CALCULATIONS GIVEN K_{eq}

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Example #1

$$CH_{4(g)} + H_2O_{(g)} <====> CO_{(g)} + 3\ H_{2(g)}$$

At 1500°C, $K_{eq} = 5.67$

$[CO] = 0.300$ M, $[H_2] = 0.800$ M and

$[CH_4] = 0.400$ M

What is $[H_2O]$ at equilibrium?

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Example #1

$$CH_{4(g)} + H_2O_{(g)} <===> CO_{(g)} + 3\ H_{2(g)}$$

At 1500°C, K_{eq} = 5.67, [CO] = 0.300 M, $[H_2]$ = 0.800 M and, $[CH_4]$ = 0.400 M

What is $[H_2O]$ at equilibrium?

$$K_{eq} = \frac{[CO][H_2]^3}{[CH_4][H_2O]}$$

$$[H_2O] = \frac{[CO][H_2]^3}{[CH_4]\ K_{eq}}$$

$$[H_2O] = \frac{[0.300M][0.800M]^3}{[0.400M]\ 5.67}$$

$$[H_2O] = 0.0677M$$

.: $[H_2O]$ at equilibrium was $6.77\text{x}10^{-2}$M

CALCULATIONS GIVEN K_{eq}

Example #2

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

What is $[NH_3]$ when $[N_2] = 0.45$ M, $[H_2] = 1.10$ and $K_{eq} = 1.7 \times 10^{-2}$?

CALCULATIONS GIVEN K_{eq}

Example #2

$$N_{2(g)} + 3\ H_{2(g)} <====> 2\ NH_{3(g)}$$

What is $[NH_3]$ when $[N_2] = 0.45$ M, $[H_2] = 1.10$ M and Keq $= 1.7 \times 10^{-2}$?

$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

$$K_{eq}\ [N_2][H_2]^3 = [NH_3]^2$$

$$0.017\ [0.45M][1.10M]^3 = [NH_3]^2$$

$$0.01018215 = [NH_3]^2$$

$$\sqrt{0.01018215} = \sqrt{[NH_3]^{\cancel{2}}}$$

$$0.10090664 = [NH_3]$$

.: the $[NH_3]$ is 1.0×10^{-1}M

CALCULATIONS GIVEN K_{eq}

Example #3

$$CO_{(g)} + H_2O_{(g)} \text{ <====> } CO_{2(g)} + H_{2(g)}$$

At equilibrium, Keq = 4.06.

If 0.100 mol of CO and 0.100 mol of $H_2O_{(g)}$ are placed in a 1.00 L container,

a) What are the concentrations of the reactants and products at equilibrium?

b) What is the final mass of $CO_{2(g)}$?

CALCULATIONS GIVEN K_{eq}

Example #3

At equilibrium, Keq = 4.06.

If 0.100 mol of CO and 0.100 mol of $H_2O_{(g)}$ are placed in a 1.00 L container,

a) What are the concentrations of the reactants and products at equilibrium?

b) What is the final mass of $CO_{2(g)}$?

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

	$CO_{(g)}$	$H_2O_{(g)}$	$CO_{2(g)}$	$H_{2(g)}$
I	0.100	0.100	0	0
C	–x	-x	+x	+x
E	0.100-x	0.100-x	x	x

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

$$4.06 = \frac{[x][x]}{[0.100-x][0.100-x]}$$

CALCULATIONS GIVEN K_{eq}

	$CO_{(g)}$ +	$H_2O_{(g)}$ <=>	$CO_{2(g)}$ +	$H_{2(g)}$
I	0.100	0.100	0	0
C	–x	-x	+x	+x
E	0.100-x	0.100-x	x	x

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

$$4.06 = \frac{[x][x]}{[0.100\text{-}x][0.100\text{-}x]}$$

$$4.06\,[0.100-x][0.100-x] = x^2$$

$$(0.406-4.06x)[0.100-x] = x^2$$

$$0.0406 - 0.406x - 0.406x + 4.06x^2 = x^2$$

$$3.06x^2 - 0.812x + 0.0406 = 0$$

Time to use the quadratic formula!

CALCULATIONS GIVEN K_{eq}

	$CO_{(g)}$	+	$H_2O_{(g)}$	<=>	$CO_{2(g)}$	+	$H_{2(g)}$
I	0.100		0.100		0		0
C	−x		-x		+x		+x
E	0.100-x		0.100-x		x		x

$$3.06x^2 - 0.812x + 0.0406 = 0$$

Quadratic equation: $ax^2 + bx + c = 0$

Quadratic formula: $x = \dfrac{-b \pm \sqrt{(b^2-4ac)}}{2a}$

$$x = \frac{-(-0.812) \pm \sqrt{[-0.812^2-4(3.06)(0.0406)]}}{2(3.06)}$$

$$x = 0.19853 \text{ or } 0.066832$$

This is the correct answer. Otherwise you may end up with negative reactant concentrations

CALCULATIONS GIVEN K_{eq}

Example #3

a) What are the concentrations of the reactants and products at equilibrium?

b) What is the final mass of $CO_{2(g)}$?

	$CO_{(g)}$	+	$H_2O_{(g)}$	<=>	$CO_{2(g)}$	+	$H_{2(g)}$
I	0.100		0.100		0		0
C	−x		-x		+x		+x
E	0.100-x		0.100-x		x		x

$$x = \mathit{0.066832}$$

E	0.0332	0.0332	0.066832	0.066832

a) .: $[CO_{(g)}]=3.32\text{x}10^{-2}M$, $[H_2O_{(g)}]=3.32\text{x}10^{-2}M$, $[CO_{2(g)}]=6.68\text{x}10^{-2}M$, $[H_{2(g)}]=6.68\text{x}10^{-2}M$

CALCULATIONS GIVEN K_{eq}

Example #3

If 0.100 mol of CO and 0.100 mol of $H_2O_{(g)}$ are placed in a 1.00 L container,

a) What are the concentrations of the reactants and products at equilibrium?

b) What is the final mass of $CO_{2(g)}$?

	$CO_{(g)}$	+	$H_2O_{(g)}$	<=>	$CO_{2(g)}$	+	$H_{2(g)}$
E	0.0332		0.0332		0.066832		0.066832

$$[CO_{2(g)}]=6.68\times10^{-2}M$$

$$C=\frac{n}{V}$$

$$n=CV$$

$$n=0.066832mol/L \times 1.00L$$

$$n=0.066832mol$$

$$n=\frac{m}{M}$$

$$m=nM$$

$$m= 0.066832mol \times 44.01g/mol$$

$$m=2.94g$$

.: 2.94g of $CO_{2(g)}$ were produced

CALCULATIONS GIVEN K_{eq}

Example #4

$$H_{2(g)} + I_{2(g)} \Longleftrightarrow 2\ HI_{(g)} \qquad K_{eq} = 55.6$$

If initial $[H_2]$ = 0.200 M and initial $[I_2]$ = 0.200 M. What is [HI] at equilibrium?

CALCULATIONS GIVEN K_{eq}

Example #4

If initial $[H_2]$ = 0.200 M and initial $[I_2]$ = 0.200 M. K_{eq} = 55.6

What is [HI] at equilibrium?

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

	$H_{2(g)}$	$I_{2(g)}$	$2\ HI_{(g)}$
I	0.200	0.200	0
C	-x	-x	+2x
E	0.200-x	0.200-x	2x

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

$$55.6 = \frac{[HI]^2}{[H_2][I_2]}$$

$$55.6 = \frac{[2x]^2}{[0.200-x][0.200-x]}$$

CALCULATIONS GIVEN K_{eq}

Example #4

If initial $[H_2]$ = 0.200 M and initial $[I_2]$ = 0.200 M. K_{eq} = 55.6

What is [HI] at equilibrium?

$$H_{2(g)} \quad + \quad I_{2(g)} \quad \Leftrightarrow 2\,HI_{(g)}$$

$$55.6 = \frac{[2x]^2}{[0.200-x][0.200-x]}$$

$$\sqrt{55.6} = \sqrt{\frac{4x^2}{[0.200-x]^2}}$$

$$7.456 = \frac{2x}{0.200-x}$$

$$1.4913 - 7.4565x = 2x$$

$$1.4913 = 9.4565x$$

$$0.1577 = x$$

CALCULATIONS GIVEN K_{eq}

Example #4

If initial $[H_2]$ = 0.200 M and initial $[I_2]$ = 0.200 M. K_{eq} = 55.6

What is [HI] at equilibrium?

	$H_{2(g)}$	+	$I_{2(g)}$	<=>	$2\ HI_{(g)}$
I	0.200		0.200		0
C	-x		-x		+2x
E	0.200-x		0.200-x		2x

x = 0.1577

2x = 2 (0.1577)

2x = 0.3154M

.: the equilibrium $[HI_{(aq)}] = 3.15 \times 10^{-1}$M

CALCULATIONS GIVEN K_{eq}

Example #5

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

At equilibrium, $K_{eq} = 10.0$

A reaction vessel is found to contain

0.80 M CO, 0.050 M H_2O, 0.50 M CO_2 and 0.40 M H_2.

Determine if the reaction is at equilibrium.

CALCULATIONS GIVEN K_{eq}

Q

When testing if conditions are at equilibrium Q is the symbol used, rather than K_{eq}.

Q is the "test K_{eq}" variable.

$$Q = \frac{[\text{products}]}{[\text{reactants}]}$$

CALCULATIONS GIVEN K_{eq}

Example #5

$$CO_{(g)} + H_2O_{(g)} <===> CO_{2(g)} + H_{2(g)}$$

At equilibrium, K_{eq} = 10.0.
A reaction vessel is found to contain
0.80 M CO, 0.050 M H_2O, 0.50 M CO_2 and 0.40 M H_2.

Determine if the reaction is at equilibrium.

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

$$Q = \frac{[0.50][0.40]}{[0.80][0.050]}$$

$$Q = 5.0$$

Since $k_{eq} \neq Q$, then the reaction is not at equilibrium

.: the reaction is not at equilibrium

CALCULATIONS GIVEN K_{eq}

Example #6

An equilibrium reaction of H_2, I_2, and HI begins with initial concentrations of 1.80 M, 1.80 M and 4.40 respectively. How much HI must be added to the equilibrium concentrations to increase the concentration of I_2 to 2.00 M? K_{eq} is equal to 6.0

CALCULATIONS GIVEN K_{eq}

Example #6

An equilibrium reaction of H_2, I_2, and HI begins with initial concentrations of 1.80 M, 1.80 M and 4.40 respectively. How much HI must be added to the equilibrium concentrations to increase the concentration of I_2 to 2.00 M? K_{eq} is equal to 6.0. The volume is 1.00 L.

	$H_{2(g)}$ +	$I_{2(g)}$	<=> $2\ HI_{(g)}$
I	1.80	1.80	4.40
C	-x	-x	+2x +y
E	2.00	2.00	4.00+y

"y" represents HI added

$$x = -0.20$$

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

$$6.0 = \frac{(4.00+y)^2}{(2.00)^2}$$

$$y = \cancel{-8.899} \text{ or } 0.899$$

.: 0.899 moles of HI(g) must be added

CALCULATIONS GIVEN K_{eq}

Summarizing ICE Tables

1. Write out the balanced equation.
2. All values in the table must have mol/L units.
3. Initial [product] = 0, unless otherwise stated.
4. Changes in concentration always occur in the same stoichiometric ratio.
5. Reactants and products will change in opposite directions from each other.