Chemistry 12
Worksheet 2-3

KEY

Calculations Involving the Equilibrium Constant K_{eq})

1. Given the equilibrium equation below:

$$A_{2(g)} + B_{2(g)} \rightleftarrows 2AB_{(g)}$$

If, *at equilibrium*, the concentrations are as follows:

2SD

$[A_2] = 3.45$ M, $[B_2] = 5.67$ M and $[AB] = 0.67$ M

a) Write the **expression** for the equilibrium constant, K_{eq}

$$K_{eq} = \frac{[AB]^2}{[A_2][B_2]}$$

b) Find the **value** of the equilibrium constant, K_{eq} at the temperature that the experiment was done.

$$K_{eq} = \frac{(0.67)^2}{(3.45)(5.67)} = 0.023$$

Answer $K_{eq} = 0.023$

2. Given the equilibrium equation:

$$X_{2(g)} + 3Y_{2(g)} \rightleftarrows 2XY_{3(g)}$$

at a temperature of 50°C, it is found that when equilibrium is reached that:

$[X_2] = 0.37$ M, $[Y_2] = 0.53$ M and $[XY_3] = 0.090$ M (All 2 SD)

a) Write the **equilibrium constant expression** (K_{eq})

$$K_{eq} = \frac{[XY_3]^2}{[X_2][Y_2]^3}$$

b) Calculate the **value** of K_{eq} at 50°C.

$$K_{eq} = \frac{(0.090)^2}{(0.37)(0.53)^3} = 0.15$$

Answer $K_{eq} = 0.15$

KEY

3. For the reaction: $A_{2(g)} + B_{(g)} \rightleftarrows 2C_{(g)}$

it is found that by adding 1.5 moles of C to a 1.0 L container, an equilibrium is established in which 0.30 moles of B are found. (*Hint: Make a table and use it to answer the questions below.*)

1/1 2/1 1 dec place

	A_2	+ B	⇄ 2 C
[I]	0	0	1.5
[C]	+ 0.30	+ 0.30	- 0.60
[E]	0.30	0.30	0.9

a) What is [A] at equilibrium? Answer 0.30M

b) What is [B] at equilibrium? Answer 0.30 M

c) What is [C] at equilibrium? Answer 0.9 M

d) Write the **expression** for the equilibrium constant, K_{eq}.

$$K_{eq} = \frac{[C]^2}{[A][B]}$$

e) Calculate the **value** for the equilibrium constant at the temperature at the experiment was done.

1 SD

$$K_{eq} = \frac{(0.9)^2}{(0.30)(0.30)} = 9$$

Answer $K_{eq} = 9$

KEY

4. Considering the following equilibrium:

$$2AB_{3(g)} \rightleftarrows A_{2(g)} + 3B_{2(g)}$$

If 0.87 moles of AB_3 are injected into a 5.0 L container at 25°C, at equilibrium the final $[A_2]$ is found to be 0.070 M.(*Hint: Make a table and use it to answer the questions below.*)

Initial $[AB_3] = \frac{0.87\text{ mol}}{5.0\text{ L}} = 0.174\text{ M}$ (limited to 2 SD's) (= 2 dps)

	$2AB_3 \xleftarrow{2/1}$ ⇌	$A_2 \xrightarrow{3/1}$ +	$3B_2$
[I]	0.174 (2dp)	0	0
[C]	−0.14	+0.070	+0.21
[E]	0.034 (2dp =1SD)	0.070	0.21

a) Calculate the equilibrium concentration of AB_3. Answer 0.03 M

b) Calculate the equilibrium $[A_2]$. Answer 0.070 M

c) Calculate the equilibrium $[B_2]$. Answer 0.21 M

5. Consider the reaction:

$$A_{(g)} + B_{(g)} \rightleftarrows C_{(g)}$$

a) In an equilibrium mixture the following concentrations were found: (2 SD)

[A] = 0.45M, [B] = 0.63M and [C] = 0.30M. Calculate the value of the equilibrium constant for this reaction.

$K_{eq} = \frac{[C]}{[A][B]} = \frac{(0.30)}{(0.45)(0.63)} = 1.0582$

Answer $K_{eq} = 1.1$

b) At the same temperature, another equilibrium mixture is analyzed and it is found that [B] = 0.21 M and [C] = 0.70 M. From this and the information above, calculate the *equilibrium [A].* (use unrounded value for keq)

$1.0582 = \frac{(0.70)}{[A](0.21)}$ so $[A] = \frac{(0.70)}{(1.0582)(0.21)} = 3.2\text{M}$

$\boxed{\text{Answer } [A] = 3.2\text{M}}$

c) In another equilibrium mixture at the same temperature, it is found that [A] = 0.35 M and the [C] = 0.86 M. From this and the information above, calculate the *equilibrium [B]*.

$$K_{eq} = \frac{[C]}{[A][B]} \qquad A_{(g)} + B_{(g)} \rightleftarrows C_{(g)}$$

$$1.0582 = \frac{(0.86)}{(0.35)[B]} \rightarrow [B] = \frac{(0.86)}{1.0582(0.35)} = 2.3M$$

Answer $\boxed{[B] = 2.3M}$

6. Two mole of gaseous NH_3 are introduced into a 1.0 L vessel and allowed to undergo partial decomposition at high temperature according to the reaction:

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

At equilibrium, 1.0 mole of $NH_{3(g)}$ remains.
(*Make a table and use it to answer the questions below:*)

$$2NH_3 \xrightarrow{1/2} N_2 \quad + \quad 3H_2 \quad (3/2)$$

	$2NH_3$	N_2	$3H_2$
[I]	2.0	0	0
[C]	-1.0	+0.50	+1.5
[E]	1.0	0.50	1.5

a) What is the equilibrium $[N_2]$? Answer 0.50 M

b) What is the equilibrium $[H_2]$? Answer 1.5 M

c) Calculate the **value** of the equilibrium constant at the temperature of the experiment.

$$K_{eq} = \frac{[N_2][H_2]^3}{[NH_3]^2} = \frac{(0.50)(1.5)^3}{(1.0)^2} = 1.7$$

Answer $\boxed{K_{eq} = 1.7}$

7. At a high temperature, 0.50 mol of HBr was placed in a 1.0 L container and allowed to decompose according to the reaction:

$$2HBr_{(g)} \rightleftarrows H_{2(g)} + Br_{2(g)}$$

At equilibrium the $[Br_2]$ was measured to be 0.13 M. What is K_{eq} for this reaction at this temperature?

	$2HBr$	H_2	Br_2
[I]	0.50	0	0
[C]	−0.26	+0.13	+0.13
[E]	0.24	0.13	0.13

$$K_{eq} = \frac{[H_2][Br_2]}{[HBr]^2} = \frac{(0.13)^2}{(0.24)^2} = 0.29$$

Answer $K_{eq} = 0.29$

8. When 1.0 mol of $NH_{3(g)}$ and 0.40 mol of $N_{2(g)}$ are placed in a 5.0 L vessel and allowed to reach equilibrium at a certain temperature, it is found that 0.78 mol of NH_3 is present. The reaction is:

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

initial $[NH_3] = \frac{1.0\ mol}{5.0\ L} = 0.20M$ initial $[N_2] = \frac{0.40\ mol}{5.0\ L} = 0.080M$

equil^m $[NH_3] = \frac{0.78\ mol}{5.0\ L} = 0.156\ M$

	$2NH_3$	$3H_2$	N_2
[I]	0.20	0	0.080
[C]	−0.044	+0.066	+0.022
[E]	0.156	0.066	0.102

a) Calculate the **equilibrium concentrations** of all three species.

$[NH_3]$ = 0.16 M $[H_2]$ = 0.066 M $[N_2]$ = 0.10M

b) Calculate the **value** of the equilibrium constant at this temperature.
(use unrounded concs, then round to 2 SD's)

$$K_{eq} = \frac{[H_2]^3[N_2]}{[NH_3]^2} = \frac{(0.066)^3(0.102)}{(0.156)^2} = 0.0012$$

Answer 1.2×10^{-3}

c) How many **moles** of H_2 are present at equilibrium?

$0.066\ M \times 5.0\ L = 0.33\ mol$

Answer 0.33 mol

d) How many **moles** of N_2 are present at equilibrium?

$0.102\ M \times 5.0\ L = 0.51\ mol$

Answer 0.51 mol

9. When 0.40 mol of PCl_5 is heated in a 10.0 L container, an equilibrium is established in which 0.25 mol of Cl_2 is present. (*Make a table and answer the questions below. Be sure to read all questions a-d before making your table!:*) Initial $[PCl_5] = \frac{0.40\ mol}{10.0\ L} = 0.040M$

	$PCl_{5(g)}$ ⇄	$PCl_{3(g)}$ +	$Cl_{2(g)}$
[I]	0.040	0	0
[C]	-0.025	+0.025	+0.025
[E]	0.015	0.025	0.025

equil'm $[Cl_2] = \frac{0.25 mol}{10.0L} = 0.025M$

a) Calculate the **equilibrium concentration** of each species.

$[PCl_5]$ = 0.015M $[PCl_3]$ = 0.025M $[Cl_2]$ = 0.025M

b) Calculate the **value** of the equilibrium constant, K_{eq} at the temperature of the experiment.

$$K_{eq} = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{(0.025)^2}{(0.015)} = 0.042$$

Answer $K_{eq} = 0.042$

c) What **amount** (moles) of PCl_3 is present at equilibrium?

$0.025M \times 10.0L = 0.25 mol$

Answer 0.25 mol

d) What **amount** (moles) of PCl_5 is present at equilibrium?

$0.015M \times 10.0L = 0.15 mol$

Answer 0.15 mol

10. A mixture of H_2 and I_2 is allowed to react at 448°C. When *equilibrium* is established, the concentrations of the participants are found to be:

$[H_2] = 0.46$ M, $[I_2] = 0.39$ M and $[HI] = 3.0$ M.

The equation is: $H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)}$

a) Calculate the **value** of K_{eq} at 448°C.

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]} = \frac{(3.0)^2}{(0.46)(0.39)} = 50.167$$

Answer 50. or 5.0×10^1

b) In another equilibrium mixture of the *same* participants at 448°C, the concentrations of I_2 and H_2 are both 0.050 M. What is the *equilibrium concentration* of HI?

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]} \rightarrow 50.167 = \frac{[HI]^2}{(0.050)^2}$$

$$[HI]^2 = 50.167 \times (0.050)^2 \rightarrow [HI] = 0.35 M$$
$$[HI]^2 = 0.1254$$

Answer 0.35 M

11. The K_{eq} for the reaction:

$$PCl_{5(g)} \rightleftarrows PCl_{3(g)} + Cl_{2(g)}$$

at 250°C is found to be ***0.042***. In an *equilibrium mixture* of these species, it is found that $[PCl_5] = 0.012$ M, and $[Cl_2] = 0.049$ M. What is the equilibrium $[PCl_3]$ at 250°C ?

$$K_{eq} = \frac{[PCl_3][Cl_2]}{[PCl_5]} \rightarrow [PCl_3] = \frac{(0.042)(0.012)}{(0.049)} = 0.010 M$$

$$0.042 = \frac{[PCl_3](0.049)}{(0.012)}$$

Answer 0.010 M

12. At a certain temperature the reaction:

$$CO_{(g)} + 2H_{2(g)} \rightleftarrows CH_3OH_{(g)}$$

has a Keq = ***0.500***. If a reaction mixture at equilibrium contains 0.210 M CO and 0.100 M H_2, what is the *equilibrium* $[CH_3OH]$?

$$K_{eq} = \frac{[CH_3OH]}{[CO][H_2]^2} \qquad [CH_3OH] = (0.500)(0.210)(0.100)^2$$
$$= 0.00105 M$$

$$0.500 = \frac{[CH_3OH]}{(0.210)(0.100)^2}$$

Answer 0.00105 M or 1.05×10^{-3} M

13. At a certain temperature the reaction: $CO_{(g)} + H_2O_{(g)} \rightleftarrows CO_{2(g)} + H_{2(g)}$

has a K_{eq} = ***0.400***. Exactly *1.00 mol* of each gas was placed in a 100.0 L vessel and the mixture was allowed to react. Find the **equilibrium concentration** of each gas.

initial [CO], [H_2O], [CO_2] & [H_2] = $\frac{1.00\ mol}{100.0\ L}$ = 0.0100M

Trial Keq = 1.00
Keq = 0.400, so rx. will shift LEFT as it approaches equilm

$CO + H_2O \rightleftarrows CO_2 + H_2$

	CO	H_2O	CO_2	H_2
[I]	0.0100	0.0100	0.0100	0.0100
[C]	+x	+x	−x	−x
[E]	0.0100+x	0.0100+x	0.0100−x	0.0100−x
[E]	0.0100 +0.002251	0.0100 +0.002251	0.0100 −0.002251	0.0100 −0.002251
[E]	0.0123	0.0123	0.0078	0.0078

$K_{eq} = \frac{[CO_2][H_2]}{[CO][H_2O]}$

$0.400 = \frac{(0.0100-x)^2}{(0.0100+x)^2}$

$\sqrt{0.400} = \frac{(0.0100-x)}{0.0100+x}$

$0.63246(0.0100+x) = 0.0100 - x$

$0.0063246 + 0.63246x = 0.0100 - x$

$1.63246x = 0.0100 - 0.0063246$

$1.63246x = 0.0036754$

$x = \frac{0.0036754}{1.63246}$

$x = 0.00225145$

(3 SD's = 5 dps)

Answer [CO] = [H_2O] = 0.0123 M, [CO_2] = [H_2] = 0.0078 M

14. The reaction: $2XY_{(g)} \rightleftarrows X_{2(g)} + Y_{2(g)}$

has a K_{eq} = ***35*** at 25°C. If 3.0 moles of XY are injected into a 1.0 L container at 25°C, find the equilibrium [X_2] and [Y_2].

Trial Keq = $\frac{[X_2][Y_2]}{[XY]^2} = \frac{0}{(3.0)^2} = 0$ (Keq = 35)
so rx. will shift RIGHT as it approaches equilm

$2XY \rightleftarrows X_2 + Y_2$

	XY	X_2	Y_2
[I]	3.0	0	0
[C]	−2x	+x	+x
[E]	3.0−2x	x	x
[E]	3.0−2(1.383)	1.383	1.383
[E]	0.2338	1.383	1.383

$K_{eq} = \frac{[X_2][Y_2]}{[XY]^2}$

$35 = \frac{x^2}{(3.0-2x)^2}$

$\sqrt{35} = \frac{x}{3.0-2x}$

$5.916(3.0-2x) = x$

$17.7482 - 11.832x = x$

$12.832x = 17.7482$

$x = \frac{17.7482}{12.832}$

$x = 1.383$

Answer [X_2] = 1.4M [Y_2] = 1.4M

15. The equilibrium constant for the reaction:

$$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)} \qquad \text{at } 448°C \text{ is } \mathbf{50}.$$

a) If 1.0 mol of H_2 is mixed with 1.0 mol of I_2 in a 0.50 L container and allowed to react at 448°C , what is the **equilibrium** [HI]?

Initial $[H_2] = \frac{1.0\,mol}{0.50\,L} = 2.0\,M$, $[I_2] = 2.0\,M$, $[HI] = 0$

$H_2 + I_2 \rightleftharpoons 2HI$

[I]	2.0	2.0	0
[C]	−x	−x	+2x
[E]	2.0−x	2.0−x	2x
[E]	2.0−1.559	2.0−1.559	2(1.559)
[E]	0.44	0.44	3.1

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

$$50. = \frac{(2x)^2}{(2.0-x)^2}$$

$$\sqrt{50.} = \frac{2x}{2-x}$$

$7.071(2-x) = 2x$

$14.142 - 7.071x = 2x$

$14.142 = 9.071x$

$x = \frac{14.142}{9.071} = 1.559$

Answer $[HI] = 3.1\,M$

b) How many **moles** of HI are formed at equilibrium? (Actual yield)

$3.118\,M \times 0.50\,L = 1.6\,mol$

Answer 1.6 moles of HI

16. Given K_{eq} for the reaction:

$$PCl_{5(g)} \rightleftharpoons PCl_{3(g)} + Cl_{2(g)}$$

is ***0.042*** at 250°C , what will happen if 2.50 mol of PCl_5, 0.600 mol of Cl_2 and 0.600 mol of PCl_3 are placed in a 1.00 flask at 250°C ? (*Will the reaction shift left, right, or not occur at all?*)

$$K_{eq} = \frac{[PCl_3][Cl_2]}{[PCl_5]}$$

$$\text{Trial } K_{eq} = \frac{(0.600)(0.600)}{(2.50)}$$

Trial $K_{eq} = 0.144$

Actual $K_{eq} = 0.042$

since Trial Keq > Keq

The reaction will shift to the LEFT as it approaches equilm. So $[PCl_5]$ will increase while $[PCl_3]$ and $[Cl_2]$ decrease.

Answer shift LEFT

17. Given the equilibrium equation: $H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)}$

at 448°C, $K_{eq} = \mathbf{50.}$ If 3.0 mol of HI, 2.0 mol of H_2, and 1.5 mol of I_2 are placed in a 1.0 L container at 448°C, will a reaction occur?

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

$$\text{Trial } K_{eq} = \frac{(3.0)^2}{(2.0)(1.5)} = 3.0$$

Trial K_{eq} (3.0) < K_{eq} (50.) so the reaction will proceed to the RIGHT in order to reach equilibrium.

Answer $\boxed{\text{Yes, it will}}$

If so, which way does the reaction shift? $\boxed{\text{to the RIGHT}}$

18. Given the equilibrium equation: $H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)}$

at 448°C, $K_{eq} = \mathbf{50.}$ If 5.0 mol of HI, 0.7071 mol of H_2, and 0.7071 mol of I_2 are placed in a 1.0 L container at 448°C, will a reaction occur? (*Round any answers off to 3 significant digits!*)

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

$$\text{Trial } K_{eq} = \frac{(5.0)^2}{(0.7071)^2} = \underline{50.}$$

since Trial K_{eq} = K_{eq}

Answer $\boxed{\text{Rx. will NOT OCCUR}}$

If so, which way does the reaction shift? $\boxed{\text{neither way.}}$

19. Determine the equilibrium constant for the reaction: $H_{2(g)} + I_{2(g)} \rightleftarrows 2HI_{(g)}$ given that an equilibrium mixture is analyzed and found to contain the following concentrations: $[H_2] = 0.0075$ M, $[I_2] = 0.000043$ M and $[HI] = 0.0040$ M

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]} = \frac{(0.0040)^2}{(0.0075)(0.000043)} = 50.$$

Answer $\boxed{K_{eq} = 50.}$

20. Given the equilibrium equation: $3A_{(g)} + B_{(g)} \rightleftharpoons 2C_{(g)}$

If 2.50 moles of A and 0.500 moles of B are added to a 2.00 L container, an equilibrium is established in which the [C] is found to be 0.250 M.

a) Find [A] and [B] at equilibrium.

initial $[A] = \frac{2.50\,mol}{2.00\,L} = 1.25\,M$, $[B] = \frac{0.500\,mol}{2.00\,L} = 0.250\,M$

3/2 1/2

	3A	+ B	⇄ 2C
[I]	1.25	0.250	0
[C]	−0.375	−0.125	+0.250
[E]	0.875	0.125	0.250

2 dec. places

Answer $[A] = 0.88\,M$ $[B] = 0.125\,M$

b) Calculate the value of the equilibrium constant K_{eq}.

$$K_{eq} = \frac{[C]^2}{[A]^3[B]} = \frac{(0.250)^2}{(0.875)^3(0.125)} = 0.746$$

Answer $K_{eq} = 0.75$

21. At 800°C, the equilibrium constant K_{eq}, for the reaction:

$CO_{2(g)} + H_{2(g)} \rightleftharpoons CO_{(g)} + H_2O_{(g)}$ is 0.279

If 1.50 moles of CO_2 and 1.50 moles of H_2 are added to a 1.00 L container, what would the [CO] be at equilibrium?

	CO_2	+ H_2	⇄ CO	+ H_2O
[I]	1.50	1.50	0	0
[C]	−x	−x	+x	+x
[E]	1.50 − x	1.50 − x	x	x
[E]	1.50−0.5185	1.50−0.5185	0.5185	0.5185
[E]	0.982	0.982	0.518	0.518

$$K_{eq} = \frac{[CO][H_2O]}{[CO_2][H_2]}$$

$$0.279 = \frac{x^2}{(1.50-x)^2}$$

$$\sqrt{0.279} = \frac{x}{1.50-x}$$

$$0.5282(1.50-x) = x$$

$$0.7923 - 0.5282x = x$$

$$0.7923 = 1.5282x$$

$$x = \frac{0.7923}{1.5282}$$

$$x = 0.5185$$

Answer equil'm $[CO] = 0.518\,M$

22. Given that the equilibrium constant K_{eq} for the reaction:

$$A_{(g)} + B_{(g)} \rightleftarrows C_{(g)} + D_{(g)} \quad \text{is } \mathbf{0.015} \text{ at } 25^oC,$$

if 1.0 mole of each gas is added to a 1.0 L container at 25°C, which way will the equation shift in order to reach equilibrium?

$$K_{eq} = \frac{[C][D]}{[A][B]}$$

$$\text{Trial } K_{eq} = \frac{(1.0)(1.0)}{(1.0)(1.0)} = 1.0$$

Trial keq > keq so reaction will shift LEFT to reach equil$^{\underline{m}}$.
(1.0) (0.015)

Answer LEFT

23. Calculate the **equilibrium constant** K_{eq} for the following reaction:

$$2A_{2(g)} + 3B_{2(g)} \rightleftarrows 2A_2B_{3(g)}$$

given that the *partial pressure* of each substance at equilibrium is as follows:

Partial Pressure of A_2 = 20.0 kPa, Partial Pressure of B_2 = 30.0 kPa, Partial Pressure of A_2B_3 = 5.00 kPa.

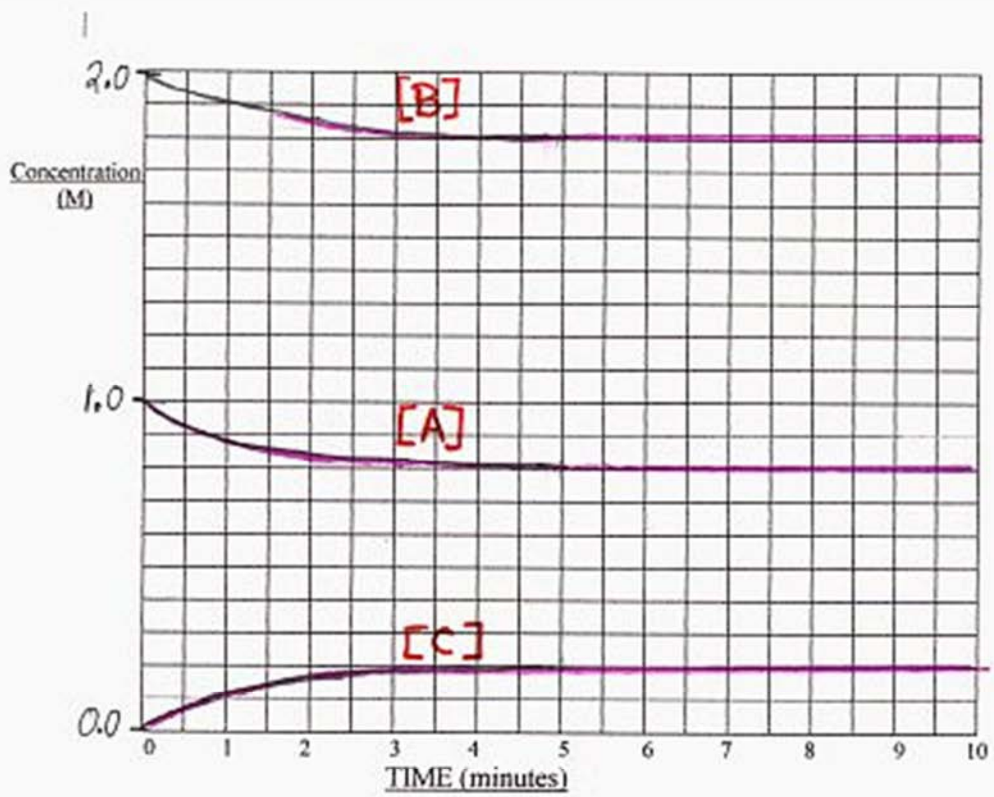
$$K_{eq} = \frac{P_{A_2B_3}^{\,2}}{P_{A_2}^{\,2} \cdot P_{B_2}^{\,3}} = \frac{(5.00)^2}{(20.0)^2\,(30.0)^3} = 2.31 \times 10^{-6}$$

Answer $K_{eq} = 2.31 \times 10^{-6}$

24. Consider the following equilibrium system: $A_{(g)} + B_{(g)} \rightleftarrows C_{(g)}$

1.0 mole of A and 2.0 moles of B are simultaneously injected into an empty 1.0 L container. At equilibrium (after 5.0 minutes), [C] is found to be 0.20 M. Make calculations and draw graphs to show how each of [A], [B] and [C] change with time over a period of 10.0 minutes. (*HINT: You have to make a table first.*)

	$A_{(g)}$ +	$B_{(g)}$ ⇄	$C_{(g)}$
[I]	1.0	2.0	0
[C]	−0.20	−0.20	+0.20
[E]	0.80	1.80	0.20



25. Given the reaction:

$$4HCl_{(g)} + O_{2(g)} \rightleftarrows 2H_2O_{(g)} + 2Cl_{2(g)} \quad \Delta H = -113 \text{ kJ}$$

How will the value of the equilibrium constant K_{eq} at 550°C compare with it's value at 450°C? Keq at 550°c will have a lower value.

Explain your answer. Since the rx. is exothermic, adding heat will cause the reaction to shift LEFT. This will decrease the value of Keq.

26. The following system is at equilibrium, in a closed container:

7 mog. 8 mog

$$4NH_{3(g)} + 3O_{2(g)} \rightleftarrows 6H_2O_{(g)} + 2N_{2(g)} + \text{Heat}$$

a) How is the *amount of N_2* in the container affected if the ***volume*** of the container is ***doubled***? Increased (decreasing P will cause a shift RIGHT producing more N_2)

b) How is the rate of the **forward reaction** affected if more water vapor is introduced into the container? Increase (First inc. $[H_2O]$ will speed up rev. rx, forming more reactants. Then forward rx. will speed up.)

c) How is the amount of O_2 in the container affected if a *catalyst* is added?

No change (A catalyst speeds up forward & rev. rx. equally)

27. At a certain temperature, K_{eq} for the reaction:

$$3C_2H_2 \rightleftarrows C_6H_6 \text{ is } \textbf{5.0}.$$

If the *equilibrium concentration* of C_2H_2 is 0.40 moles/L, what is the *equilibrium concentration* of C_6H_6?

$$K_{eq} = \frac{[C_6H_6]}{[C_2H_2]^3} \rightarrow 5.0 = \frac{[C_6H_6]}{(0.40)^3} \rightarrow [C_6H_6] = (5.0)(0.40)^3 = 0.32M$$

Answer 0.32 M