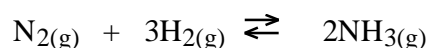


Chemistry 12
Tutorial 6 - SOLUTIONS
Calculations Involving K_{eq}

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



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

$$[\text{N}_2] = 2.12\text{M}, \quad [\text{H}_2] = 1.75\text{M} \quad \text{and} \quad [\text{NH}_3] = 84.3\text{M}$$

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} .

$$K_{eq} = \frac{[\text{NH}_3]^2}{[\text{N}_2] [\text{H}_2]^3} = \frac{(84.3)^2}{(2.12) (1.75)^3} = 625 \quad (\text{or } 6.25 \times 10^2)$$

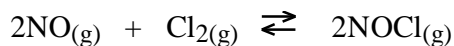
2. At 200°C, the K_{eq} for the reaction: $\text{N}_{2(g)} + 3\text{H}_{2(g)} \rightleftharpoons 2\text{NH}_{3(g)}$ is 625

If the $[\text{H}_2] = 0.430 \text{ M}$, and the $[\text{NH}_3] = 0.10 \text{ M}$, at equilibrium, calculate the equilibrium $[\text{N}_2]$.

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

$$\begin{aligned} K_{eq} &= \frac{[\text{NH}_3]^2}{[\text{N}_2] [\text{H}_2]^3} \\ [\text{N}_2] &= \frac{[\text{NH}_3]^2}{K_{eq} \times [\text{H}_2]^3} \\ &= \frac{(0.10)^2}{(625) (0.430)^3} = 2.0 \times 10^{-4} \text{ M} \end{aligned}$$

3. Consider the following equilibrium system:



0.80 moles of NO and 0.60 moles of Cl₂ are placed into a 1.0 L container and allowed to establish equilibrium. At equilibrium [NOCl] = 0.56 M.

- Calculate the equilibrium [NO]
- Calculate the equilibrium [Cl₂]
- Determine the value of K_{eq} 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)}	\rightleftharpoons	2NOCl _(g)
[I]	0.80		0.60		0
[C]	- 0.56		- 0.28		+ 0.56
[E]	0.24		0.32		0.56

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

- The equilibrium [NO] = 0.24 M
- The equilibrium [Cl₂] = 0.32 M
- The value of K_{eq} = 17

4. Given the equilibrium equation:



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 $\frac{2.0\text{mol}}{10.0\text{ L}} = 0.20\text{ M}$

The initial concentration of B is $\frac{4.0\text{mol}}{10.0\text{ L}} = 0.40\text{ M}$

The equilibrium concentration of C is $\frac{1.4\text{mol}}{10.0\text{ L}} = 0.14\text{ M}$

	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: $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$ has a $K_{\text{eq}} = 0.50$ 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:** $K_{\text{eq}} = \frac{[\text{PCl}_3] [\text{Cl}_2]}{[\text{PCl}_5]}$

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

$$\text{Trial } K_{\text{eq}} = \frac{[\text{PCl}_3] [\text{Cl}_2]}{[\text{PCl}_5]} = \frac{(0.60) (0.45)}{(0.26)} = 1.038 = 1.0 \text{ (2SD's)}$$

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.

($[\text{PCl}_5]$ will increase and the $[\text{PCl}_3]$ and $[\text{Cl}_2]$ will decrease.)

6. Consider the reaction:



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 CO₂ and CO.

Remember to determine which way the reaction has to shift in order to reach equilibrium using a Trial K_{eq} (Sometimes this can be done in your head)

Solution:

$$\text{Trial } K_{\text{eq}} = \frac{[\text{H}_2\text{O}][\text{CO}]}{[\text{H}_2][\text{CO}_2]} = \frac{(0.500)(0.500)}{(0.500)(0.500)} = 1.0 \text{ which is } < K_{\text{eq}} (1.5) \text{ so reaction will move RIGHT.}$$

Let “x” = the change (decrease) in the [H₂] as the reaction moves to the right.

Make an ICE table:

	H ₂ (g)	+	CO ₂ (g)	⇌	H ₂ O(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:

$$K_{\text{eq}} = \frac{[\text{H}_2\text{O}][\text{CO}]}{[\text{H}_2][\text{CO}_2]} = \frac{(0.500 + x)(0.500 + x)}{(0.500 - x)(0.500 - x)}$$

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

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

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

$$1.2247 (0.500 - x) = (0.500 + x) \text{ (continued on the next page...)}$$

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

$$0.61237 - 1.2247 x = 0.500 + x$$

$$0.61237 - 0.500 = 1.2247 x + x$$

$$0.11237 = 2.2247 x$$

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

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

	H ₂ (g)	+	CO ₂ (g)	⇌	H ₂ O(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 [CO₂] = 0.450 M

Answers to Self-Test on Tutorial 6

1. Given the equilibrium: $3A(g) + B(g) \rightleftharpoons 2C(g)$

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

Equilibrium [A] = 1.4 M

Equilibrium [B] = 0.8 M

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

b) Calculate the value of K_{eq} at the temperature at which this was carried out.

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

2. Given the equilibrium equation:



The value of K_{eq} 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]							

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

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

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

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

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

$$1.1764 - 5.882x = 0.200 + x \text{ (add 5.882 x to both sides)}$$

$$1.1764 = 0.200 + 6.882x \text{ (subtract 0.200 from both sides)}$$

$$6.882x = 0.9764 \text{ (divide each side by 6.882)}$$

$$6.882x = 0.9764 \text{ (divide each side by 6.882)}$$

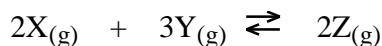
$$x = \frac{0.9764}{6.882} = \mathbf{0.142}$$

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

	A _(g)	+	B _(g)	⇌	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:



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

2 SD's

Calculate the value of K_{eq} for this reaction.

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

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

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

4. The K_{eq} for the reaction: $A_{(g)} + B_{(g)} \rightleftharpoons 2C_{(g)}$ 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