Unit 13: Chemical Equilibrium

ICE Tables

In this worksheet, you will continue to develop your understanding of ICE Tables by familiarizing yourself with trickier problems involving **quadratic equations**. However, as you progress through the worksheet, you will also learn to spot shortcuts to make these problems easier, including the **5% Rule** and taking the **Square Root**.

1) 4.00 moles of chlorine and 3.00 moles of bromine are placed in a 5.00 L flask. Calculate equilibrium concentrations of all species. $K_C = 4.7 \times 10^{-3}$

$$Cl_{2\,(g)} + Br_{2\,(g)} \rightleftarrows 2\,BrCl_{(g)}$$

2) For the reaction below, what are the equilibrium concentrations when the reaction starts with 0.20 M nitrogen and 0.15 M oxygen? $K_C = 4.10 \times 10^{-4}$

$$N_{2\,(g)} + O_{2\,(g)} \rightleftarrows 2\,NO_{(g)}$$

3) For the reaction below at 25°C, K_C = 115. If 0.050 moles of each reactant and 0.600 moles of product are placed in a 1.00 L flask, what are the equilibrium concentrations?

$$H_{2\,(g)} + F_{2\,(g)} \rightleftarrows 2\,HF_{(g)}$$

4) In a study of halogen bond strengths, 0.50 mol of iodine gas was heated in a 2.5 L vessel, and the following reaction occurred:

$$I_{2\,(g)} \rightleftarrows 2\,I_{(g)}$$

a) Calculate [I_2] and [I] at equilibrium at 600 K; K_C = 2.94×10^{-10}

b) Calculate [I_2] and [I] at equilibrium at 2000 K; K_C = 0.209

Unit 13: Chemical Equilibrium

ANSWER KEY:

1) 4.00 moles of chlorine and 3.00 moles of bromine are placed in a 5.00 L flask. Calculate equilibrium concentrations of all species. $K_C = 4.7 \times 10^{-3}$

$$Cl_{2\,(g)} + Br_{2\,(g)} \rightleftarrows 2\,BrCl_{(g)}$$

1. Remember that when filling ICE tables, molarity is used for the values. So make sure to divide your moles by the volume of the flask!

2. As for the Change and Equilibrium values, a variable must be used.

	$Cl_{2\,(g)}$	$Br_{2\,(g)}$	$BrCl_{(g)}$
Initial	(4.00 mol /5.00 L) or 0.800 M	(3.00 mol/5.00 L) or 0.600 M	0 M
Change	-x M	-x M	+2x M
Equilibrium	0.800-x M	0.600-x M	2x M

3. For this problem, we can use something called the 5% Rule, that allows us to "ignore" the "-x" term in the equilibrium values of the reactants, decreasing the amount of algebra needed. We can only use this when the K constant is multiplied to 10^{-3} or smaller.

Using the approximation:

$$\frac{[BrCl]^2}{[Br_2][Cl_2]} = 4.7 \times 10^{-3}$$

$$\frac{(2x)^2}{(0.600)(0.800)} = 4.7 \times 10^{-3}$$

$$4x^2 = 4.7 \times 10^{-3}(0.600)(0.800)$$

$$4x^2 = 0.00226$$

$$x^2 = 5.64 \times 10^{-4}$$

$$x = 0.024\ M$$

4. However, it is important to note that this rule assumes that the "x" term is small enough compared to the initial values to be considered negligible. So we must make sure that the "x" value is less than 5% of the initial concentrations.

$$\frac{0.024\ M}{0.600\ M} \times 100\% = 4.0\%$$

Also note how you can calculate this percentage for the smaller concentration as if the "x" term is less than 5% of the smaller one, it will also be less than 5% of the larger one.

5. Since it obeys the 5% Rule, the answers are:

   $[Cl_2]$ = 0.800 M - 0.024 M = **0.776 M**
   $[Br_2]$ = 0.600 M - 0.034 M = **0.576 M**
   [BrCl] = 2(0.024 M) = **0.048 M**

2) For the reaction below, what are the equilibrium concentrations when the reaction is started with 0.20 M nitrogen and 0.15 M oxygen? K_C = 4.10×10^{-4}

$$N_{2\,(g)} + O_{2\,(g)} \rightleftarrows 2\,NO_{(g)}$$

1. The setup is exactly the same as the last problem

	N_2	O_2	NO
Initial	0.20 M	0.15 M	0 M
Change	-x M	-x M	+2x M
Equilibrium	0.20-x M	0.15-x M	2x M

2. Once again, we can use the approximation method as K_c is small enough:

$$\frac{[NO]^2}{[N_2][O_2]} = 4.10 \times 10^{-4}$$

$$\frac{(2x)^2}{(0.20)(0.15)} = 4.10 \times 10^{-4}$$

$$4x^2 = 4.10 \times 10^{-4}(0.20)(0.15)$$

$$4x^2 = 1.23 \times 10^{-5}$$

$$x^2 = 3.08 \times 10^{-6}$$

$$x = 1.8 \times 10^{-3}\ M$$

3. Verify that it obeys the 5% Rule:

$$\frac{1.8\times10^{-3}\,M}{0.15\,M}\times 100\% = 1.2\%$$

4. Since it obeys the 5% Rule, the answers are:

$[N_2] = 0.20M - 1.8\times10^{-3}M = 0.1982\,M = 0.20M$
$[O_2] = 0.15M - 1.8\times10^{-3}M = 0.1482\,M = 0.15M$
$[NO] = 1.8\times10^{-3} * 2 = 3.5\times10^{-3}M$

3) For the reaction below at 25°C, K_C = 115. If 0.050 moles of each reactant and 0.600 moles of product are placed in a 1.00 L flask, what are the equilibrium concentrations?

$$H_{2\,(g)} + F_{2\,(g)} \rightleftarrows 2\,HF_{(g)}$$

Remember that when filling ICE tables, molarity is used for the values. So make sure to divide your moles by the volume of the flask! In this case the moles will be the same as molarity given the volume is 1.00 L.

Since we are given the amount of moles of both the products and reactants, we have to determine in what direction the reaction will go to reach equilibrium. We must use the Equilibrium Quotient:

1. $Q = \frac{(0.600)^2}{(0.050)^2} = 144;\ K < Q,\ meaning\ there\ is\ more\ product\ than\ there\ should\ be\ at\ equilibrium\ so\ more\ reactant\ will\ be\ made$

This means that in our ICE table, the signs for change on the reactants side will be positive and for the product will be negative.

	H_2	F_2	HF
Initial	0.050 M	0.050 M	0.600 M
Change	+x M	+x M	-2x M
Equilibrium	0.050+x M	0.050+x M	0.600-2x M

2. We can calculate X as we have done in other problems using K_c

$$\frac{[HF]^2}{[H_2][F_2]} = 115$$

$$\frac{(0.600-2x)^2}{(0.050+x)^2} = 115;\ \textit{We can take the square root of both sides to solve for x easier}$$

$$\frac{0.600-2x}{0.050+x} = 10.72$$

$$0.600 - 2x = 10.72(0.050 + x)$$

$$0.600 - 2x = 0.536 + 10.72x$$

$$0.0638 = 12.72x$$

$$x = 0.0050\ M$$

3. Therefore, at equilibrium we have
   $[H_2]$ = 0.050 M + 0.0050 M= 0.055 M
   $[F_2]$ = 0.050 M + 0.0050 M= 0.055 M
   [HF] = 0.600 M - 2(0.0050 M) = 0.590 M

4) In a study of halogen bond strengths, 0.50 mol of iodine gas was heated in a 2.5 L vessel, and the following reaction occurred:

$$I_{2\,(g)} \rightleftarrows 2\,I_{(g)}$$

a) Calculate $[I_2]$ and [I] at equilibrium at 600 K; $K_C = 2.94 \times 10^{-10}$

First we fill in the ICE table. Remember to divide the number of moles by volume to get the molarity. In this case,

$$\frac{0.50\ mol}{2.5\ L} = 0.20\ M$$

	$I_{2\,(g)}$	$2\,I_{(g)}$
Initial	0.20 M	0 M
Change	-x M	+2x M
Equilibrium	0.20 - x M	2x M

1. After setting up the ICE table, let's use the approximation method:

$\frac{[I_2]^2}{[I_2]} = 2.94 \times 10^{-10}$

$\frac{(2x)^2}{0.20-x} = \frac{(2x)^2}{0.20} = 2.94 \times 10^{-10}$

$4x^2 = 5.88 \times 10^{-11}$

$x^2 = 1.47 \times 10^{-11}$

$x = 3.8 \times 10^{-6} M$

2. Does it obey the 5% Rule? $\frac{3.8 \times 10^{-6} M}{0.20 M} \times 100\% = 0.0019\%$

3. Since it does, the equilibrium concentrations are:

$[I_2] = 0.20 M - 3.8 \times 10^{-6} = 0.199 M = 0.20 M$

$[I] = 2 * 3.8 \times 10^{-6} M = 7.7 \times 10^{-6} M$

b) Calculate $[I_2]$ and [I] at equilibrium at 2000 K; $K_C = 0.209$

1. Since the K value is greater than 10^{-3}, we cannot use the approximation method here and must resort to the quadratic formula…

$\frac{[I_2]^2}{[I_2]} = 0.209$

$\frac{(2x)^2}{0.20-x} = 0.209$

$4x^2 = 0.209(0.20 - x)$

$4x^2 = 0.0418 - 0.209x$

$4x^2 + 0.209x - 0.0418 = 0$

Recall: the quadratic formula is:

$x = \frac{-b \pm \sqrt{b^2-4ac}}{2a}$

Where an equation is in the form:

$ax^2 + bx + c = 0$

Using our equation we can plug these values into the quadratic formula

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a} = \frac{-0.209 \pm \sqrt{(0.209)^2 - 4(4)(-0.0418)}}{2(4)}$$

Plugging directly into our calculator, we get

$x = 0.079$ $\qquad$ $x = -0.132$

But as concentration cannot be negative, our value of x is 0.079 M.

2. As a result,
   $[I_2] = 0.20\,M - 0.079\,M = 0.121\,M = 0.12\,M$
   $[I] = 0.079\,M * 2 = 0.158\,M = 0.16\,M$