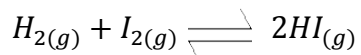


## SCH 4U

### EQUILIBRIUM CONSTANT

Consider the reaction:



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**

↓

Equilibrium constant

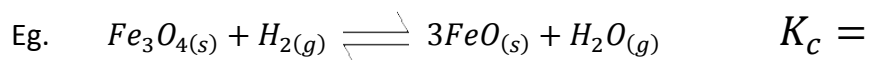
↘

Molar concentration (mol/L)

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

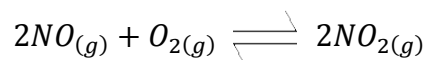
#### NOTE:

- $K_c$  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.



## GIVEN [EQUILIBRIUM]'S, DETERMINE $K_c$

Eg. At 460°C, the reaction

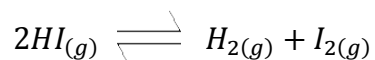


is at equilibrium. In a 3.00 L container, there are 0.300 mol of NO, 0.0420 mol of O<sub>2</sub>, and 0.300 mol of NO<sub>2</sub>. Determine  $K_c$ .

## 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  $1.84 \times 10^{-2}$ .



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				
Equilibrium				

$I + C =$

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

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

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

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

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

← perfect squares, therefore we can square root both sides.

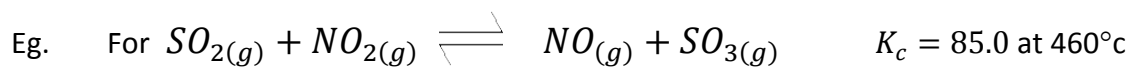
$$0.136 = \frac{x}{0.100-2x}$$

$$0.0136 - 0.272x = x$$

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

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

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



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**

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?



### TYPE C: APPROXIMATION METHOD

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

1. If Reactants  $\rightleftharpoons$  Products, then  $[P] \gg [R] \quad \therefore K_c \gg 1$
2. If Reactants  $\rightleftharpoons$  Products, then  $[P] \approx [R] \quad \therefore K_c \approx 1$
3. If Reactants  $\rightleftharpoons$  Products, then  $[P] \ll [R] \quad \therefore K_c \ll 1$

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

Eg. At  $727^\circ\text{C}$ ,  $K_c = 3.80 \times 10^{-5}$  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!

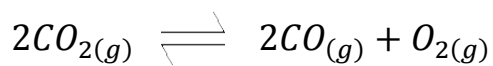
$$1.90 \times 10^{-6} = x^2$$

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

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

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

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



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