

# Day 2: Changing to Equilibrium

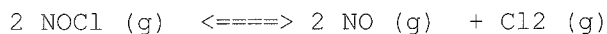
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PS#8

Period 20

- Return and go over quiz on Le Chatelier's Principle
- Go over Keq worksheet #1 → ICE Tables.
- More Mathematical Examples - Before and after situations

①

Example #4 :



Initially 2.00 mol of NOCl was placed in a 1.00 L flask at 462°C and left to reach the equilibrium shown above. When the equilibrium is established, it was found that 0.66 mol of NO were present. Calculate Keq.

Answer #4 :

We can start a table with the initial values given in the question:

Concentrations	$2\text{NOCl} \rightleftharpoons$	$2\text{NO}$	+	$\text{Cl}_2$
Initial	2.00	0.0		0.0
Change in	-	+		+
Equilibrium				

- NOTES :
1. Because the equilibrium is shifting right, the products will increase and the reactants will decrease, hence the signs - and +.
  2. Note the coefficients of the equation are translated into factors in the change in concentration. For example, if we are to make 0.5 mol of  $\text{Cl}_2$  we MUST also make 2 times as much NO (from equation).

Now let's see what other information the question gives us :

It states that we have 0.66 mol of NO at equilibrium. Therefore the  $[\text{NO}]_{\text{eq}}$  must be 0.66 M.

Since we produced 0.66 mol NO, we must have produced 0.33 mol of  $\text{Cl}_2$  according to the stoichiometry (for every 2 NO molecules produced, 1  $\text{Cl}_2$  is produced). Since we started with 0 mol  $\text{Cl}_2$ , the  $[\text{Cl}_2]_{\text{eq}}$  is 0.33 M

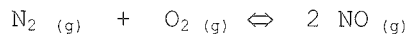
Furthermore, stoichiometrically if we produce 2 NO's we must have used 2 NOCl's. Similarly, if we produced 0.66 mol NO we must have used up 0.66 mol NOCl. Therefore  $[\text{NOCl}] = 2.00 - 0.66 = 1.34 \text{ M}$

Completing this table we get

Concentrations	$2 \text{NOCl}$	$\rightleftharpoons$	$2\text{NO}$	+	$\text{Cl}_2$
Initial	2.00		0.0		0.0
Change in	-0.66		+0.66		+0.33
Equilibrium	1.34		0.66		0.33

②

Example #6 :



2.00 mol  $\text{N}_2$  and 2.00 mol  $\text{O}_2$  are placed in a 2.00 L flask and allowed to come to equilibrium. What is the equilibrium  $[\text{NO}]$  ?  $K_{\text{eq}} = 4.51 \times 10^{-3}$

Answer #6 :

	$\text{N}_2$	$\text{O}_2$	$\text{NO}$
I	1.00	1.00	0.00
C	-X	-X	+2X
E	$1.00-X$	$1.00-X$	$2X$

$$K_{\text{eq}} = \frac{[\text{NO}]^2}{[\text{N}_2][\text{O}_2]} = 4.51 \times 10^{-3} = \frac{(2X)^2}{(1.00-X)(1.00-X)}$$

square root both sides

we get

$$6.72 \times 10^{-2} = \frac{2X}{(1.00-2X)}$$

Algebra :

$$\begin{aligned} (1.00-2X)(0.0672) &= 2X \\ 0.0672 - 0.0672X &= 2X \\ 0.0672 &= 2.0672X \\ 0.0325 &= X \end{aligned}$$

Is this what we wanted ?

Go back and look at the table and the question.

We wanted  $[\text{NO}]$ . According to our table,  $[\text{NO}] = 2X$   
 therefore  $[\text{NO}] = 2(0.0325) = 0.0650 \text{ M}$

Assignment :

- Read section 19-9 - LOOK at examples
- Keq Worksheet #2 - 1-4