

SCH 4U

EQUILIBRIUM CONSTANT

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

TRIAL	$[HI]_{(aq)}$	$[H_2]_{(aq)}$	$[I_2]_{(aq)}$	$\frac{[HI]_{(aq)}^2}{[H_2]_{(aq)}[I_2]_{(aq)}}$
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

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

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

The Equilibrium Expression

↓

Molar concentration (mol/L)

Equilibrium constant

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

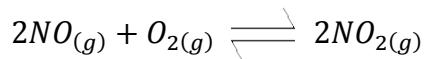
NOTE:

- K_c does not depend on [initial]'s, only depends on [equilibrium]'s.
- K_c changes with temperature.
- K_c does not include [] of solids and liquids... **ONLY gases and solutions.**

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

GIVEN [EQUILIBRIUM]'S, DETERMINE K_c

Eg. At 460°C, the reaction

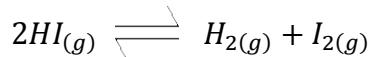


is at equilibrium. In a 3.00 L container, there are 0.300 mol of NO, 0.0420 mol of O₂, and 0.300 mol of NO₂. Determine K_c.

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} .



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			
$I + C =$			

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

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

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

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

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

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

$$0.0136 - 0.272x = x$$

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

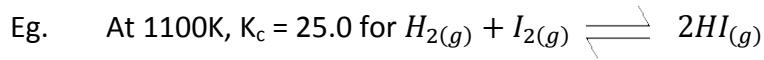
$$\therefore [H_2] = [I_2] = 1.07 \times 10^{-2} \text{ mol/L}$$

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



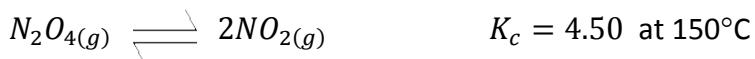
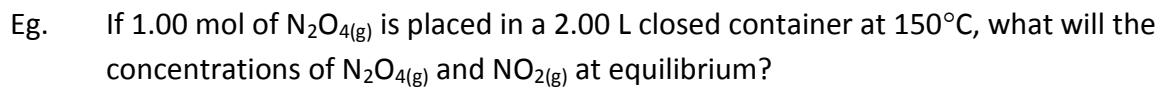
If 0.300 mol of $SO_{2(g)}$ and 0.300 mol of $NO_{2(g)}$ is placed in a 3.00 L container, what will the concentrations of NO_2 and NO be at equilibrium?

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

TYPE B: QUADRATIC EQUATIONS

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					



TYPE C: APPROXIMATION METHOD

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

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

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

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

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

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

If $\frac{\text{smallest [initial]}}{K_c} > 500$ = 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	$(0.200 - x)$	$2x$

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

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

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

APPROXIMATION METHOD???

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

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

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

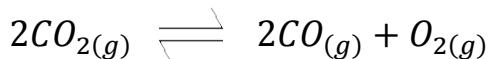
$$1.90 \times 10^{-6} = x^2$$

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

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

$$\begin{aligned} \therefore [I_2] &= 0.200 - 2.76 \times 10^{-3} \text{ mol/L} \\ &= 0.199 \text{ mol/L} \text{ (very little change)} \end{aligned}$$

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



If 5.00 mole of CO₂ is placed in a 5.00 L container, what will be the [CO] at equilibrium?