IX) Type 3 K_{eq} Problems

Type 3 K_{eq} problems give some initial data and a K_{eq} constant. You must find equilibrium concentrations. OR You will be provided with some equilibrium data and a K_{eq} constant. You must find an initial concentration.

Example:

1. $H_{2(g)} + I_{2(g)} \Leftrightarrow 2HI_{(g)}$ ③ $K_{eq} = 55.6 = \frac{[HI]^2}{[H_2][I_2]}$ $\sqrt{= \frac{2x^2}{(0.200M-x)^2}} = \sqrt{55.6}$

① Knowns: $[H_2]_i = 0.200M$, $[I_2]_i = 0.200M$

What is the equilibrium [HI]?

R:	$H_{2(g)}$ +	$I_{2(g)}$ ⇔	$2HI_{(g)}$
I	0.200M	0.200M	0M
C	$-x$	$-x$	$+2x$
E	$0.200-x$	$0.200-x$	$2x$

② Let $x = \Delta[H_2]$

∴ H_2 @ equil = $0.200M - x$

$\frac{2x}{(0.200-x)} = 7.45654$ s.f.

*DO NOT round early!

$7.45654(0.200-x) = 2x$

$1.491308 - 7.45654x = 2x$

$+7.45654x$

$\frac{1.491308}{9.45654} = \frac{9.45654x}{9.45654}$

∴ $0.15770123 = x$ 3 s.f.

④ $[HI]_{eq} = (2)\cdot(0.157701) = 0.315402M$

$\boxed{∴ 0.315M}$

Example:

2. $CO_{(g)} + H_2O_{(g)} \Leftrightarrow CO_{2(g)} + H_{2(g)}$ $K_{eq} = 4.06$

① Knowns: $[CO]_i = 0.10M$, $[H_2O]_i = 0.10M$ (2 s.f.)

What are the equilibrium concentrations of all species?

R	$CO_{(g)}$ +	$H_2O_{(g)}$ ⇔	$CO_{2(g)}$ +	$H_{2(g)}$
I	0.10M	0.10M	0M	0M
C	$-x$	$-x$	$+x$	$+x$
E	$(0.10-x)$	$(0.10-x)$	x	x

③ $K_{eq} = 4.06 = \frac{[CO_2][H_2]}{[CO][H_2O]}$

$\sqrt{4.06} = \sqrt{\frac{x^2}{(0.10-x)^2}}$ $\sqrt{\ }$ to cancel x^2

$2.015 = \frac{x}{(0.10-x)}$

$2.015(0.10-x) = x$

$0.2015 - 2.015x = x$

$+2.015x$

$\frac{0.2015}{3.015} = \frac{3.015x}{3.015}$

∴ $x = 0.06683$

② Let $x = \Delta[CO]_{(g)}$

④ $0.10 - 0.06683 = 0.03317$

∴ $[CO]_{eq} = 0.033M$ $[CO_2]_{eq} = 0.067M$
$[H_2O]_{eq} = 0.033M$ $[H_2]_{eq} = 0.067M$

3 s.f

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

Example:

3. A certain amount of H_2O was placed in a 2.00L closed flask. When equilibrium was reached, the $[H_2]$ was 0.500M. If $K_{eq} = 16.0$ at this temperature, how many moles of H_2O were originally placed in the flask?

① knowns.

R	$2H_2O_{(g)}$	⇔	$2H_2{}_{(g)}$	+	$O_2{}_{(g)}$
I	x		0M		0M
C	−0.500M		+0.500M		+0.250M
E	$(x - 0.500)$		0.500M		0.250M

② Let $x = [H_2O]_i$

③ $K_{eq} = \frac{[H_2]^2[O_2]}{[H_2O]^2}$

$$16.0 = \frac{(0.500)^2(0.250)}{(x-0.500)^2}$$

$$\frac{16(x-0.500)^2}{16} = \frac{(0.500)^2(0.250)}{16}$$

$$\sqrt{(x-0.500)^2} = \sqrt{3.906 \times 10^{-3}}$$

$$(x - 0.500) = 0.0625$$

$$+0.500 \qquad +0.500$$

$$x = 0.5625\text{ M}$$

④ $[H_2O]_i = 0.5625M = \frac{n}{2.00L}$

moles of $H_2O_i = (0.5625M)(2.00L) = \boxed{1.13 \text{ moles}}$

Assignment 8: Type 3 Exercises

1. For the following reaction:

$$H_{2(g)} + CO_{2(g)} \Leftrightarrow H_2O_{(g)} + CO_{(g)} \qquad K_{eq} = 0.771$$

If 0.0100mol of H_2 and 0.0100mol of CO_2 are mixed in a 1.00L container, what are the concentrations of all substances at equilibrium?

2. For the following reaction:

$$2IBr_{(g)} \Leftrightarrow I_{2(g)} + Br_{2(g)} \qquad K_{eq} = 8.5 \times 10^{-3}$$

If 0.0600mol of IBr is placed in a 1.0L container, what are the concentrations of all three substances at equilibrium?

3. For the following reaction:

$$FeO_{(s)} + CO_{(g)} \Leftrightarrow Fe_{(s)} + CO_{2(g)} \qquad K_{eq} = 0.403$$

If 0.0500mol of CO and excess solid FeO are placed in a 1.00L container, what are the concentrations of CO and CO_2 when equilibrium has been attained?

4. For the following reaction:

$$2HCl_{(g)} \Leftrightarrow H_{2(g)} + Cl_{2(g)} \qquad K_{eq} = 4.00$$

An unknown [HCl] was added to a 2.00L flask and allowed to reach equilibrium. At equilibrium, $[H_2] = 0.200M$. How many moles of HCl was originally placed in the flask?

5. For the following reaction:

$$N_2O_{2(g)} + H_{2(g)} \Leftrightarrow N_2O_{(g)} + H_2O_{(g)} \qquad K_{eq} = 1.00$$

If 0.150mol each of N_2O and H_2O are introduced into a 1.00L flask and allowed to come to equilibrium, what concentration of N_2O_2 will be present at equilibrium?

X) Type 4 K_{eq} Problems

Type 4 problems use a TRIAL K_{eq} (called 'Q' in Hebden), which can be used to predict whether or not a system is at equilibrium. If the system is not at equilibrium, the trial K_{eq} value compared to the actual K_{eq} can be used to predict in what direction the system is shifting to attain equilibrium.

Example:

1. Is the following reaction at equilibrium? If not, in which direction must the reaction shift to reach equilibrium?

$CO_{(g)}$	+	$H_2O_{(g)}$	$\Leftrightarrow$	$CO_{2(g)}$	+	$H_{2(g)}$	$K_{eq} = 10.0$
0.80M		0.050M		0.50M		0.40M	

$$\text{Trial } K_{eq} = \frac{[CO_2][H_2]}{[CO][H_2O]} = \frac{(0.50M)(0.40M)}{(0.80M)(0.050M)} = 5.0$$

The Trial $K_{eq} \neq K_{eq}$ so the system is NOT currently at equilibrium.

Trial K_{eq} of $5.0 <$ K_{eq} of 10.0

∴ system must shift RIGHT to attain equilibrium. ([products] need to ↑ and [reactants] ↓)

Example:

2. Is the following reaction at equilibrium? If not, in which direction must the reaction shift in order to attain equilibrium?

$H_{2(g)}$	+	$I_{2(g)}$	$\Leftrightarrow$	$2HI_{(g)}$	$K_{eq} = 5.0$
2.0M		1.5M		3.0M	

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

Trial K_{eq} of $3.0 <$ K_{eq} 5.0

∴ system must shift RIGHT to attain equilibrium

(to ↑ Trial K_{eq}, [products] must ↑ and [reactants] ↓)

$Br_2 = \frac{0.60\,mol}{2.0L}$ $\quad$ $Cl_2 = \frac{0.80\,mol}{2.0L}$

$[Br_2] = 0.30M$ $\quad$ $[Cl_2] = 0.40M$

$BrCl = \frac{2.20\,mol}{2.0L}$

$[BrCl] = 1.1M$

Example:

3. The following reaction occurs in a 2.0L container:

$$Br_{2(g)} + Cl_{2(g)} \Leftrightarrow 2BrCl_{(g)} \qquad K_{eq} = 3.2 \times 10^{-2}$$

Quantities of gases were found to be as follows:

$Br_2 = 0.60$ mol, $Cl_2 = 0.80$ mol, $BrCl = 2.20$ mol

What will happen to the $[Br_2]$ as the system approaches equilibrium?

$$\text{Trial } K_{eq} = \frac{[BrCl]^2}{[Br_2][Cl]} = \frac{(1.1M)^2}{(0.30)(0.40)} = 1.0 \times 10^1$$

Trial K_{eq} of 1.0×10^1 > K_{eq} of 3.2×10^{-2} (not currently @ equilibrium)

∴ system must shift LEFT to attain equilibrium

(to ↓ Trial K_{eq}, [reactants] must ↑ and [products] ↓)

∴ shift left will ↑ $[Br_2]$

Assignment 9: Type IV Exercises

1. For the following reaction: $2O_{3(g)} \Leftrightarrow 3O_{2(g)}$ $\quad K_{eq} = 75$
   Predict the direction in which the equilibrium will shift, if any, when the following substances are introduced into a 10.0L container? 0.60mol O_3 and 3.0mol O_2
2. Consider the following reaction: $2SO_{2(g)} + O_{2(g)} \Leftrightarrow 2SO_{3(g)}$ $\quad K_{eq} = 75$
   A student places 0.50moles SO_2, 0.080mol O_2, and 1.0mol SO_3 into a 1.0L flask. The student predicts that the $[SO_2]$ will decrease as equilibrium is established. Do you agree with the student's prediction? Explain using appropriate calculations.
3. Consider the following reaction:
   $$2C_{(s)} + O_{2(g)} \Leftrightarrow 2CO_{(g)} \qquad K_{eq} = 1.20 \times 10^{-2}$$
   If 2.0mol C, 0.800mol O_2, and 0.600mol CO are placed into a 1.0L flask, in which direction will the equilibrium shift in order to achieve equilibrium? What will happen to the [C]? Show all calculations.
4. For the following reaction: $H_{2(g)} + Cl_{2(g)} \Leftrightarrow 2HCl_{(g)}$ $\quad K_{eq} = 0.15$
   Equal moles of each of the three gases are in a 1.0L vessel. What direction will the reaction shift in order to reach equilibrium?
5. *Type III & Type IV hybrid question ☺