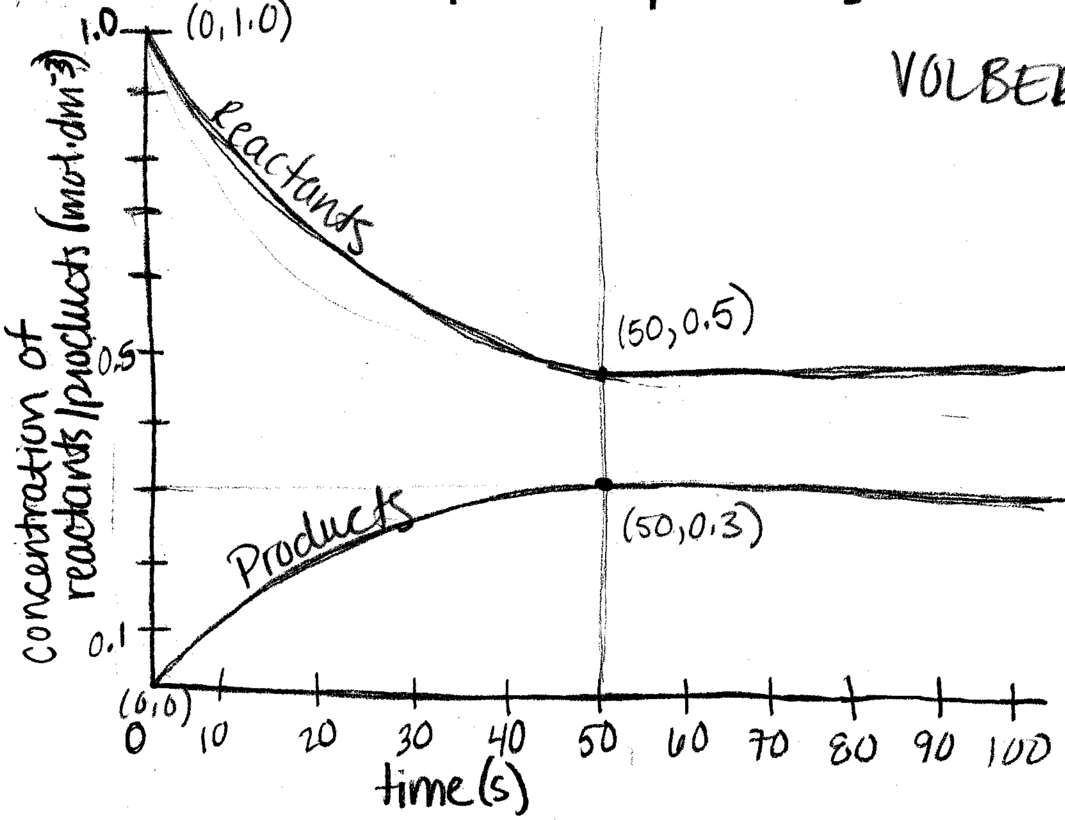
IB Chem #117 extra practice problems

VOLBERS KEY

1.



a) calculate the average rate of reaction over the first 50 seconds, stating the units.

$\frac{\Delta y}{\Delta x}$ $\frac{0.5-1.0}{50-0} = -0.01\ mol \cdot dm^{-3} \cdot s^{-1}$ (reactants) $\frac{0.3-0}{50-0} = +0.006\ mol \cdot dm^{-3} \cdot s^{-1}$ (products)

b) why are the concentrations of the reactants and products no longer changing after 50 seconds?

They are @ equilibrium (concentrations do not change)

c) If a catalyst was added to this reaction, how would it change the equilibrium? Would the K_c value be affected? Would the reaction rates be affected?

catalyst do not affect the equilibrium, they only affect the time it would take to reach equilibrium (and the rate to get there)

K_c value does not change

the forward and reaction rates are increased by the same amount so the overall rate is not affected

② The following reaction is observed to reach equilibrium at a constant temperature:

$$2NO_{(g)} + 2H_{2(g)} \rightleftharpoons N_{2(g)} + 2H_2O_{(g)}$$

the initial concentrations of NO and H_2 are both $0.200\ mol \cdot dm^{-3}$. The K_c value is $3.87 * 10^{-7}$. What is the equilibrium concentration of $H_2O_{(g)}$?

	$2NO$	$+ 2H_2 \rightleftharpoons$	N_2	$+ 2H_2O$
R				
I	0.200	0.200	0	0
C	$-2x$	$-2x$	$+x$	$+2x$
E	0.200	0.200	0.000537	0.00107

$$K_c = \frac{[N_2][H_2O]^2}{[NO]^2[H_2]^2} = \frac{(x)(2x)^2}{(0.200)^2(0.200)^2} = 3.87 * 10^{-7}$$

$x = 0.000537$

$[H_2O]$ @ equilibrium is 0.00107

$\boxed{1.07 * 10^{-3}\ mol \cdot dm^{-3}}$

③ During the Haber Process under constant temperature of 273 K, the following equilibrium is observed:

$$N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$$

Ⓐ If the initial concentrations of N_2 and H_2 are $0.0300\ mol \cdot dm^{-3}$, and the equilibrium concentration of NH_3 is 0.0100, what is the K_c value for the reaction?

	N_2	$+ 3H_2 \rightleftharpoons$	$2NH_3$
R			
I	0.03	0.03	0
C	$-x$	$-3x$	$+2x$
E	0.025	0.015	0.0100

$2x = 0.0100$

$x = 0.005$

$$K_c = \frac{[NH]^2}{[N_2][H_2]^3} = \frac{(0.01)^2}{(0.025)(0.015)^3} = 1185 = \boxed{1.19 * 10^3}$$

Ⓑ Calculate ΔG^θ for this reaction. -16056

$\Delta G^\theta = -RT \ln K$

it's spontaneous!

$$= -(8.31\ J/K\,mol)(298K)(\ln(1.19 * 10^3)) = -17537\ J/mol$$

$$= \boxed{-17.5\ kJ/mol}$$

Ⓒ If the K_c value decreases as the temperature increases for this reaction, what can you assume about ΔH? negative

$K_c \downarrow$ as $T \uparrow$ means equil. shifts towards reactants so exothermic