Chemistry 12

K_{eq} Calculations Worksheet

Name: Key
Date:
Block:

1. Given the equilibrium equation below:

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

If, at equilibrium, the concentrations are as follows:

$[A_2] = 3.45$ 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.

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

2. For the reaction:

$$A_{2(g)} + B_{(g)} \leftrightharpoons 2C_{(g)}$$

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.

	A_2	+ B_2	⇌ 2C
I	0M	0M	1.5M
C	+0.30M	+0.30M	−0.60M
E	0.30M	0.30M	0.90M

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

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

c) What is [C] at equilibrium? 0.90M

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

$$K_{eq} = \frac{[C]^2}{[A_2][B]}$$

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

$$K_{eq} = \frac{(0.90)^2}{(0.30)(0.30)} = \boxed{9.0}$$

3. Considering the following equilibrium:

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

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.

	$2AB_3$	A_2	$3B_2$
I	0.174M	0M	0M
C	−0.14M	+0.070M	+0.21M
E	0.034M	0.070M	0.21M

a) Calculate the equilibrium $[AB_3]$. 0.034M

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

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

4. Consider the reaction:

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

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

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

$$K_{eq} = \frac{[C]}{[A][B]} = \frac{(0.30)}{(0.45)(0.63)} = \boxed{1.1}$$

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].

$$1.1 = \frac{(0.70)}{[A](0.21)}$$

$$[A] = \boxed{3.0M}$$

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].

$$1.1 = \frac{(0.86)}{(0.35)[B]}$$

$$[B] = \boxed{2.2M}$$

5. 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:

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

At equilibrium, 1.0 mole of $NH_{3(g)}$ remains.

	$2NH_3$	$\rightleftharpoons$	N_2	+	$3H_2$
I	2.0M		0M		0M
C	−1.0M		+0.50M		+1.5M
E	1.0M		0.50M		1.5M

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

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

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

$$K_{eq} = \frac{(1.5)^3(0.50)}{(1.0)^2} = \boxed{1.7}$$

6. 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:

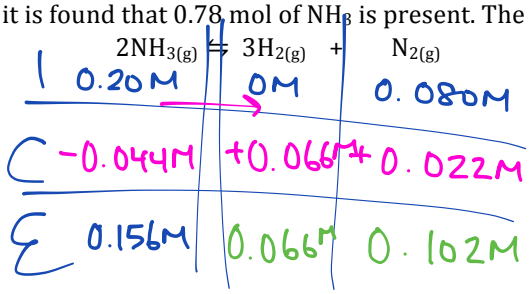
$$2HBr_{(g)} \leftrightharpoons H_{2(g)} + Br_{2(g)}$$

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

	$2HBr$	$\rightleftharpoons$	H_2	+	Br_2
I	0.50M		0M		0M
C	−0.26M		+0.13M		+0.13M
E	0.24M		0.13M		0.13M

$$K_{eq} = \frac{(0.13)(0.13)}{(0.24)^2} = \boxed{0.29}$$

7. 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:



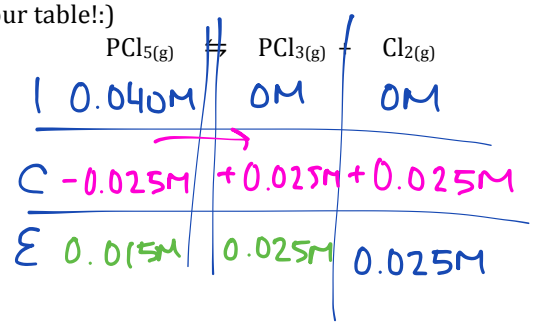
a) Calculate the equilibrium concentrations of all three species.

$[NH_3]$ = 0.156M $[H_2]$ = 0.066M $[N_2]$ = 0.102M

b) Calculate the value of the equilibrium constant at this temperature.

$$K_{eq} = \frac{(0.102)(0.066)^3}{(0.156)^2} = 1.2 \cdot 10^{-3}$$

8. 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!:)



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.

$$K_{eq} = \frac{(0.025)(0.025)}{(0.015)} = 4.2 \cdot 10^{-2}$$

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

$$\frac{0.025 mol}{1 L} \times \frac{10.0 L}{1} = 0.25 mol\ PCl_3$$

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

$$\frac{0.015 mol}{1 L} \times \frac{10.0 L}{1} = 0.15 mol\ PCl_5$$

9. 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:

$$H_{2(g)} + I_{2(g)} \leftrightharpoons 2HI_{(g)}$$

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

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]} = \frac{(3.0)^2}{(0.46)(0.39)} = \boxed{50.}$$

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

$$50. = \frac{[HI]^2}{(0.050)^2}$$

$$[HI] = \sqrt{(50.)(0.050)^2}$$

$$= \boxed{0.35M}$$

Eq V

10. At a certain temperature the reaction:

$$CO_{(g)} + H_2O_{(g)} \leftrightharpoons CO_{2(g)} + H_{2(g)}$$

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.

trial $K_{eq} = \frac{(0.01)^2}{(0.01)^2} = 1$

trial > K_{eq} ∴ shift left

	CO	+ H_2O	⇌ CO_2	+ H_2
I	0.01M	0.01M	0.01M	0.01M
C	+x	+x	-x	-x
E	0.01+x	0.01+x	0.01-x	0.01-x

$$\sqrt{0.400} = \sqrt{\frac{(0.01-x)^2}{(0.01+x)^2}}$$

$$0.63245 = \frac{(0.01-x)}{(0.01+x)}$$

$$0.0063245 + 0.63245x = 0.01 - x$$

$$1.63245x = 0.0036755$$

$$x = 2.2515 \times 10^{-3}$$

$$\boxed{[CO] = [H_2O] = 1.23 \cdot 10^{-2} M} \quad \boxed{[CO_2] = [H_2] = 7.75 \cdot 10^{-3} M}$$

11. The reaction:

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

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 $[Y_2]$.

	2XY	⇌	X_2	+	Y_2
I	3.0M		0M		0M
C	−2x		+x		+x
E	3.0−2x		x		x

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

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

$$5.916 = \frac{x}{3.0-2x}$$

$$17.748 - 11.832x = x$$

$$17.748 = 12.832x$$

$$1.38M = x$$

$$[X_2] = [Y_2] = 1.38M$$

12. The equilibrium constant for the reaction:

$$H_{2(g)} + I_{2(g)} \leftrightharpoons 2HI_{(g)}$$ at 448°C is 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]?

	H_2	+	I_2	⇌	2HI
I	2.0M		2.0M		0M
C	−x		−x		+2x
E	2.0−x		2.0−x		2x

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

$$x = 1.56M$$

$$[HI] = 2(1.56)$$

$$= 3.12M$$

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

$$\frac{3.12\,mol}{1L} \times \frac{0.50L}{1} = 1.56\,mol\ HI$$

Equil IV

13. Given K_{eq} for the reaction:

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

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?)

$$\text{Trial } K_{eq} = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{(0.600)(0.600)}{(2.50)}$$

$$= 0.144 > 0.042$$

Trial K_{eq} Actual K_{eq}

Shift left towards reactants

14. Given the equilibrium equation:

$$H_{2(g)} + I_{2(g)} \leftrightharpoons 2HI_{(g)}$$

at 448°C, K_{eq} = 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, which way does the reaction shift?

$$\text{Trial } K_{eq} = \frac{[HI]^2}{[H_2][I_2]} = \frac{(3.0)^2}{(2.0)(1.5)}$$

$$= 3 < 50$$

Trial K_{eq} Actual K_{eq}

Shift right towards products

15. Given the equilibrium equation:

$$H_{2(g)} + I_{2(g)} \leftrightharpoons 2HI(g)$$

at 448°C, Keq = 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, which way does the reaction shift?

$$\text{Trial } K_{eq} = \frac{[HI]^2}{[H_2][I_2]} = \frac{(5.0)^2}{(0.7071)^2}$$

$$= 50 = 50$$

Trial K_{eq} Actual K_{eq}

The reaction is <u>at</u> equilibrium and will not shift

16. Determine the equilibrium constant for the reaction:

$$H_{2(g)} + I_{2(g)} \leftrightharpoons 2HI_{(g)}$$

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

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

$$= 49.6 = \boxed{50.}$$

17. Given the equilibrium equation:

$$3A_{(g)} + B_{(g)} \leftrightharpoons 2C_{(g)}$$

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.

	3A	+ B	⇌ 2C
I	1.25M	0.250M	0M
C	−0.375M	−0.125M	+0.250M
E	0.875M	0.125M	0.250M

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

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

18. At 800°C, the equilibrium constant $K_{eq} = 0.279$ for the reaction:

$$CO_{2(g)} + H_{2(g)} \leftrightharpoons CO_{(g)} + H_2O_{(g)}$$

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	1.50M	1.50M	0M	0M
C	−x	−x	+x	+x
E	1.50−x	1.50−x	x	x

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

$$x = \boxed{0.518M = [CO]}$$

19. Given that the equilibrium constant K_{eq} = 0.015 at 25°C for the reaction:

$$A_{(g)} + B_{(g)} \leftrightharpoons C_{(g)} + D_{(g)}$$

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?

1.0 > 0.015

Trial Keq Actual Keq

Shift left towards reactants

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

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

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.

21. Given the reaction:

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

How will the value of the equilibrium constant K_{eq} at 550°C compare with its value at 450°C? Explain your answer.

The reaction is exothermic. Increasing the temp to 550°C from 450°C will shift the equilibrium to the left and the Keq value will decrease.

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

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

7 8

a) How is the amount of N_2 in the container affected if the volume of the container is doubled?

If volume ↑, the rxn will shift right and $[N_2]$ ↑ (pressure ↓)

b) How is the rate of the forward reaction affected if more water vapor is introduced into the container?

$[H_2O]$ ↑

The forward rate would ↑ as $[NH_3]$ and $[O_2]$ ↑

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

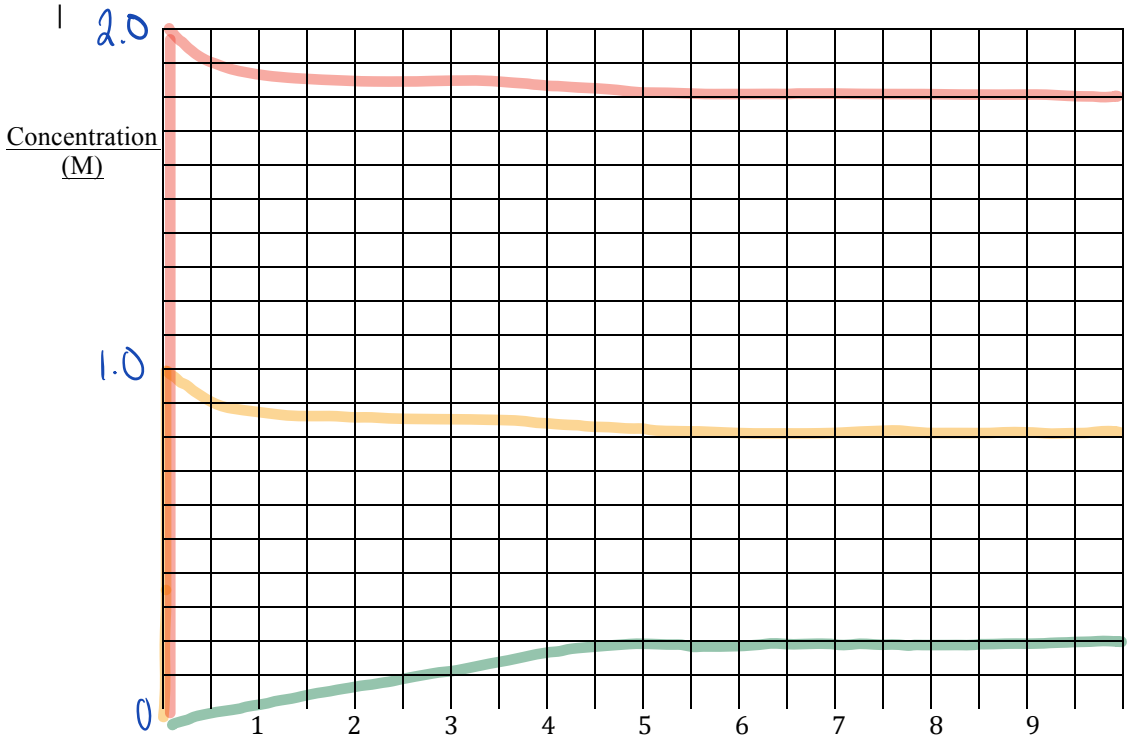
Not affected

23. Consider the following equilibrium system:

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

1.0 mole of A and 2.0 moles of B are simultaneously injected into an empty 1.0 L container. After 5.0 minutes, equilibrium is reached and [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	+ B	⇌ C
I	1.0M	2.0M	0M
C	−0.20M	−0.20M	+0.20M
E	0.80M	1.8M	0.20M



[A]

[B]

[C]