C12 - 2.0 - Reaction Rates/Graphs

$\boxed{Rate = \frac{amount}{time}}$

Find rate if 8g used up in 4 minutes.

$$Rate = \frac{8g}{4\,min}$$

$$Rate = 2\frac{g}{min}$$

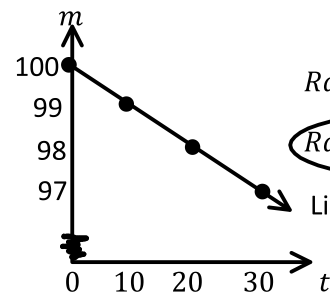
If rate of water production is $0.5\frac{mol}{min}$, find the rate of oxygen gas consumption.

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(l)}$$

$$0.5\frac{mol\ H_2O}{min} \times \frac{1\ mol\ O}{2\ mol\ H_2O} = 0.25\frac{mol\ O_2}{min}$$

Rates :
Moles Reactants used up = Moles Products are made
(Coefficients*)

$Time(s)$	$Mass(g)$
0	100
10	99
20	98
30	97

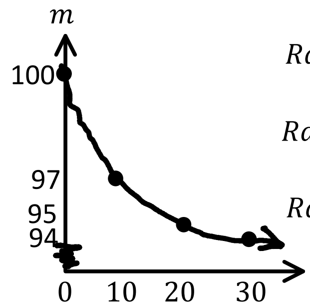


$$Rate = \frac{97-100}{30-0}$$

$$Rate = -0.1\frac{g}{min}$$

Linear/Constant Rate

$Time(s)$	$Mass(g)$
0	100
10	97
20	95
30	94

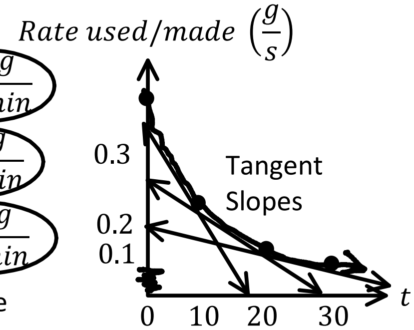


$$Rate_{0-10} = \frac{97-100}{10-0} = -0.3\frac{g}{min}$$

$$Rate_{10-20} = \frac{95-97}{20-10} = -0.2\frac{g}{min}$$

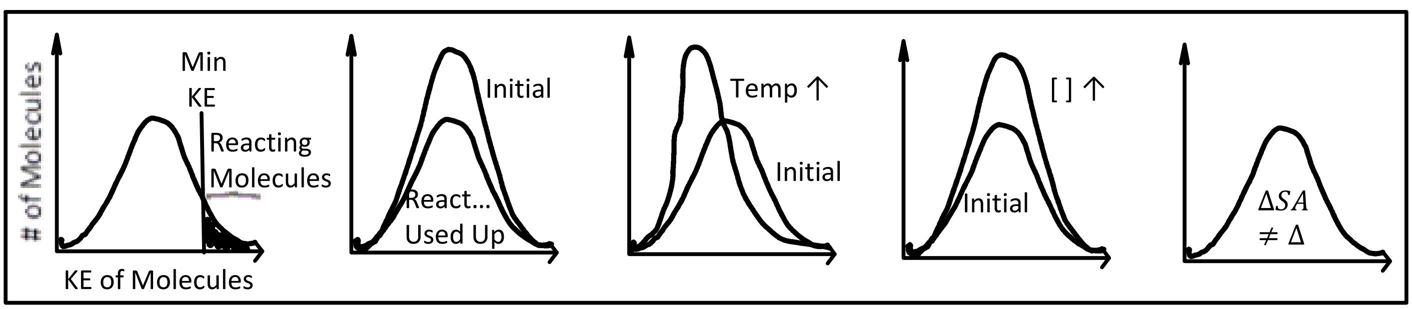
$$Rate_{20-30} = \frac{94-95}{30-20} = -0.1\frac{g}{min}$$

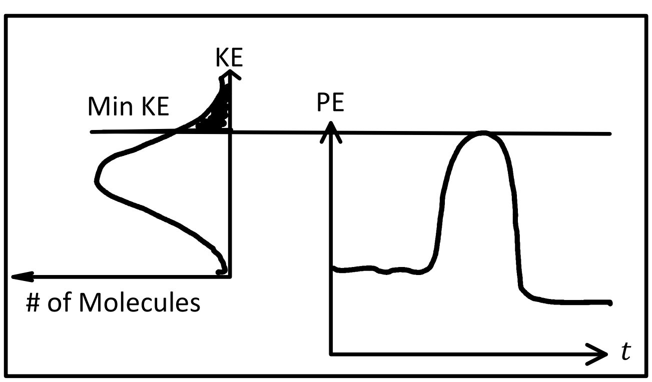
Exponential/Decreasing Rate

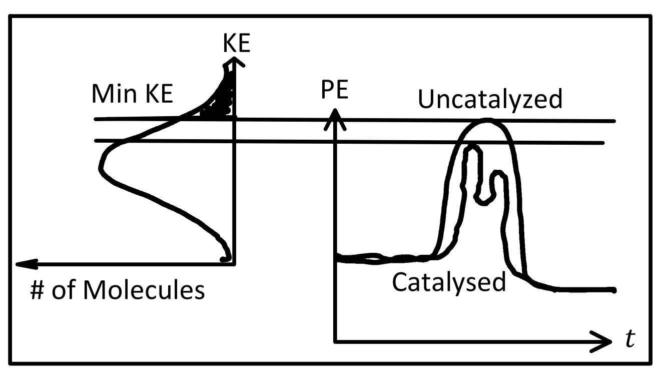


Faster
$Li^+ + F^- \rightarrow LiF$
-1 bonds made

Slower
$H-H + O=O \rightarrow 2H-O-H$
-2 bonds broken, 4 bonds made







C12 - 2.0 - Equilibrium (Eq)

Don't have to calculate the rates*

Equilibrium - $Rate_{forward}$; $K_f = K_r$; $Rate_{reverse}$. (Microscopic changes*)

K : Rate constant

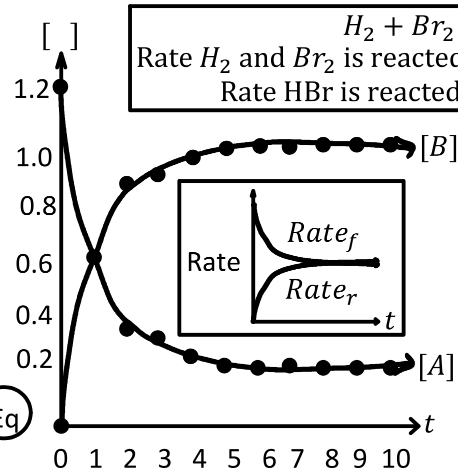
$K_f \propto$ [Reactant]
$K_r \propto$ [Product]

Rate of consumption of reactants = rate of production of reactants. (A system will tend towards Eq.)
[Reactants] & [Products] constant in time. [Reactants] differ from [Products] in general.

$A \rightleftharpoons B \quad [A] = 1.2\ M \quad [B] = 0\ M \quad K_{fi} = 0.5 \quad K_{ri} = 0.1$ $\boxed{i : initial}$

Time (min)	R_f	R_r	$[A]$	$[B]$
0	0.600	0.000	1.200	0.000
1	0.300	0.060	0.600	0.600
2	0.180	0.084	0.360	0.840
3	0.132	0.094	0.264	0.936
4	0.113	0.097	0.226	0.974
5	0.105	0.099	0.210	0.990
6	0.102	0.100	0.204	0.996
7	0.101	0.100	0.202	0.998
8	0.100	0.100	0.201	0.999
9	0.100	0.100	0.200	1.000
10	0.100	0.100	0.200	1.000

Eq



$H_2 + Br_2 \rightleftharpoons 2HBr$
Rate H_2 and Br_2 is reacted=Rate H_2 and Br_2 is produced
Rate HBr is reacted=Rate HBr is produced

$$\frac{K_{fi}}{K_{ri}} = \frac{0.5}{0.1} = 5$$

$$\frac{[Product]}{[Reactant]_{eq}} = \frac{1}{0.2} = 5$$

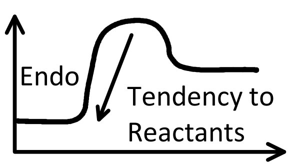
Spontaneous Change- A change that occurs by itself, without outside assistance. (Exothermic*)

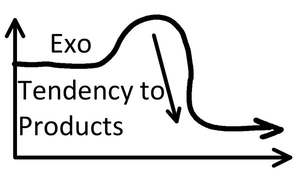
Entropy - Amount of randomness in a system. $\boxed{gases \gg solutions \gg liquids \gg solids}$

Tendency for a reaction to go to the side of Maximum Entropy (Most particles of most random phase)/Minimum Energy (Enthalpy).

$$A_{(s)} + B_{(l)} \underset{}{\overset{> Random}{\rightleftharpoons}} 2C_{(g)} + D_{(aq)}$$

$$A_{(g)} + B_{(s)} \underset{> Random}{\rightleftharpoons} 2C_{(g)} + D_{(s)}$$





Le Chatelier's Principle - If a closed system (Nothing can enter or leave) at equilibrium is subject to a change, processes will occur that tend to counteract that change. (Whatever we do, nature tends to undo*.)

Effects

$\boxed{2A_{(g)} + B_{(g)} \rightleftharpoons 2C_{(g)} + kJ}$

↑ [B] Shifts to use up more [A].

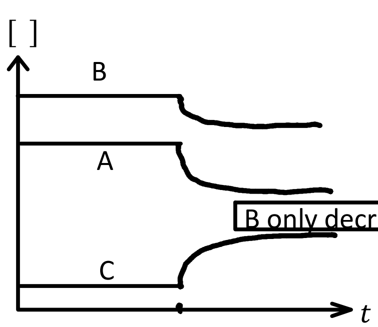
Initially : ↑r_f, ↓r_r, Then : ↓r_f, ↑r_r

$A + B \rightleftharpoons C$ +kj

End : ↑[B],↑[C],↑[D],↓[A]

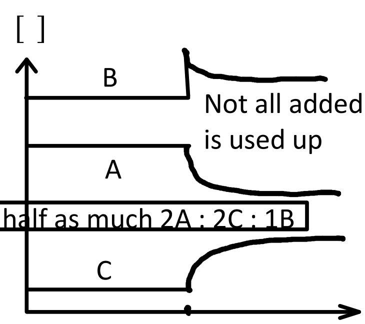
↓ heat. Shifts to produce more heat.
$A \rightleftharpoons B + kJ$

↓ V, ↑ P, ↑[gases]
Shifts to lower overall pressure.
(Side with lower moles.)

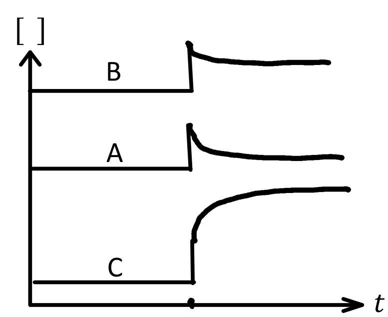


Temp ↓

B only decreases half as much 2A : 2C : 1B



[B]↑



P↑

C12 - 2.0 - Equilibrium

$4.0\ mol\ NO_{2(g)} \rightarrow 2.0L\ Bulb$ $\quad 2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)}$ $\quad Find\ k_{eq}\ if\ @\ Eq\ 0.50\ mol\ NO_{2(g)}$

$$k_{eq} = \frac{[NO_{2(g)}]^2}{[NO_{(g)}]^2[O_{2(g)}]} \qquad [NO_{2(g)}]_{start} = \frac{4.0\ mol}{2.0\ L} = 2.0\ M \qquad [NO_{(g)}]_{eq} = \frac{0.5\ mol}{2.0\ L} = 0.25\ M$$

$\frac{mol}{L}$	$2NO_{(g)}$	$+ O_{2(g)}$	$\rightleftharpoons 2NO_{2(g)}$
$Start$	0	0	2
$+\Delta$	0.25	0.125	− 0.25
$= Eq$	0.25	0.125	1.75

$\boxed{Mole\ Ratio\ -\ 2:1:2}$

$$\Delta[O_2] = \Delta[NO] \times \frac{1\ mol\ O_2}{2\ mol\ NO} = 0.25 \times \frac{1}{2} = 0.125$$

$$\Delta[NO_2] = \Delta[NO] \times \frac{2\ mol\ NO_2}{2\ mol\ NO} = 0.25 \times \frac{2}{2} = 0.25$$

$$k_{eq} = \frac{[1.75]^2}{[0.25]^2[0.125]} = 392$$

$"x"\ mol\ NO_{2(g)} \rightarrow 5.0L\ Bulb$ $\quad 2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)}$ $\quad [NO_{(g)}]_{eq} = 0.80\ M$ $\quad k_{eq} = 24$ $\quad [NO_{2(g)}] = ?$

$\frac{mol}{L}$	$2NO_{(g)}$	$+ O_{2(g)}$	$\rightleftharpoons 2NO_{2(g)}$
$Start$	0	0	x
$+\Delta$	0.80	0.40	− 0.80
$= Eq$	0.80	0.40	$(x - 0.80)$

$\boxed{Mole\ Ratio\ -\ 2:1:2}$

$$k_{eq} = \frac{[NO_{2(g)}]^2}{[NO_{(g)}]^2[O_{2(g)}]}$$

$$24 = \frac{[x - 0.80]^2}{[0.8]^2[0.4]}$$

$$0.256 \times 24 = \frac{[x - 0.80]^2}{0.256} \times 0.256$$

$$\sqrt{6.144} = \sqrt{[x - 0.80]^2}$$

$$2.4787 = x - 0.80$$

$$x = 3.2786$$

$$[NO_{2(g)}] = 3.2786\ M$$

$$\#\ mol\ N0_{2(g)} = 3.2786 \frac{mol}{L} \times 5.0\ L = 16.39\ mol\ NO_{2(g)}$$

$k_{eq} = 48$ $\quad 2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)}$ $\quad 2.0\ mol\ NO_{(g)}$ $\quad 0.2\ mol\ O_{2(g)}$ $\quad 0.4\ mol\ NO_{2(g)}$ $\quad 2.0L\ Bulb$

Which direction will the reaction shift?

$Q = Reaction\ Quotient$

$$Q = \frac{[NO_{2(g)}]^2}{[NO_{(g)}]^2[O_{2(g)}]} \qquad k_{eq} = \frac{[NO_{2(g)}]^2}{[NO_{(g)}]^2[O_{2(g)}]}$$

$$[NO_{(g)}]_{start} = \frac{2.0\ mol}{2.0\ L} = 1.0\ M \qquad [O_{2(g)}]_{start} = \frac{0.2\ mol}{2.0\ L} = 0.1\ M \qquad [NO_{2(g)}]_{start} = \frac{0.4\ mol}{2.0\ L} = 0.2\ M$$

$$\boxed{Q = \frac{[0.2]^2}{[1.0]^2[0.1]} = 0.4 = \frac{[Products]}{[Reactants]} \uparrow < k_{eq} = 49}$$

Shift towards products

$k_{eq} = 4$ $\quad N_{2(g)} + O_{6(g)} \rightleftharpoons 2NO_{3(g)}$ $\quad 1.0\ mol\ N_{2(g)}$ $\quad 1.0\ mol\ O_{6(g)}$ $\quad 4.0L\ Bulb$ $\quad Find\ [\ ]_{eq}$

$\frac{mol}{L}$	$N_{2(g)}$	$+$	$O_{6(g)}$	$\rightleftharpoons$	$2NO_{3(g)}$
$Start$	0.25		0.25		0
$+\Delta$	$-x$		$-x$		x
$= Eq$	$(0.25 - x)$		$(0.25 - x)$		x

$\boxed{Mole\ Ratio\ -1:1:1}$

$$[NO_{(g)}]_{start} = \frac{1.0\ mol}{4.0\ L} = 0.25\ M \qquad [O_{2(g)}]_{start} = \frac{1.0\ mol}{4.0\ L} = 0.25\ M$$

$$k_{eq} = \frac{[NO_{3(g)}]}{[NO_{(g)}][O_{2(g)}]}$$

$$4 = \frac{[x]^2}{[0.25 - x][0.25 - x]}$$

$$4 = \frac{x^2}{[0.25 - x]^2}$$

$$2 = \frac{x}{0.25 - x}$$

$$2(0.25 - x) = x$$

$$0.5 - 2x = x$$

$$0.5 = 3x$$

$$x = 0.167$$

Square Root both sides
Cross Multiply
Distribute/Subtract/Divide

$$[NO_{3(g)}] = 0.17\ M$$

$$[N_{2(g)}] = [O_{6(g)}] = (0.25 - x) = (0.25 - 0.167) = 0.08$$