Chemistry 12

Equilibrium IV

II.8
(pg 63-72)

Name:
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1. ICE Tables

ICE Tables

Example:

(1) A 2.0 L bulb contains 6.00 mol of $NO_{2\,(g)}$, 3.0 mol of $NO_{(g)}$ and 0.20 mol of O_2 at equilibrium. What is the K_{eq} for this reaction?

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

- Determine the K_{eq} expression:

$$K_{eq} = \frac{[NO_2]^2}{[NO]^2[O_2]}$$

- Calculate the concentration of each species:

$$[NO_2] = \frac{6.00\text{ mol}}{2.0\text{ L}} = 3.0\text{ M}$$

$$[NO] = 1.5\text{ M}$$

$$[O_2] = 0.1\text{ M}$$

- Substitute into the K_{eq} expression:

$$K_{eq} = \frac{(3.0)^2}{(1.5)^2(0.1)} = \boxed{40}$$

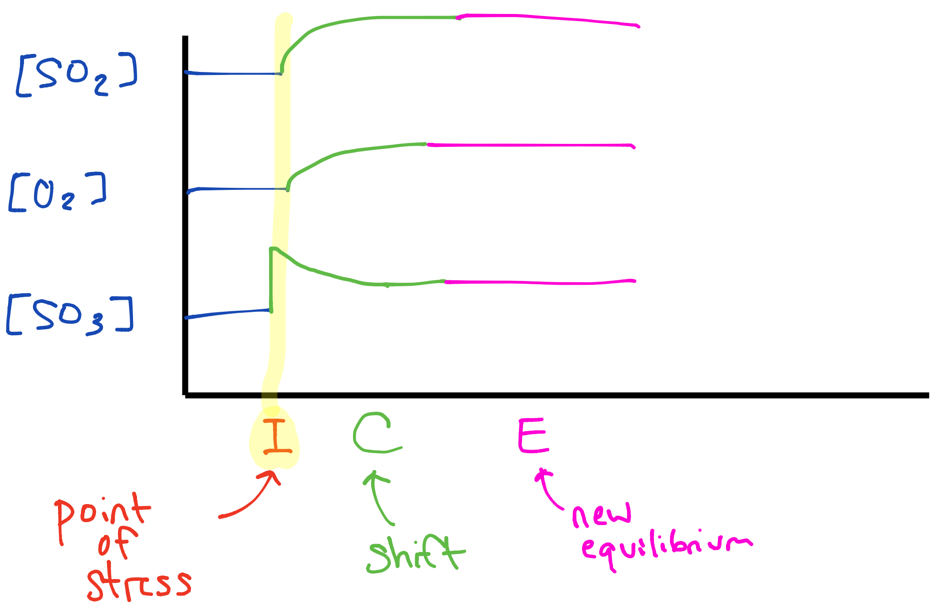
stress

(2) 4.00 mol of SO_3 is introduced into a 2.00 L bulb. After 5 minutes, equilibrium is established and is found that 0.500 mol of $O_{2\,(g)}$ exists. What is the K_{eq} value?

@ equil'm

$$2\,SO_{2\,(g)} + O_{2\,(g)} \leftrightharpoons 2\,SO_{3\,(g)}$$

- The system was at equilibrium – then a stress was introduced. What is this stress? SO_3 added
- The addition of product means that the equilibrium will shift __left__ forming more __reactants__ and decreasing concentration of the __product__.



- **All changes in concentration must obey stoichiometric proportions.**

	2 $SO_{2\,(g)}$	+	$O_{2\,(g)}$	$\leftrightharpoons$	2 $SO_{3\,(g)}$
Step 1 ***Original Equilibrium***	0 M		0 M		0 M
Step 1 **Initial** *(Where the stress is introduced)*	0 M		0 M		$\frac{4.00\text{ mol}}{2.0\text{ L}} = 2.0\text{ M}$
Step 3 **Change** *(How the system responds to the stress)*	+0.500 M		+0.250 M		−0.500 M
Step 2 (O_2) Step 4 (SO_2, SO_3) **Equilibrium** *(New equil'm concentrations)*	0.500 M		$\frac{0.500\text{ mol}}{2.0\text{ L}} = 0.250\text{ M}$		1.50 M

- K_{eq} is calculated using the new equilibrium concentrations:

$$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(1.50)^2}{(0.500)^2(0.250)}$$

$$= \boxed{36.0}$$

(3) A 1.00L reaction vessel contains 0.750 mol of CO and 0.275 mol of H_2O only. After 1 hour, equilibrium is reached according to:

$$\_\_\_\_ CO_{(g)} + \_\_\_\_ H_2O_{(g)} \leftrightharpoons \_\_\_\_ CO_{2\,(g)} + \_\_\_\_ H_{2\,(g)}.$$

Analysis shows 0.250 mol of CO_2 is present at equilibrium. What is the K_{eq}?

	$CO_{(g)}$	+	$H_2O_{(g)}$	$\leftrightharpoons$	$CO_{2\,(g)}$	+	$H_{2\,(g)}$
Original Equilibrium	0M		0M		0M		0M
Initial	0.750M		0.275M		0M		0M
Change	−0.250M		−0.250M		+0.250M		+0.250M
Equilibrium	0.500M		0.025M		0.250M		0.250M

$$K_{eq} = \frac{[\text{Products}]}{[\text{Reactants}]} = \frac{[CO_2][H_2]}{[CO][H_2O]}$$

$$= \frac{(0.250)(0.250)}{(0.500)(0.025)}$$

$$= \boxed{5.0}$$

(4) A student placed 7.00 mol NH_3 in a 0.500L flask. At equilibrium, 6.2 M N_2 was found in the flask. What is the equilibrium constant for this reaction?

$$\underline{2}\ NH_{3\,(g)} \leftrightharpoons \_\_\_\_ N_{2\,(g)} + \underline{3}\ H_{2\,(g)}$$

	NH_3	N_2	H_2
I	14.0M	0M	0M
C	−12.4M	+6.2M	+18.6M
E	1.6M	6.2M	18.6M

$$K_{eq} = \frac{[N_2][H_2]^3}{[NH_3]^2} = \frac{(6.2)(18.6)^3}{(1.6)^2}$$

$$= 15584$$

$$= \boxed{16000}$$

*2 s.f.

Determining Equilibrium Concentrations from K_{eq} and the Initial Concentrations

(1) The following gases are injected into a 1.00L flask: 1.20 mol of $H_{2\,(g)}$ and 1.20 mol of $F_{2\,(g)}$. What will the concentration of HF be when equilibrium is achieved?

$$\_\_\_\_ H_{2\,(g)} + \_\_\_\_ F_{2\,(g)} \leftrightharpoons \underline{2}\ HF_{(g)} \qquad K_{eq} = 2.50$$

	$H_{2\,(g)}$	+	$F_{2\,(g)}$	⇋	2 $HF_{(g)}$
Initial	1.20M		1.20M		0M
Change	$-x$		$-x$		$+2x$
Equilibrium	$1.20-x$		$1.20-x$		$2x$

$$K_{eq} = 2.50 = \frac{[HF]^2}{[H_2][F_2]}$$

$$\sqrt{2.50} = \frac{\sqrt{(2x)^2}}{\sqrt{(1.20-x)^2}}$$

$$1.58 = \frac{2x}{1.20-x}$$

$$1.897 - 1.58x = 2x$$

$$1.897 = 3.58x$$

$$x = 0.5298$$

$$[HF] = 2x$$

$$= \boxed{1.06M}$$

(2) 4.00 mol H_2 and 4.00 mol C_2N_2 are injected into a 2.00 L flask where they establish equilibrium. What is the $[C_2N_2]$ when equilibrium is achieved? $K_{eq} = 5.00$

$$\_\_\_\_ H_{2\,(g)} + \_\_\_\_ C_2N_{2\,(g)} \leftrightharpoons \underline{2}\ HCN_{(g)}$$

	H_2	C_2N_2	HCN
I	2.00M	2.00M	0M
C	$-x$	$-x$	$+2x$
E	$2.00-x$	$2.00-x$	$2x$

$$K_{eq} = 5.00 = \frac{[HCN]^2}{[H_2][C_2N_2]}$$

$$\sqrt{5.00} = \sqrt{\frac{(2x)^2}{(2.00-x)^2}}$$

$$2.236 = \frac{2x}{2.00-x}$$

$$4.47 - 2.236x = 2x$$

$$4.47 = 4.236x$$

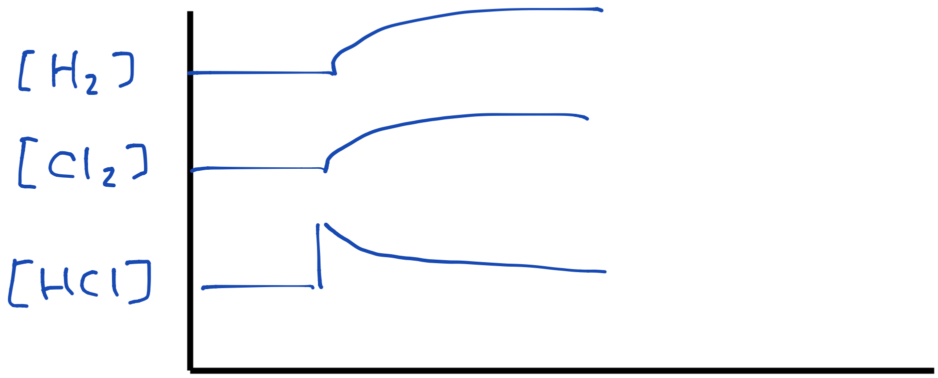
$$x = 1.0557$$

$$[C_2N_2] = 2.00 - x$$

$$= \boxed{0.944M}$$

(3) A 3.00 L flask contains 6.00 M H_2, 6.00 M Cl_2 and 3.00 M HCl at equilibrium. An additional 15 mol of HCl is injected into the flask. What is the $[Cl_2]$ when equilibrium is re-established?

What is K_{eq}? $K_{eq} = \frac{[HCl]^2}{[H_2][Cl_2]} = \frac{(3.00)^2}{(6.00)(6.00)} = \boxed{0.250}$



	$H_{2\,(g)}$	+	$Cl_{2\,(g)}$	⇋	2 $HCl_{(g)}$
Original Equilibrium	6.00M		6.00M		3.00M
Initial	6.00M		6.00M		$\frac{15 mol}{3.00 L}$ = 5.00M added (8.00M)
Change	$+x$		$+x$		$-2x$
Equilibrium	$6.00+x$		$6.00+x$		$8.00-2x$

$$K_{eq} = \sqrt{0.250} = \sqrt{\frac{(8.00-2x)^2}{(6.00+x)^2}}$$

$$0.500 = \frac{8.00-2x}{6.00+x}$$

$$3 + 0.5x = 8 - 2x$$

$$2.5x = 5$$

$$x = 2$$

$$[Cl_2] = 6.00 + x = \boxed{8.00M}$$

(4) The table below shows the molarity of three gases at equilibrium. The concentration of HF is then decreased as shown. What is the [HF] when equilibrium is re-established?

$$K_{eq} = \frac{[HF]^2}{[H_2][F_2]} = \frac{(12.0)^2}{(6.00)(6.00)} = 4.00$$

	$H_{2\,(g)}$	+	$F_{2\,(g)}$	⇋	2 $HF_{(g)}$
Original Equilibrium	6.00		6.00		12.0
Initial	6.00		6.00		5.00
Change	$-x$		$-x$		$+2x$
Equilibrium	$6.00-x$		$6.00-x$		$5.00+2x$

stress = removal of HF

$$K_{eq} = \sqrt{4.00} = \sqrt{\frac{(5.00+2x)^2}{(6.00-x)^2}}$$

$$2.00 = \frac{5.00+2x}{6.00-x}$$

$$x = 1.75$$

$$[HF] = 5.00 + 2(1.75) = \boxed{8.50\,M}$$

Hebden Workbook: Pg. 70 #47-49, 53, 55, 60