Chemistry 12
Tutorial 6 - SOLUTIONS
Calculations Involving K_{eq}

1. The equilibrium equation for the formation of ***ammonia*** is:

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

In an **equilibrium mixture** at 200 °C, the concentrations were found to be as follows:

$[N_2]$ = 2.12M, $[H_2]$ = 1.75M and $[NH_3]$ = 84.3M

Notice the 3 SD's in all your data

Calculate the value of the Equilibrium Constant for this reaction at 200°C.

Solution: All concentrations are given at equilibrium, so we can write the K_{eq} expression and plug the values in to calculate K_{eq}.

$$K_{eq} = \frac{[NH_3]^2}{[N_2]\,[H_2]^3} = \frac{(84.3)^2}{(2.12)\,(1.75)^3} = 625 \text{ (or } 6.25 \times 10^2\text{)}$$

2. At 200°C, the Keq for the reaction: $N_{2(g)} + 3H_{2(g)} \rightleftarrows 2NH_{3(g)}$ is ***625***

If the $[H_2]$ = 0.430 M, and the $[NH_3]$ = 0.10 M, at equilibrium, calculate the equilibrium $[N_2]$.

Solution: Write out the K_{eq} expression and solve it for the unknown:

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

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

$$= \frac{(0.10)^2}{(625)\,(0.430)^3} = 2.0 \times 10^{-4} \text{ M}$$

3. Consider the following equilibrium system:

$$2NO_{(g)} + Cl_{2(g)} \rightleftarrows 2NOCl_{(g)}$$

0.80 moles of NO and 0.60 moles of Cl_2 are placed into a 1.0 L container and allowed to establish equilibrium. At equilibrium [NOCl] = 0.56 M.

a) Calculate the equilibrium [NO]
b) Calculate the equilibrium [Cl_2]
c) Determine the value of Keq at this temperature.

NOTE: In a 1.0 Litre container, concentration is moles/ 1.0 litre, so concentration is the same as the moles. In other words, if 0.80 moles of NO are placed in a 1.0 L container, the ***initial concentration of NO*** = 0.80 M

	$2NO_{(g)}$ +	$Cl_{2(g)}$ $\rightleftarrows$	$2NOCl_{(g)}$
[I]	**0.80**	**0.60**	**0**
[C]	**- 0.56**	**- 0.28**	**+ 0.56**
[E]	**0.24**	**0.32**	**0.56**

$$\mathbf{K_{eq} = \frac{[NOCl]^2}{[NO]^2\ [Cl]^2} = \frac{(0.56)^2}{(0.24)^2\ (0.32)} = 17}$$

a). The equilibrium [NO] = 0.24 M

b). The equilibrium [Cl_2] = 0.32 M

c). The value of K_{eq} = 17

4. Given the equilibrium equation:

$$A + 2B \rightleftharpoons C$$

When 2.0 moles of A and 4.0 moles of B are added to a 10.0 L container, an equilibrium established in which 1.4 moles of C are found.
Find the equilibrium concentrations of A, B and C.

Solution:

The initial concentration of A is $\mathbf{\dfrac{2.0\,mol}{10.0\ L} = 0.20\ M}$

The initial concentration of B is $\mathbf{\dfrac{4.0\,mol}{10.0\ L} = 0.40\ M}$

The equilibrium concentration of C is $\mathbf{\dfrac{1.4\,mol}{10.0\ L} = 0.14\ M}$

	A +	2B $\rightleftharpoons$	C
[I]	0.20	0.40	0
[C]	-0.14	-0.28	+ 0.14
[E]	0.06	0.12	0.14

Subtraction, so answer is to 2 dec. places.

The equilibrium [A] = 0.06 M
The equilibrium [B] = 0.12 M
The equilibrium [C] = 0.14 M

5. The equilibrium equation: $PCl_{5(g)} \rightleftarrows PCl_{3(g)} + Cl_{2(g)}$ has a Keq = 0.50 at 25°C.

If 0.60 moles of PCl_3, 0.45 moles of Cl_2 and 0.26 moles of PCl_5 are all placed in a 1.0 L container, will the reaction move to the left, right or not at all in order to reach equilibrium?

Solution: **Keq expression:** $K_{eq} = \dfrac{[PCl_3]\,[Cl_2]}{[PCl_5]}$

Plug in the values for initial concentration to calculate the Trial K_{eq} (Trial Quotient "Q")

$$\text{Trial } K_{eq} = \frac{[PCl_3]\,[Cl_2]}{[PCl_5]} = \frac{(0.60)\,(0.45)}{(0.26)} = 1.038 = 1.0 \text{ (2SD's)}$$

Trial K_{eq} (1.0) > K_{eq} (0.50) so the ratio of Products/Reactants is too high.

As equilibrium is approached, the reaction will move to the LEFT.

($[PCl_5]$ will increase and the $[PCl_3]$ and $[Cl_2]$ will decrease.)

6. Consider the reaction:

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

At a certain temperature the K_{eq} for this reaction = 1.50

If the initial concentration of ***all four species*** = 0.500 M, calculate the ***equilibrium concentration*** of CO_2 and CO.
Remember to determine which way the reaction has to shift in order to reach equilibrium using a Trial Keq (Sometimes this can be done in your head)

Solution:

Trial $K_{eq} = \frac{[H_2O]\,[CO]}{[H_2]\,[CO_2]} = \frac{(0.500)\;(0.500)}{(0.500)\;(0.500)} = 1.0$ **which is** $< K_{eq}$ **(1.5) so reaction will move RIGHT.**

Let "x" = the change (decrease) in the $[H_2]$ **as the reaction moves to the right.**

Make an ICE table:

	$H_{2(g)}$ +	$CO_{2(g)}$ $\rightleftarrows$	$H_2O_{(g)}$ +	$CO_{(g)}$
[I]	0.500	0.500	0.500	0.500
[C]	$-x$	$-x$	$+x$	$+x$
[E]	0.500 - x	0.500 - x	0.500 + x	0.500 + x
[E]				
[E]				

Plug [E]'s into the K_{eq} **Expression:**

$$K_{eq} = \frac{[H_2O]\,[CO]}{[H_2]\,[CO_2]} = \frac{(0.500 + x)\;(0.500 + x)}{(0.500 - x)\;(0.500 - x)}$$

$$K_{eq} = \frac{(0.500 + x)^2}{(0.500 - x)^2}$$

$$\sqrt{K_{eq}} = \sqrt{\frac{(0.500 + x)^2}{(0.500 - x)^2}}$$

$$\sqrt{1.50} = \frac{(0.500 + x)}{(0.500 - x)}$$

1.2247 (0.500 - x) = (0.500 + x) (*continued on the next page…*)

1.2247 (0.500 - x) = (0.500 + x)

0.61237 - 1.2247 x = 0.500 + x

0.61237 - 0.500 = 1.2247 x + x

0.11237 = 2.2247 x

$$x = \frac{0.11237}{2.2247} = 0.05051 \text{ M}$$

Put the value for "x" (0.05051 M) back into the table and solve for the concentrations:

	$H_{2(g)}$ +	$CO_{2(g)}$ ⇄	$H_2O_{(g)}$ +	$CO_{(g)}$
[I]	0.500	0.500	0.500	0.500
[C]	-*x*	-*x*	+*x*	+*x*
[E]	0.500 - x	0.500 - x	0.500 + x	0.500 + x
[E]	0.500 - 0.05051	0.500 - 0.05051	0.500 + 0.05051	0.500 + 0.05051
[E]	0.4495	0.4495	0.5505	0.5505

Answer **Equilibrium [CO] = 0.551 M**

Equilibrium [CO_2] = 0.450 M

Answers to Self-Test on Tutorial 6

1. Given the equilibrium: $3A_{(g)} + B_{(g)} \rightleftarrows 2C_{(g)}$

If 4.0 moles of A and 2.0 moles of B are added to a 2.0 L container, an equilibrium is established in which the [C] = 0.40 M.

a) Calculate the equilibrium [A] and [B]

Solution: Initial [A] = 4.0 mol/2.0L = 2.0 M
Initial [B] = 2.0 mol/2.0 L = 1.0 M

	3A(g) +	**B(g) ⇄**	**2C(g)**
[I]	**2.0**	**1.0**	**0**
[C]	**- 0.60**	**- 0.20**	**+ 0.40**
[E]	**1.4**	**0.8**	**0.40**

1 dp = 1SD due to subtracting from 1.0 (1 dp)

Equilibrium [A] = 1.4 M

Equilibrium [B] = 0.8 M

b) Calculate the value of Keq at the temperature at which this was carried out.

$$\mathbf{Keq} = \frac{[C]^2}{[A]^3\,[B]} = \frac{(0.40)^2}{(1.4)^3\,(0.8)} = \mathbf{0.073 = 0.07 \text{ (to 1 SD)}}$$

2. Given the equilibrium equation:

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

The value of Keq for this reaction at 25 °C is ***34.6***

0.200 moles of A, B, C & D are all added to a 1.0 L container.

Calculate the [B] at equilibrium.

*HINT: The reaction will have to shift to the **right** in order to reach equilibrium, **because Trial K_{eq} (1.0)< K_{eq} (34.6)***

	$A_{(g)}$ +	$B_{(g)}$ $\rightleftharpoons$	$C_{(g)}$ +	$D_{(g)}$
[I]	0.200	0.200	0.200	0.200
[C]	-*x*	-*x*	+*x*	+*x*
[E]	0.200 - x	0.200 - x	0.200 + x	0.200 + x
[E]				
[E]				

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

$$34.6 = \frac{(0.200 + x)^2}{(0.200 - x)^2}$$

$$\sqrt{34.6} = \sqrt{\frac{(0.200 + x)^2}{(0.200 - x)^2}}$$

$$5.882 = \frac{(0.200 + x)}{(0.200 - x)}$$

$5.882\,(0.200 - x) = 0.200 + x$

$1.1764 - 5.882\,x = 0.200 + x$ (add 5.882 x to both sides)

$1.1764 = 0.200 + 6.882\,x$ (subtract 0.200 from both sides)

6.882 x = 0.9764 (divide each side by 6.882)
6.882 x = 0.9764 (divide each side by 6.882)

$$x = \frac{0.9764}{6.882} = \mathbf{0.142}$$

Now we have x = 0.142, we put this into the last two rows of the ICE table:

	$A_{(g)}$ +	$B_{(g)}$ $\rightleftharpoons$	$C_{(g)}$ +	$D_{(g)}$
[I]	0.200	0.200	0.200	0.200
[C]	*-x*	*-x*	*+x*	*+x*
[E]	0.200 - x	0.200 - x	0.200 + x	0.200 + x
[E]	0.200 – 0.142	0.200 – 0.142	0.200 + 0.142	0.200 + 0.142
[E]	0.058	0.058	0.342	0.342

So Equilibrium [B] = 0.058 M

3. Consider the equation:

$$2X_{(g)} + 3Y_{(g)} \rightleftharpoons 2Z_{(g)}$$

An equilibrium mixture is analyzed and [X] is 0.030M, [Y] = 0.500M and [Z] = 0.600M

2 SD's

Calculate the value of Keq for this reaction.

Everything is at equilibrium already, so there is no ICE table. We just plug [E]'s into the K_{eq} expression:

$$K_{eq} = \frac{[Z]^2}{[X]^2\,[Y]^3}$$

$$= \frac{(0.600)^2}{(0.030)^2\,(0.500)^3} = \mathbf{3200 \ \ or \ \ 3.2 \times 10^2}$$

4. The Keq for the reaction: $A_{(g)} + B_{(g)} \rightleftarrows 2C_{(g)}$ is ***1.20***

A mixture of A, B and C is analyzed and found to contain 3.0M A, 0.40M B and 2.50M C.

This reaction will shift which way (left, right or not at all) in order to reach equilibrium?

We start by calculating a Trial K_{eq} :

$$\text{Trial } K_{eq} = \frac{[C]^2}{[A]\,[B]} = \frac{(2.50)^2}{(3.0)\,(0.40)} = \mathbf{5.21}$$

Since Trial K_{eq} (5.21) > K_{eq} (1.20), the reaction will move to the LEFT in order to reach equilibrium. As it moves to the left, [A] and [B] will increase and [C] will decrease.

This is the end of Tutorial 6 Solutions