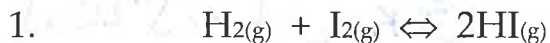


IX) Type 3 K_{eq} Problems

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 Knowns: $[H_2]_i = 0.200M$, $[I_2]_i = 0.200M$
What is the equilibrium $[HI]$?

R:	$H_2(g)$	+	$I_2(g)$	\rightleftharpoons	$2HI(g)$
I	0.200M		0.200M		0M
C	-x		-x		+2x
E	0.200-x		0.200-x		2x

② Let $x = \Delta[H_2]$

$\therefore H_2 @ \text{equil} = 0.200M - x$

Example:



① Knowns: $[CO]_i = 0.10M$, $[H_2O]_i = 0.10M$
What are the equilibrium concentrations of all species?

R	$CO(g)$	+	$H_2O(g)$	\rightleftharpoons	$CO_2(g)$	+	$H_2(g)$
I	0.10M		0.10M		0M		0M
C	-x		-x		+x		+x
E	(0.10-x)		(0.10-x)		x		x

② Let $x = \Delta[CO]_{eq}$

④ $0.10 - 0.06683 = 0.03317$

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

③ $K_{eq} = 55.6 = \frac{[HI]^2}{[H_2][I_2]} = \frac{(2x)^2}{(0.200-x)^2} = 55.6$

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

DON'T round early!

$7.45654(0.200-x) = 2x$

$1.491308 - 7.45654x = 2x$

$1.491308 = 9.45654x$

$\therefore 0.15770123 = x$ *s.f.*

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

$\therefore 0.315M$

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

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

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

$2.015(0.10-x) = x$

$0.2015 - 2.015x = x$

$0.2015 = 3.015x$

$\therefore x = 0.06683$

Example:

3. A certain amount of H_2O was placed in a 2.00L closed flask. When equilibrium was reached, the $[\text{H}_2]$ was 0.500M. If $K_{\text{eq}} = 16.0$ at this temperature, how many moles of H_2O were originally placed in the flask?

3 s.f.

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

① known

R	$2\text{H}_2\text{O}(\text{g})$	\rightleftharpoons	$2\text{H}_2(\text{g})$	+	$\text{O}_2(\text{g})$
I	x		0M		0M
C	-0.500M		$+0.500\text{M}$		$+0.250\text{M}$
E	$(x - 0.500)$		0.500M		0.250M

③ $K_{\text{eq}} = \frac{[\text{H}_2]^2 [\text{O}_2]}{[\text{H}_2\text{O}]^2}$

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

$16(x - 0.500)^2 = (0.500)^2 (0.250)$

$\sqrt{(x - 0.500)^2} = \sqrt{\frac{0.625}{16}}$

$(x - 0.500) = \frac{0.625}{4}$

$x = 0.5625\text{M}$

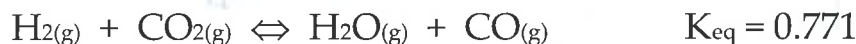
② Let $x = [\text{H}_2\text{O}]_i$

④ $[\text{H}_2\text{O}]_i = 0.5625\text{M} = \frac{n}{2.00\text{L}}$

moles of $\text{H}_2\text{O}_i = (0.5625\text{M})(2.00\text{L}) = 1.13\text{ moles}$

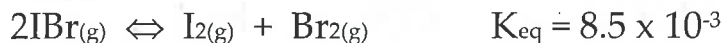
Assignment 8: Type 3 Exercises

1. For the following reaction:



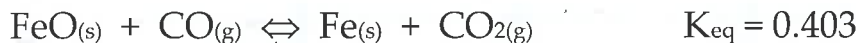
If 0.0100mol of H_2 and 0.0100mol of CO_2 are mixed in a 1.00L container, what are the concentrations of all substances at equilibrium?

2. For the following reaction:



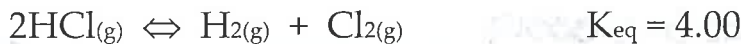
If 0.0600mol of IBr is placed in a 1.0L container, what are the concentrations of all three substances at equilibrium?

3. For the following reaction:



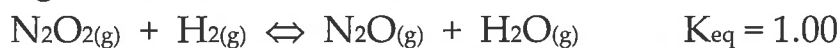
If 0.0500mol of CO and excess solid FeO are placed in a 1.00L container, what are the concentrations of CO and CO_2 when equilibrium has been attained?

4. For the following reaction:



An unknown [HCl] was added to a 2.00L flask and allowed to reach equilibrium. At equilibrium, [H₂] = 0.200M. How many moles of HCl was originally placed in the flask?

5. For the following reaction:



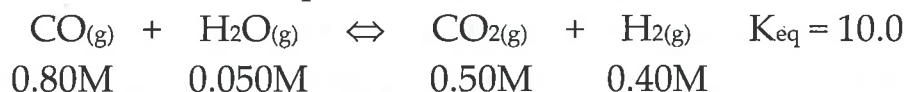
If 0.150mol each of N₂O and H₂O are introduced into a 1.00L flask and allowed to come to equilibrium, what concentration of N₂O₂ will be present at equilibrium?

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?



$$\text{Trial } K_{eq} = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]} = \frac{(0.50\text{M})(0.40\text{M})}{(0.80\text{M})(0.050\text{M})} = 5.0$$

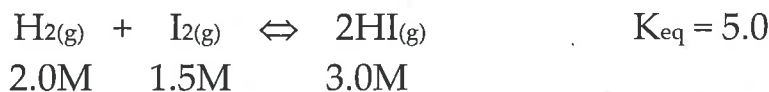
Trial K_{eq} of 5.0 < K_{eq} of 10.0

∴ 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?



$$\text{Trial } K_{eq} = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = \frac{(3.0\text{M})^2}{(2.0\text{M})(1.5\text{M})} = 3.0 \quad \text{Trial } K_{eq} \text{ of } 3.0 < K_{eq} \text{ } 5.0$$

∴ system must shift RIGHT to attain equilibrium

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

$$\text{Br}_2 = \frac{0.60 \text{ mol}}{2.0 \text{ L}} \\ [\text{Br}_2] = 0.30 \text{ M}$$

$$\text{Cl}_2 = \frac{0.80 \text{ mol}}{2.0 \text{ L}} \\ [\text{Cl}_2] = 0.40 \text{ M}$$

$$\text{BrCl} = \frac{2.20 \text{ mol}}{2.0 \text{ L}} \\ [\text{BrCl}] = 1.1 \text{ M}$$

Example:

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



Quantities of gases were found to be as follows:

$\text{Br}_2 = 0.60 \text{ mol}$, $\text{Cl}_2 = 0.80 \text{ mol}$, $\text{BrCl} = 2.20 \text{ mol}$

What will happen to the $[\text{Br}_2]$ as the system approaches equilibrium?

$$\text{Trial } K_{eq} = \frac{[\text{BrCl}]^2}{[\text{Br}_2][\text{Cl}_2]} = \frac{(1.1 \text{ M})^2}{(0.30)(0.40)} = 1.0 \times 10^1$$

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

\therefore system must shift LEFT to attain equilibrium
(to \downarrow Trial K_{eq} , $[\text{reactants}]$ must \uparrow and $[\text{products}] \downarrow$)

\therefore shift left will $\uparrow [\text{Br}_2]$

Assignment 9: Type IV Exercises

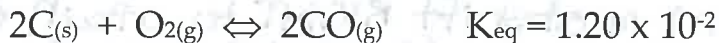
1. For the following reaction: $2\text{O}_3(g) \rightleftharpoons 3\text{O}_2(g) \quad K_{eq} = 75$

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: $2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g) \quad K_{eq} = 75$

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 $[\text{SO}_2]$ will decrease as equilibrium is established. Do you agree with the student's prediction? Explain using appropriate calculations.

3. Consider the following reaction:



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 $[\text{C}]$? Show all calculations.

4. For the following reaction: $\text{H}_2(g) + \text{Cl}_2(g) \rightleftharpoons 2\text{HCl}(g) \quad K_{eq} = 0.15$

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 ☺