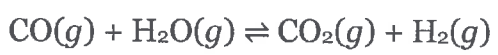


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CHEM 101B Chapter 12 Equilibrium – PreEquilibrium and i.c.e. tables

e.g.1 At 700K,  $K_c = 5.10$  for the following reaction:

$$K_c = \frac{[CO_2][H_2]}{[CO][H_2O]} = 5.10$$



A 1.000-liter flask is charged with 1.000 mole of each species. Calculate the equilibrium concentrations of all species.

i.c.e.  
add

	$CO(g)$	+	$H_2O(g)$	$\rightleftharpoons$	$CO_2(g)$	+	$H_2(g)$
Initial	1.000		1.000		1.000		1.000
Change	-x		-x		+x		+x
Equilibrium	1.000 - x <i>1.000 - .387</i>		1.000 - x <i>0.613M</i>		1.000 + x <i>1.000 + 387</i>		1.000 + x <i>1.387M</i>

$$Q = \frac{(1.000)(1.000)}{(1.000)(1.000)} = 1.000 < K$$

x = change (absolute)

$$\sqrt{5.10} = \frac{(1.000+x)^2}{(1.000-x)^2}$$

$$2.26 = \frac{1.000+x}{1.000-x}$$

check

$$K = \frac{(1.387)^2}{(.613)^2} = 5.12 \checkmark$$

3SF

$$2.26 - 2.26x = 1 + x$$

-1      +2.26x      -1      +2.26x

$$1.26 = 3.26x$$

$$x = \frac{1.26}{3.26} = 0.387$$

57. At 35°C,  $K = 1.6 \times 10^{-5}$  for the reaction



$$K = \frac{[\text{NO}]^2 [\text{Cl}_2]}{[\text{NOCl}]^2}$$

Calculate the concentrations of all species at equilibrium for each of the following original mixtures.

1.0 mole of NOCl and 1.0 mole of NO in a 1.0-L flask

**5% Rule:** Test the assumption that the "change" is less than 5% of the original value:

$$\frac{x}{[\text{orig conc}]} < 0.05$$



i.	1.0 M	1.0 M	0 M
c.	-2x	+2x	+x
e.	1.0 - 2x	1.0 + 2x	x

$$Q = 0 \longrightarrow$$

$$2x \ll 1.0$$

$$1.6 \times 10^{-5} = \frac{(1.0 + 2x)x}{1.0 - 2x}$$

*v. small*

$$1.6 \times 10^{-5} = x$$

$$.000016 = x$$

$$\frac{1.0}{0.000016}$$

$$1.000016$$

check

$$K = \frac{1.6 \times 10^{-5} (1.0)^2}{1.0}$$

check K

$$\frac{2(0.000016)}{1.0} < 0.05$$

$$= 1.6 \times 10^{-5}$$

59. At a particular temperature,  $K = 2.0 \times 10^{-6}$  for the reaction



If 2.0 moles of  $\text{CO}_2$  is initially placed into a 5.0-L vessel, calculate the equilibrium concentrations of all species.

$$K = \frac{[\text{CO}]^2 [\text{O}_2]}{[\text{CO}_2]^2}$$

$$Q = \frac{(0)^2 (0)}{(1.40)^2} = 0$$

→ shift

$$2\text{CO}_2 \rightleftharpoons 2\text{CO} + \text{O}_2$$

I.	.40	0	0
C.	-2x	+2x	+x
E.	.40-2x	2x	x

$$2.0 \times 10^{-6} = \frac{(2x)^2 (x)}{(.40 - 2x)^2}$$

$$2.0 \times 10^{-6} = \frac{4x^3}{(.40)^2}$$

$$2.0 \times 10^{-6} = \frac{4x^3}{.16}$$

$$8.0 \times 10^{-8} = x^3$$

$$x = \sqrt[3]{8.0 \times 10^{-8}}$$

$$x = 0.0043$$

$$[\text{CO}_2]_{\text{eq}} = .40 - 0.0043 \cdot 2 = 0.38 \text{ M}$$

$$[\text{CO}]_{\text{eq}} = 0.0043 \cdot 2 = 0.0086 \text{ M}$$

$$[\text{O}_2]_{\text{eq}} = .0043 \text{ M}$$

Assume  
 $2x \ll .40$

5% check

$$\frac{2(.0043)}{.40} < .05 ?$$

$$= .02 < .05 \checkmark$$

61. At 25°C,  $K_p = 2.9 \times 10^{-3}$  for the reaction



In an experiment carried out at 25°C, a certain amount of  $\text{NH}_4\text{OCNH}_2$  is placed in an evacuated rigid container and allowed to come to equilibrium. Calculate the total pressure in the container at equilibrium.

$$K_p = P_{\text{NH}_3}^2 P_{\text{CO}_2}$$

Any species that is not in the equilibrium phase—usually solution or gas phase—is not part of equilibrium expression (i.e. exclude solids)



i.	—	0	0
c.	—	+2x	+x
e.	—	2x	x

$$2.9 \times 10^{-3} = (2x)^2 (x)$$

$$= 4x^3$$

$$7.25 \times 10^{-4} = x^3$$

$$x = \sqrt[3]{7.25 \times 10^{-4}}$$

$$= 0.0898$$

$$x = 0.090$$

$$P_{\text{NH}_3} = 2(0.090)$$

$$= 0.18 \text{ atm}$$

$$P_{\text{CO}_2} = 0.090 \text{ atm}$$

$$P_{\text{tot}} = 0.27 \text{ atm}$$