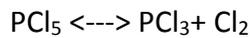


## Answer on the Question #55539 - Chemistry - Other

### Question:

1. Some  $\text{PCl}_5$  is pumped into a 500 mL flask. The  $\text{PCl}_3$  at equilibrium is 1.50 M. What was the initial  $[\text{PCl}_5]$ ?



$$K_{eq} = 2.14$$

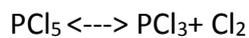
2.  $K_{eq}$  for the reaction  $2\text{HI} \rightleftharpoons \text{H}_2 + \text{I}_2$  has a value of  $1.85 \times 10^{-2}$  at 425 degrees celsius. If 0.18 mol of HI is placed in a 2.0 L flask and allowed to come to equilibrium at this temperature. What will the equilibrium of  $[\text{I}_2]$  be?

3. 0.020 mol of each  $\text{SO}_2$  and  $\text{O}_2$  and  $\text{SO}_3$  is placed in a 1.0 L flask and allowed to come to equilibrium. The equilibrium of  $[\text{SO}_2]$  is found to be 0.0080M. What is the value of  $K_{eq}$  for the reaction:  $2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$

4. A 3.00 L flask contains 6.00M  $\text{H}_2$ , 6.00M  $\text{Cl}_2$ , 3.00M  $\text{HCl}$  at equilibrium. An additional 15 mol of  $\text{HCl}$  is injected into the flask. What is the  $[\text{Cl}_2]$  when equilibrium is re-established?.

### Answer:

1. Some  $\text{PCl}_5$  is pumped into a 500 mL flask. The  $\text{PCl}_3$  at equilibrium is 1.50 M. What was the initial  $[\text{PCl}_5]$ ?



$$K_{eq} = 2.14$$

Let's consider the reaction equation. One can note, that the number of the moles of  $\text{PCl}_5$  consumed in the reaction is equal to the number of the moles of  $\text{PCl}_3$  and  $\text{Cl}_2$  produced (simple stoichiometry). Then if we will look at the expression for the equilibrium constant, we can substitute the values of  $[\text{PCl}_5]$  and  $[\text{Cl}_2]$  by  $(c_0 - [\text{PCl}_3])$  and  $[\text{PCl}_3]$ , respectively. Using simple algebra, we derive the initial concentration of phosphorus pentachloride  $c_0$ :



$$K_{eq} = \frac{[\text{PCl}_3] \cdot [\text{Cl}_2]}{[\text{PCl}_5]} = \frac{[\text{PCl}_3]^2}{(c_0 - [\text{PCl}_3])} = 2.14$$

$$c_0 = \frac{[\text{PCl}_3]^2 + 2.14 * [\text{PCl}_3]}{2.14} = 2.6 \text{ M}$$

2.  $K_{eq}$  for the reaction  $2HI \rightleftharpoons H_2 + I_2$  has a value of  $1.85 \times 10^{-2}$  at 425 degrees Celsius. If 0.18 mol of HI is placed in a 2.0 L flask and allowed to come to equilibrium at this temperature. What will the equilibrium concentration of  $I_2$  be?

The equilibrium constant expression is:

$$K_{eq} = \frac{[H_2] \cdot [I_2]}{[HI]^2} = 1.85 \cdot 10^{-2}$$

The equilibrium concentration of hydrogen is equal to the concentration of iodine  $[H_2] = [I_2]$ . Also, the equilibrium concentration of hydrogen iodide can be expressed as the difference between the initial concentration and the amount of free iodine produced:  $[HI] = c_0 - 2[I_2]$ .

$$K_{eq} = \frac{[I_2]^2}{(c_0 - 2[I_2])^2}$$

Initial concentration of hydrogen iodide is the ratio of number of the moles to volume of the system:

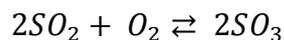
$$c_0 = \frac{n}{V} = \frac{0.18}{2.0} = 0.09 \text{ mol L}^{-1}$$

Using this value and solving the square equation, equilibrium iodine concentration is calculated:

$$[I_2] = 0.0096 \text{ mol L}^{-1}$$

3. 0.020 mol of each  $SO_2$  and  $O_2$  and  $SO_3$  is placed in a 1.0 L flask and allowed to come to equilibrium. The equilibrium concentration of  $[SO_2]$  is found to be 0.0080 M. What is the value of  $K_{eq}$  for the reaction:  $2SO_2 + O_2 \rightleftharpoons 2SO_3$

Let's write the reaction equation and understand what is going on in the system.



$c_0$	0.02	0.02	0.02
$\Delta$	$-2x$	$-x$	$+2x$
$[c]$	$(0.02 - 2x)$	$(0.02 - x)$	$(0.02 + 2x)$

Then, the x is:

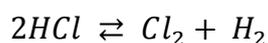
$$x = \frac{0.02 - 0.008}{2} = 0.006$$

And equilibrium constant is:

$$K_{eq} = \frac{(0.02 - x) \cdot (0.02 + 2x)^2}{0.008^2} = \frac{0.014 \cdot 0.032^2}{0.008^2} = 0.224$$

4. A 3.00 L flask contains 6.00M H<sub>2</sub>, 6.00M Cl<sub>2</sub>, 3.00M HCl at equilibrium. An additional 15 mol of HCl is injected into the flask. What is the [Cl<sub>2</sub>] when equilibrium is re-established?

The reaction equation is:



The equilibrium constant is:

$$K_{eq} = \frac{[Cl_2][H_2]}{[HCl]^2} = \frac{6 \cdot 6}{3^2} = 4$$

When HCl is added, the concentration of Cl<sub>2</sub> will increase by x, also H<sub>2</sub> concentration will increase by x, and the HCl concentration will decrease by 2x:

$$K_{eq} = \frac{(6 + x)(6 + x)}{(15 + 3 - 2x)} = 4$$

$$x = 1.662$$

Then, new equilibrium concentration of Cl<sub>2</sub> will be:

$$[Cl_2] = 6 + x = 7.66 \text{ mol L}^{-1}$$