}}{\mathsf{T}_0} \frac{\mathsf{P}_0}{\mathsf{P}} \\ \mathsf{C}_\mathsf{A} &= \frac{\mathsf{F}_\mathsf{A}}{\upsilon} = \frac{\mathsf{F}_\mathsf{A} 0 \left(1 - X \right)}{\upsilon_0 \left(1 + \varepsilon X \right)} \frac{\mathsf{T}_0}{\mathsf{T}} \frac{\mathsf{P}}{\mathsf{P}_0} = \mathsf{C}_\mathsf{A} 0 \left(\frac{\left(1 - X \right)}{\left(1 + \varepsilon X \right)} \frac{\mathsf{T}_0}{\mathsf{T}} \frac{\mathsf{P}}{\mathsf{P}_0} \right) \\ \mathsf{C}_\mathsf{B} &= \frac{\mathsf{F}_\mathsf{B}}{\upsilon} = \frac{\mathsf{F}_\mathsf{A} 0 \left(\frac{\mathsf{\Theta}_\mathsf{B}}{\mathsf{H}} - \frac{\mathsf{b}}{\mathsf{a}} X \right)}{\upsilon_0 \left(1 + \varepsilon X \right)} \frac{\mathsf{T}_0}{\mathsf{T}} \frac{\mathsf{P}}{\mathsf{P}_0} = \mathsf{C}_\mathsf{A} 0 \left(\frac{\mathsf{\Theta}_\mathsf{B}}{\left(1 + \varepsilon X \right)} \frac{\mathsf{b}}{\mathsf{T}} \frac{\mathsf{T}_0}{\mathsf{P}_0} \frac{\mathsf{P}}{\mathsf{P}_0} \right) \\ \end{split}$$

etc.

Again, these equations give us information about $-r_A = f(X)$, which we can use to size reactors.

For example if the gas phase reaction has the rate law

$$-\mathbf{r}_{a} = \mathbf{k}_{a} \mathbf{C}_{a}^{2} \mathbf{C}_{b}$$



	Batch			
	<u>Species</u>	<u>Initial</u>	<u>Change</u>	<u>Remaining</u>
	А	N _{AO}	-N _{A0} X	$N_A = N_{A0}(1-X)$
	В	0	+N _{A0} X/2	$N_{B} = N_{A0}X/2$
		$N_{TO} = N_{AO}$		$N_{T} = N_{A0} - N_{A0}X/2$
Constant V	/olume V =	V ₀		

$$C_{A} = \frac{N_{A}}{V} = \frac{N_{A}}{V_{0}} = \frac{N_{A0}(1-X)}{V_{0}} = C_{A0}(1-X)$$
$$C_{B} = \frac{N_{B}}{V} = \frac{N_{A0}X/2}{V_{0}} = \frac{C_{A0}X}{2}$$

$$K_{C} = \frac{C_{Be}}{C_{Ae}^{2}} = \frac{C_{A0} \frac{X_{e}}{2}}{C_{A0}^{2} (1 - X_{e})^{2}}$$

$$2K_{C}C_{A0} = \frac{X_{e}}{(1 - X_{e})^{2}} = (2\left(20\frac{dm^{3}}{mol}\right)\left(0.2\frac{mol}{dm^{3}}\right) = 8$$

$$8X_{e}^{2} - 17X_{e} + 8 = 0$$
Solving
Batch X = 0.7

<u>Species</u>	<u>Fed</u>	<u>Change</u>	<u>Remaining</u>
А	F _{AO}	-F _{AO} X	F _A = F _{A0} (1-X)
В	0	+F _{A0} X/2	$F_{B} = F_{A0}X/2$
	F _{T0} = F _{A0}		$F_{T} = F_{A0} - F_{A0}X/2$
$C_{Ae} = \frac{C_{A0}}{(1+1)}$ $Be = \frac{C_{A0}}{24}$	$\frac{(1-X_e)}{eX_e}$		

$$K_{C} = \frac{\frac{C_{A0}X_{e}}{2(1+\varepsilon X_{e})}}{\left[\frac{C_{A0}(1-X_{e})}{(1+\varepsilon X_{e})}\right]^{2}} = \frac{X_{e}(1+\varepsilon X_{e})}{2C_{A0}(1-X_{e})^{2}}$$

$$2K_{C}C_{A0} = \frac{X_{e}+\varepsilon X_{e}^{2}}{(1-2X_{e}+X_{e}^{2})} = 2\left(20\frac{\mathrm{dm}^{3}}{\mathrm{mol}}\right)\left(0.2\frac{\mathrm{mol}}{\mathrm{dm}^{3}}\right) = 8$$

$$\varepsilon = y_{A0}\delta = 1\left(\frac{1}{2}-1\right) = -\frac{1}{2}$$

$$8 = \frac{X_{e}-0.5X_{e}^{2}}{(1-2X_{e}+X_{e}^{2})}$$

$$8.5X_{e}^{2}-17X_{e}+8=0$$

Flow : $X_{e} = 0.757$



Basis of Calculation

The stoichiometric ratio is one toluene to two ethylene (1/2). However, the feed is one toluene to three ethylene (1/3) and there is not sufficient toluene to consume all the ethylene. Therefore toluene is the limiting reactant and thus the basis of calculation.

Entering Concentrations of ethylene and toluene

Let A = toluene, B = ethylene, C = ethyl benzene and D = propylene A + 2B \longrightarrow C + D C_{A0} = y_{A0}C_{T0} = y_{A0} $\frac{P_0}{RT_0}$ = 0.25 $\frac{8.2 \text{ atm}}{\left(0.082 \frac{\text{atm} \cdot \text{dm}^3}{\text{mol} \cdot \text{K}}\right)^2} = (0.25)\left(0.2 \frac{\text{mol}}{\text{dm}^3}\right) = 0.05 \frac{\text{mol}}{\text{dm}^3}$ C_{A0} = $\frac{P_0}{RT_0} = 0.2 \frac{\text{mol}}{\text{dm}^3}$ Ethylene C_{B0} = y_{B0}C_{T0} = (0.75)0.20 $\frac{\text{mol}}{\text{dm}^3} = 0.15 \frac{\text{mol}}{\text{dm}^3}$

Since toluene, i.e. A, is the limiting reactant and has a stoichiometric coefficient of 1

$$y_{AB} = 0.25$$
 A + 2B \longrightarrow C + D
 $\delta = (1 + 1 - 1 - 2) = -1$
 $\sigma = y_{A0}\delta = (0.25)(-1) = -0.25$

Species	Symbol	Entering	Change	Leaving			
Toluene	A	F _{AO}	-F _{A0} X	F _A =F _{A0} (1-X)			
Ethylene	В	$F_B = \theta_B F_{A0}$	-2FAOX	$F_{\rm B} = F_{\rm A0} \left(\theta_{\rm B} - 2 X \right)$			
	8	$B = \frac{y_{B0}}{y_{A0}} = \frac{.75}{.25} =$	3				
Leaving $F_{B} = F_{A0}(3-2X)$							

Species	Symbol	Entering	Change	Leaving
Toluene	А	F _{AO}	-F _{AO} X	F _A =F _{A0} (1-X)
Ethylene	В	$F_B = 3F_{A0}$	-2F _{A0} X	$F_{\rm B} = F_{\rm A0} (3 - 2X)$
Ethyl benzene	С	0	+F _{A0} X	$F_{\rm C} = F_{\rm A0} X$
Proplyene	D	0	+F _{A0} X	$F_D = F_{A0}X$
Total		$F_{\rm m} = 4F_{\rm A0}$		$F_{T} = 4F_{A,0} - F_{A,0}$

Write the volumetric flow rate in terms of conversion

$$\begin{aligned}
& v = v_0 \left(1 + sX\right) \frac{P_0}{P} \frac{T}{T_0} \\
& P = P_0 \text{ and } T = T_0 \\
& \varepsilon = y_{A0} \delta = (0.25)(1 + 1 - 1 - 2) = -0.25 \\
& v = v_0 (1 - 0.25X)
\end{aligned}$$
In terms of conversion

$$C_A = \frac{F_A}{v}$$
For a flow system at constant T and P

$$C_A = \frac{F_A}{v} - \frac{F_{A0}(1 - X)}{v_0(1 + sX)} - C_{A0} \frac{(1 - X)}{(1 + sX)}$$

$$C_A = \frac{0.05(1 - X)}{1 - 0.25X}$$

$$C_B = \frac{F_B}{v} = \frac{F_{A0}(3 - 2X)}{v_0(1 + sX)} = C_{A0} \frac{(3 - 2X)}{(1 - 0.25X)}$$

In terms of conversion
$$-r_{A} = kC_{A}C_{B}^{2}$$
$$C_{A} = \frac{C_{A0}(1-X)}{(1+eX)} \qquad C_{B} = C_{A0}\frac{(3-2X)}{(1+eX)}, \quad s = -0.25$$
$$-r_{A} = kC_{A0}^{3}\frac{(1-X)(3-2X)^{2}}{(1-0.25X)^{3}}$$
$$kC_{A0}^{3} = y_{A0}^{3}kC_{T0}^{2} = (0.25)^{3}\left[\frac{2\times10^{-1}}{s}\right]\left(\frac{mol}{dm^{-3}}\right) = \frac{3.13\times10^{-3}}{s}\left(\frac{mol}{dm^{-3}}\right)$$
$$-r_{A} = \frac{3.13\times10^{-3}}{s}\left(\frac{mol}{dm^{-3}}\right)\frac{(1-X)(3-2X)^{2}}{(1-0.25X)^{3}}$$





atch System Stoichiometric Reactor (similarly for TSIR and PFR!!)					
pecies	<u>Symbol</u>	Initial	<u>Change</u>	Remaining	
А	А	N _{A0}	$-N_{A0} \times$	$N_A = N_{A0}(1 - X)$	
В	В	$N_{B0} = N_{A0} \odot_B$	$-\frac{b}{a}N_{A0}X$	$N_{B} = N_{A0} \left(\Theta_{B} - \frac{b}{a} X \right)$	
С	с	$N_{C0} = N_{A0} \Theta_C$	$+\frac{c}{a}N_{A0}X$	$N_{C} = N_{A0} \left(\Theta_{C} + \frac{c}{a} X \right)$	
D	D	$N_{D0} = N_{A0} \Theta_D$	$+\frac{d}{a}N_{A0}X$	$N_{D} = N_{A0} \left(\Theta_{D} + \frac{d}{a} X \right)$	
Inert	Ι	$N_{I} = N_{A0}\Theta_{I}$		$N_{I} = N_{A0} \Theta_{I}$	
		N _{TO}		$N_T = N_{T0} + \delta N_{A0} \times$	