Warmup: *Calculating* Equilibrium Concentrations

Nitrogen gas and hydrogen gas react to produce ammonia

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

If 0.750 mol of nitrogen and 2.250 mol of hydrogen are placed in a 5L (5.0L) vessel, what is the equilibrium constant if the equilibrium mixture contains 0.060 mol of ammonia.?

Amount	N_2	$3H_2$	$2NH_3$
Initial	$\frac{0.750\text{ mol}}{5L} = 0.15\text{ mol/L}$	$\frac{2.250\text{ mol}}{5L} = 0.45\text{ mol/L}$	0
Change	$-x$	$-3x$	$+2x$
Equilibrium	$0.15 - x$	$0.45 - 3x$	$2x = 0.012$

$\frac{0.060\text{ mol}}{5L} = 0.012\text{ mol/L}$

N_2: $= 0.15 - 0.006 = 0.144$

H_2: $= 0.45 - 3(0.006) = 0.432$

$x = \frac{0.012}{2} = 0.006$

$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{(0.012)^2}{(0.144)(0.432)^3} = 0.0124 = 0.012$$

∴ the equilibrium constant is 0.012.

Station #1

When 1.0 mol CO and 3.0 mol H_2 are placed in a 10.00L vessel and allowed to come to equilibrium the mixture is found to contain 0.387 mol H_2O. Calculate the K for the reaction.

$$\_\_CO_{(g)} + \underline{3}\,H_2 \leftrightharpoons \_\_CH_{4(g)} + \_\_H_2O_{(g)}$$

Amount	CO	$3H_2$	CH_4	H_2O
Initial	$\frac{1.0\,mol}{10\,L} = 0.1\ mol/L$	$\frac{3.0\,mol}{10\,L} = 0.3\ mol/L$	0	0
Change	$-x$	$-3x$	$+x$	$+x$
Equilibrium	$0.1-x$	$0.3-3x$	x	$x = \frac{0.387\,mol}{10\,L}$

$= 0.1 - 0.0387$
$= 0.0613\ \frac{mol}{L}$

$= 0.3 - 3(0.0387)$
$= 0.1839\ \frac{mol}{L}$

$= 0.0387\ \frac{mol}{L}$

$= 0.0387\ mol/L$

$$K_{eq} = \frac{[CH_4][H_2O]}{[CO][H_2]^3}$$

$$= \frac{(0.0387)(0.0387)}{(0.0613)(0.1839)^3}$$

$$= 3.9$$

$\therefore$ the K_{eq} is 3.9.

Station #2

At 25°C the value of Kc for the following reaction is 82

$$\_\_\_ I_{2(g)} + \_\_\_ Cl_{2(g)} \leftrightharpoons \underline{2}\ ICl_{(g)}$$

If 0.83 moles of both $I_{2(g)}$ and $Cl_{2(g)}$ are placed in a 10L container at 25°C, what are the concentrations of the three gases at equilibrium?

10.0 L

Amount	$I_{2(g)}$	$Cl_{2(g)}$	$2ICl_{(g)}$
Initial	0.083 mol/L	0.083 mol/L	0
Change	$-x$	$-x$	$+2x$
Equilibrium	$0.083-x$	$0.083-x$	$2x$

$$K = \frac{[ICl]^2}{[I_2][Cl_2]}$$

$$82 = \frac{(2x)^2}{(0.083-x)(0.083-x)}$$

$$82 = \frac{(2x)^2}{(0.083-x)^2}$$

$$\pm\sqrt{82} = \frac{2x}{0.083-x}$$

$$\pm 9.055 = \frac{2x}{0.083-x}$$

$$9.055 = \frac{2x}{0.083-x}$$

$$0.7516 - 9.055x = 2x$$

$$0.7516 = 11.055x$$

$$x = 0.0679$$

$$-9.055 = \frac{2x}{0.083-x}$$

$$-0.7516 + 9.055x = 2x$$

$$-0.7516 = -7.055x$$

$$x = 0.1065$$

↑ too big will give a −ve conc.

$\therefore [I_2] = 0.083 - 0.0679$
$= 0.015$ mol/L

$[Cl_2] = 0.015$ mol/L

$[ICl] = 2(0.0679)$
$= 0.14$ mol/L

Station #3

The Keq for the following reaction is 4.8

$$\_\_\_ SO_{2(g)} + \_\_\_ NO_{2(g)} \rightleftharpoons \_\_\_ NO_{(g)} + \_\_\_ SO_{3(g)}$$

In a 1.0L container the chemist added 1.7×10^{-1} mol of $SO_{2(g)}$ to 1.1×10^{-1} mol of $NO_{2(g)}$. What are the equilibrium concentrations of all four gasses?

Amount	SO_2	NO_2	NO	SO_3
Initial	0.17 mol/L	0.11 mol/L	0	0
Change	−x	−x	+x	+x
Equilibrium	0.17−x	0.11−x	x	x

$$K_{eq} = \frac{[NO][SO_3]}{[SO_2][NO_2]}$$

$$4.8 = \frac{(x)(x)}{(0.17-x)(0.11-x)}$$

$$4.8 = \frac{x^2}{0.0187 - 0.28x + x^2}$$

$$0.08976 - 1.344x + 4.8x^2 = x^2$$

$$3.8x^2 - 1.344x + 0.08976 = 0$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$= \frac{+1.344 \pm \sqrt{(1.344)^2 - 4(3.8)(0.08976)}}{2(3.8)} = \frac{1.344 \pm 0.6648}{7.6} = 0.2643 \text{ OR } 0.08937 ✓$$

↑ too big

$[SO_2] = 0.17 - x$
$= 0.17 - 0.08937$
$= 0.081$ mol/L

$[NO_2] = 0.11 - x$
$= 0.11 - 0.08937$
$= 0.021$ mol/L

$[NO] = 0.089$ mol/L

$[SO_3] = 0.089$ mol/L

Station #4

When nitrogen gas and chlorine gas react to form nitrogen trichloride gas, Keq = 4.15×10^{-5}.

$$\_\_\_ N_{2(g)} + \underline{3}\ Cl_{2(g)} \leftrightharpoons \underline{2}\ NCl_{3(g)}$$

If 2.74 mol of nitrogen gas and 0.84 mol of chlorine gas are put in a 2.0 L reaction vessel, what is the equilibrium concentration of the nitrogen trichloride?

Amount	N_2	$3Cl_2$	$2NCl_3$
Initial	$\frac{2.74\ mol}{2.0\ L} = 1.37\ mol/L$	$\frac{0.84\ mol}{2.0\ L} = 0.42\ mol/L$	0
Change	$-x$	$-3x$	$+2x$
Equilibrium	$1.37 - x$	$0.42 - 3x$	$2x$

Check: $\frac{\text{Inital concentration}}{K_{eq}}$

$$= \frac{0.42}{4.15 \times 10^{-5}}$$

$$= 10120 \quad > 1000$$

$\therefore$ x is really small in comparison to the starting concentrations

$[N_2] = 1.37\ mol/L$

$[Cl_2] = 0.42\ mol/L$

$$K_{eq} = \frac{[NCl_3]^2}{[N_2][Cl_2]^3}$$

$$4.15 \times 10^{-5} = \frac{(2x)^2}{(1.37)(0.42)^3}$$

* lost the x value

$$4.212 \times 10^{-6} = 4x^2$$

$$x = \pm\sqrt{\frac{4.212 \times 10^{-6}}{4}} = 0.001026$$

$[NCl_3] = 2(0.001026)$

$= 0.00205\ mol/L$

$= 0.0021\ mol/L$