SCH 4U

EQUILIBRIUM CONSTANT

Consider the reaction: $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

TRIAL	$[HI]_{(aq)}$	$[H_2]_{(aq)}$	$[I_2]_{(aq)}$	$\frac{[HI]^2_{(aq)}}{[H_2]_{(aq)}[I_2]_{(aq)}}$
1	0.156	0.0220	0.0220	50.3
2	1.00	0.820	0.0243	50.2
3	0.750	0.106	0.106	50.1

∴ For $aA + bB \rightleftharpoons cC + dD$ at equilibrium,

$$\boxed{K_{eq} = K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}}$$

The Equilibrium Expression

Molar concentration (mol/L)

Equilibrium constant

Eg. What is the equilibrium expression for $N_{2(g)} + 3H_{2(g)} \rightleftharpoons 2NH_{3(g)}$

NOTE:

1. K_c does not depend on [initial]'s, <u>only depends</u> on [equilibrium]'s.

2. K_c changes with temperature.

3. K_c does not include **[ ]** of solids and liquids... **ONLY** <u>gases and solutions</u>.

Eg. $Fe_3O_{4(s)} + H_{2(g)} \rightleftharpoons 3FeO_{(s)} + H_2O_{(g)}$ $\quad K_c =$

GIVEN [EQUILIBRIUM]'S, DETERMINE K_c

Eg. At 460°C, the reaction

$$2NO_{(g)} + O_{2(g)} \rightleftharpoons 2NO_{2(g)}$$

is at equilibrium. In a 3.00 L container, there are 0.300 mol of NO, 0.0420 mol of O_2, and 0.300 mol of NO_2. Determine K_c.

GIVEN [INITIAL]'S AND K_c, DETERMINE [EQUILIBRIUM]

TYPE A: PERFECT SQUARES

Eg. At 430°C, the equilibrium constant, K_c, for the following reaction is 1.84×10^{-2}.

$$2HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)}$$

If 0.0500 mol of HI is placed in a 500.0 mL container and allowed to reach equilibrium, determine the equilibrium concentrations of all species.

***STEP 1**: Set up an ICE table with equation.*

		$2HI_{(g)} \rightleftharpoons$	$H_{2(g)}$ +	$I_{2(g)}$
	Initial		0	0
	Change			
$I + C =$	Equilibrium			

Product of coefficients of substance and x,
+ve if forming, **-ve** if disappearing.

STEP 2: Substitute K_c and [eq] into equilibrium expression.

$$K_c = \frac{[H_2][I_2]}{[HI]^2}$$

$$1.84 \times 10^{-2} = \frac{(x)(x)}{(0.100-2x)^2}$$

$$1.84 \times 10^{-2} = \frac{x^2}{(0.100-2x)^2}$$ ⟵ perfect squares, therefore we can square root both sides.

$$0.136 = \frac{x}{0.100-2x}$$

$$0.0136 - 0.272x = x$$

$$\therefore\ x = 1.07 \times 10^{-2}$$

$\therefore$ $[H_2] = [I_2] = 1.07 \times 10^{-2}\ mol/L$

and $[HI] = 1.00 - 2(1.07 \times 10^{-2}) = 7.86 \times 10^{-2}\ mol/L$

Eg. For $SO_{2(g)} + NO_{2(g)} \rightleftharpoons NO_{(g)} + SO_{3(g)}$ $K_c = 85.0$ at 460°c

If 0.300 mol of $SO_{2(g)}$ and 0.300 mol of $NO_{2(g)}$ is placed in a 3.00 L container, what will the concentrations of NO_2 and NO be at equilibrium?

	$SO_{2(g)}$ +	$NO_{2(g)}$ $\rightleftharpoons$	$NO_{(g)}$ +	$SO_{3(g)}$
Initial				
Change				
Equilibrium				

TYPE B: QUADRATIC EQUATIONS

Eg. At 1100K, K_c = 25.0 for $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

4.00 mol of $H_{2(g)}$ and 6.00 mol of $I_{2(g)}$ are placed in a 2.00 L container at 1100K. What is the equilibrium concentration of each gas?

	$H_{2(g)}$ +	$I_{2(g)} \rightleftharpoons$	$2HI_{(g)}$
Initial			
Change			
Equilibrium			

Eg. If 1.00 mol of $N_2O_{4(g)}$ is placed in a 2.00 L closed container at 150°C, what will the concentrations of $N_2O_{4(g)}$ and $NO_{2(g)}$ at equilibrium?

$N_2O_{4(g)} \rightleftharpoons 2NO_{2(g)}$ $\qquad K_c = 4.50$ at 150°C

TYPE C: APPROXIMATION METHOD

$$K_c = \frac{[Products]^x}{[Reactants]^y}$$

1. If Reactants $\rightleftharpoons$ Products, then $[\boldsymbol{P}] \gg [\boldsymbol{R}]$ $\therefore K_c \gg 1$

2. If Reactants $\rightleftharpoons$ Products, then $[\boldsymbol{P}] \approx [\boldsymbol{R}]$ $\therefore K_c \approx 1$

3. If Reactants $\rightleftharpoons$ Products, then $[\boldsymbol{P}] \ll [\boldsymbol{R}]$ $\therefore K_c \ll 1$

In #3, very little reactant converts to product, so in the ICE table,
the $-x$ term in the CHANGE row is negligible and so [reactants]$_{\text{initial}}$ = [reactants]$_{\text{equilibrium}}$.

How do we know when to ignore the $-x$ term?
– this will make calculations much easier.

If $\boxed{\dfrac{smallest\ [initial]}{K_c} > 500}$ = ignore the x in any bracketed terms on the **E** line!!

Eg. At 727°C, $K_c = 3.80 \times 10^{-5}$ for $I_{2(g)} \rightleftharpoons 2I_{(g)}$.

If the original [I_2] is 0.200 mol/L, what is the equilibrium [I]?

	$I_{2(g)}$ $\rightleftharpoons$	$2I_{(g)}$
Initial	0.200 mol/L	0
Change	$-x$	$+2x$
Equilibrium	$(0.200 - x)$	$2x$

$$K_c = \frac{[I]^2}{\{I_2]}$$

$$3.80 \times 10^{-5} = \frac{(2x)^2}{(0.200 - x)}$$

APPROXIMATION METHOD???

$$3.80 \times 10^{-5} = \frac{4x^2}{0.200}$$

$$\frac{0.200}{3.8 \times 10^{-5}} = 5263 > 500$$

$$7.60 \times 10^{-6} = 4x^2$$

Therefore, approximation method works, so ignore $-x$ of reactants inside brackets!

$$1.90 \times 10^{-6} = x^2$$

$$\therefore \; x = 1.38 \times 10^{-3}$$

$\therefore \; [I] = 2(1.38 \times 10^{-3}) = 2.76 \times 10^{-3} \; mol/L$

$\therefore \; [I_2] = 0.200 - 2.76 \times 10^{-3} \; mol/L$

$= 0.199 \; mol/L$ (very little change)

Eg. At 2000°C, the following has an equilibrium constant, K_c, of 6.40×10^{-7}.

$$2CO_{2(g)} \rightleftharpoons 2CO_{(g)} + O_{2(g)}$$

If 5.00 mole of CO_2 is placed in a 5.00 L container, what will be the [CO] at equilibrium?