Chemistry 40S

Equilibrium
Lesson 2 Homework Assignment

1.

a) $$K_c = \frac{[SO_3][NO]}{[SO_2][NO_2]}$$

b) $$K_c = \frac{[C_2H_6]}{[H_2]^3}$$

c) $$K_c = \frac{[O_3]^2}{[O_2]^3}$$

d) $K_c = [CO_2]$

e) $$K_c = \frac{[H^+]^6}{[Bi^{3+}]^2[H_2S]^3}$$

f) $$K_c = \frac{1}{[I_2]}$$

g) $$K_c = \frac{[PCl_5]}{[Cl_2][PCl_3]}$$

2.

$$K_c = \frac{[NH_3]^2}{[H_2]^3[N_2]}$$

$$K_c = \frac{(0.10)^2}{(0.15)^3(0.25)}$$

$$= \frac{(0.010)}{(0.003375)(.25)}$$

$$K_c = 11.85$$

Since the volume is 1 L, the concentrations can be determined by inspection.
Substitute concentrations and calculate.

3. $Br_2 = Cl_2 = 4.00$ moles
$BrCl = ?$

According to the reaction stoichiometry, if 4.00 moles of Cl_2 are produced the same amount of Br_2 will be produced.

$$K_c = \frac{[Cl_2][Br_2]}{[BrCl]^2}$$

$$[BrCl]^2 = \frac{[Cl_2][Br_2]}{K_c}$$

$$[BrCl] = \sqrt{\frac{[Cl_2][Br_2]}{K_c}}$$

$$[BrCl] = \sqrt{\frac{(4.00)(4.00)}{11.1}}$$

We can substitute these values into the equilibrium law and calculate the amount of BrCl at equilibrium.

$$[BrCl] = \sqrt{1.44}$$

$$\mathbf{[BrCl] = 1.20\ mol/L}$$

4. Let $H_2 = x$ and $I_2 = x$

According to the reaction stoichiometry, the moles of hydrogen and iodine produced should be equal. Since we do not know their equilibrium amounts, we can assign them a value of x.

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

$$0.020 = \frac{x \cdot x}{(0.50)^2}$$

$$x^2 = (0.020)(0.50)^2$$

Solve for x.

$$x = \sqrt{(0.020)(0.25)}$$

$$x = 0.07071 \approx 0.071$$

$H_2 = I_2 = 0.071$ moles

5. In (a) and (b), products are favoured over reactants since K>1. However, products are much more favoured in (a) than (b). The values in (c) and (d) favour the production of reactants as they are both less than 1.

6. $$K = \frac{[CO][H_2O]}{[CO_2][H_2]}$$

The volume is 1 L so concentrations can be done by inspection.

$$= \frac{(0.448)(0.448)}{(0.552)(0.552)}$$

Just substitute values and solve for K.

$K = 0.659$

7.

$$K = \frac{[CH_4][H_2S]^2}{[H_2]^4[CS_2]}$$

Rearrange the equilibrium law and solve for H_2S.

$$[H_2S]^2 = \frac{K\,[H_2]^4[CS_2]}{[CH_4]}$$

$$[H_2S] = \sqrt{\frac{(0.256)(0.316)^4(0.0898)}{(0.00108)}}$$

Don't forget H_2S is **squared**. You need to take the **square root**.

$$= \sqrt{\frac{(0.256)(0.009971)(0.0898)}{(0.00108)}}$$

$$= \sqrt{0.21225}$$

$$\mathbf{[H_2S] = 0.461\ mol/L}$$

8.

$$C = \frac{\text{moles}}{\text{Litres}} = \frac{n}{V}$$

First, find the concentrations of all species.

$$C_{P_4} = \frac{10.0\ \text{mol}}{5.00\ \text{L}} = 2.00\ \text{mol/L}$$

$$C_{H_2} = \frac{25.0\ \text{mol}}{5.00\ \text{L}} = 5.00\ \text{mol/L}$$

$$C_{PH_3} = \frac{5.00\ \text{mol}}{5.00\ \text{L}} = 1.00\ \text{mol/L}$$

$$K_c = \frac{[PH_3]^4}{[P_4][H_2]^6}$$

Write equilibrium law and substitute values.

$$K_c = \frac{(1.00)^4}{(5.00)(2.00)^6}$$

$$K_c = \frac{1.00}{(5.00)(64.0)}$$

$$\mathbf{K_c = 0.003125}$$

9.

$$C = \frac{\text{moles}}{\text{Litres}} = \frac{n}{V}$$

$$C_{CO} = \frac{3.00\text{ mol}}{0.500\text{ L}} = 6.00\text{ mol/L}$$

$$C_{CO_2} = \frac{4.00\text{ mol}}{0.500\text{ L}} = 8.00\text{ mol/L}$$

First, find the concentrations of all species.

$$K_c = \frac{[CO_2]}{[CO]}$$

$$K_c = \frac{(8.00)}{(6.00)}$$

$$\mathbf{K_c = 1.33}$$

Write equilibrium law and substitute values.

10.

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

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

$$[I_2] = \frac{(0.50)^2}{(0.10)(46.0)}$$

$$\mathbf{[I_2] = 0.054}$$

Write equilibrium law and rearrange to solve for $[I_2]$

11.

$$K = \frac{[NH_3]^2}{[H_2]^3[N_2]}$$

$$[NH_3]^2 = K[H_2]^3[N_2]$$

$$[NH_3] = \sqrt{K[H_2]^3[N_2]}$$

$$[NH_3] = \sqrt{(10.0)(0.600)^3(0.100)}$$

$$[NH_3] = \sqrt{0.216}$$

$$[NH_3] = 0.465\text{ mol/L}$$

Write equilibrium law and rearrange to solve for $[NH_3]$

$$n = C \bullet V$$

$$n = (0.465\text{ mol/L})(2.00\text{ L})$$

$$\mathbf{n = 0.930\text{ moles } NH_3}$$

Use the concentration to calculate the number of moles.

Chemistry 40S

Equilibrium
Lesson 3 Homework Assignment Key

1. $$[H_2] = [Cl_2] = \frac{n}{V} = \frac{2.00\ \text{mol}}{0.500\ \text{L}} = 4.00\ \text{mol/L}$$

	$H_{2(g)}$ +	$Cl_{2(g)}$	$\rightleftharpoons$ 2 $HCl_{(g)}$
I	4.00	4.00	0
C	$-x$	$-x$	$+2x$
E	$4.00-x$	$4.00-x$	$2x$

$$K_c = \frac{[HCl]^2}{[H_2][Cl_2]}$$

Write equilibrium law and substitute the equilibrium values.

$$76.0 = \frac{(2x)^2}{(4.00-x)^2}$$

$$\sqrt{76.0} = \sqrt{\frac{(2x)^2}{(4.00-x)^2}}$$

$$8.718 = \frac{2x}{4.00-x}$$

$$(4.00-x)8.718 = \left(\frac{2x}{4.00-x}\right)(4.00-x)$$

$$34.871 - 8.718x = 2x$$

$$34.871 = 2x + 8.718x = 10.718x$$

$$\frac{34.871}{10.718} = x$$

$$3.25 = x$$

$$[H_2] = [Cl_2] = 4.00 - x = 4.00 - 3.25 = 0.75\ \text{mol/L}$$

$$[HCl] = 2x = 2(3.25) = 6.50\ \text{mol/L}$$

Determine equilibrium concentrations using the value of x

CHEMICAL Equilibrium Assignment #2 - Key

2. $K = 1.00 \times 10^{-4}$

$$N_2(g) + O_2(g) \rightleftarrows 2NO(g)$$

$[N_2] = [O_2] = 1.00$ mol/L

	N_2	O_2	$\rightleftarrows 2NO$
I	1.00	1.00	0
C	$-x$	$-x$	$+2x$
E	$1.00-x$	$1.00-x$	$2x$

$$K_c = \frac{[NO]^2}{(1.00-x)(1.00-x)}$$

$$1.00\times10^{-4} = \frac{(2x)^2}{(1.00-x)^2}$$

$$\sqrt{1.00\times10^{-4}} = \sqrt{\frac{(2x)^2}{(1.00-x)^2}}$$

$$0.01 = \frac{2x}{1.00-x}$$

$$0.01 - 0.01x = 2x$$

$$0.01 = 2.01x$$

$$4.98\times10^{-3} = x$$

$[N_2] = [O_2] \cdot 1.00 - 4.98\times10^{-3} = 9.95\times10^{-1}$ mol/L

$[NO] = 2(4.98\times10^{-3}) = 9.96\times10^{-3}$ mol/L

moles [NO] in 10.0 L container $= \dfrac{9.96\times10^{-3}\text{ mol}}{1.00\,\cancel{L}} \times 10.0\,\cancel{L} = 9.96\times10^{-2}$ mol

3. We can disregard A, since it is solid and does not appear in the equilibrium law. We must first find the concentration of B:

$$C = \frac{n}{V} = \frac{4.84 \text{ moles}}{2.00 \text{ L}} = 2.42 \text{ mol/L B}$$

	$A_{(s)}$ +	$2\,B_{(g)}$	qe $2\,C_{(g)}$
I	---	2.42	0
C	---	$-x$	$+x$
E	---	$2.42-x$	x

Since the coefficients for both B and C are the same, every mole of B used up, produces a mole of C

$$K_c = \frac{[C]^2}{[B]^2}$$

Write equilibrium law and substitute the equilibrium values.

$$78.0 = \frac{x^2}{(2.42-x)^2}$$

$$\sqrt{78.0} = \sqrt{\frac{x^2}{(2.42-x)^2}}$$

$$8.832 = \frac{x}{2.42-x}$$

$$(2.42-x)8.832 = \left(\frac{x}{2.42-x}\right)(2.42-x)$$

$$21.37 - 8.832x = x$$

$$21.37 = 9.832x$$

$$\frac{21.37}{9.832} = x$$

$$2.17 = x$$

Determine equilibrium concentration of B using the value of x.

$$[B] = 2.42 - x = 2.42 - 2.17 = 0.25 \text{mol/L}$$

$$\text{moles B} = C \bullet V = (0.25 \text{ mol/L})(2.00 \text{ L})$$

$$= \mathbf{0.50 \text{ moles B}}$$

Find the number of moles of B by multiplying concentration times volume.

4. Determine the concentrations of oxygen and carbon dioxide. The carbon can be disregarded as it is a solid.

$$C_{O_2} = \frac{n}{V} = \frac{50.0 \text{ moles}}{2.00 \text{ L}} = 25.0 \text{ mol/L}$$

$$C_{CO_2} = \frac{n}{V} = \frac{2.00 \text{ moles}}{2.00 \text{ L}} = 1.00 \text{ mol/L}$$

Set up the chemical equation and use "ICE"

	$C(s)$ +	$O_2(g)$	qe $CO_2(g)$
I		25.0	1.00
C		$-x$	$+x$
E		$25.0 - x$	$1.00 + x$

$$K_c = \frac{[CO_2]}{[O_2]}$$

Write the equilibrium law and substitute equilibrium concentrations.

$$25.0 = \frac{1.00 + x}{25.0 - x}$$

$$25.0(25.0 - x) = \left(\frac{1.00 + x}{25.0 - x}\right)(25.0 - x)$$

Eliminate denominator by multiplying both sides by the denominator.

$$625 - 25.0x = 1.00 + x$$

$$624 = 26x$$

$$\frac{624}{26} = \frac{26x}{26}$$

$$24.0 = x$$

$$[CO_2] = 1.00 + x = 25.0 \text{mol/L}$$

Solve for x, then for the equilibrium concentration of CO_2.

$$\text{moles} = CV = (25.0 \text{ mol/L})(2.00\text{L})$$

$$\text{moles} = \mathbf{50.0 \text{ moles } CO_2}$$

Find moles of CO_2.

5. We can disregard NH_4Cl since it is a solid.

	$NH_4Cl_{(s)}$	$\rightleftharpoons NH_{3(g)}$	$+ HCl_{(g)}$
I	---	0.200	0.200
C	---	$-x$	$-x$
E	---	$0.200-x$	$0.200-x$

$$K_c = [NH_3][HCl]$$

$$3.50\times10^{-4} = (0.200-x)^2$$

$$\sqrt{3.50\times10^{-4}} = \sqrt{(0.200-x)^2}$$

$$0.01871 = 0.200-x$$

$$0.181 = x$$

$$[NH_3] = 0.200-x = 0.200-0.181$$

$$[NH_3] = \mathbf{0.019\ mol/L}\,[NH_3]$$

6. a) $Q=\dfrac{[C]^2}{[B]^2}=\dfrac{(30.0)^2}{(2.0)^2}=225$

b) The reverse reaction is favoured since Q>K.

c) [C] is increasing and [B] is decreasing.

7. Find the concentrations of each species, then determine the value of K_c.

$$[CO_2] = [CO] = \frac{5.0 \text{ mol}}{2.0 \text{ L}} = 2.5 \text{ mol/L}$$

$$[O_2] = \frac{0.20 \text{ mol}}{2.0 \text{ L}} = 0.10 \text{ mol/L}$$

$$K = \frac{[CO_2]^2}{[CO]^2[O_2]}$$

$$K = \frac{(2.5)^2}{(2.5)^2(0.10)}$$

$$K = 10.$$

$$Q = \frac{[CO_2]^2}{[CO]^2[O_2]} = \frac{(15.8)^2}{(10.0)^2(0.25)} \approx 10.$$

$Q = K$ so the system is at equilibrium.

8. Set up a modified "ICE" table:

	$2\,SO_{2(g)}$ +	$O_{2(g)}$	qe $2\,SO_{3(g)}$
Initial	5.0	10.0	0.0
Change	−4.2	−2.1	+4.2
Final	0.8	7.9	4.2

$$Q = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(4.2)^2}{(0.8)^2(7.9)} = 3.5$$

$Q < K$. This means that the system is not at equilibrium. The forward reaction is favoured.

$[SO_2]$ and $[O_2]$ are decreasing and $[SO_3]$ is increasing.

CHEMICAL EQUILIBRIUM ASSIGNMENT #3 – key

Q. $4NH_{3(g)} + 5O_{2(g)} \rightleftarrows 4NO_{2(g)} + 6H_2O_{(g)}$

at equilibrium: $[H_2O] = 0.100$ mol/L : $[O_2] = 2.00$ mol/L

$[NO_2] = 0.200$ mol/L $[NH_3] = 0.500$ mol/L

$$K_{eq} = \frac{[NO_2]^4[H_2O]^6}{[NH_3]^4[O_2]^5}$$

$$K_{eq} = \frac{(0.200)^4(0.100)^6}{(0.500)^4(2.00)^5}$$

$$K_{eq} = 8.00 \times 10^{-10}$$

IF: 0.75 mol H_2O $c = \frac{n}{V} = \frac{0.75 \text{ mol } H_2O}{3.0 \text{ L}} = 0.25$ mol/L H_2O

12.0 mol NO_2 $c = \frac{n}{V} = \frac{12.0 \text{ mol NO}}{3.0 \text{ L}} = 4.0$ mol/L NO_2

30.0 mol O_2 $c = \frac{n}{V} = \frac{30.0 \text{ mol } O_2}{3.0 \text{ L}} = 10.0$ mol/L O_2

0.30 mol NH_3 $c = \frac{n}{V} = \frac{0.30 \text{ mol } NH_3}{3.0 \text{ L}} = 0.10$ mol/L NH_3

$$Q = \frac{[NO_2]^4[H_2O]^6}{[NH_3]^4[O_2]^5}$$

$$Q = \frac{(4.0)^4(0.25)^6}{(0.10)^4(10.0)^5}$$

$$Q = 6.25 \times 10^{-3}$$

$Q > K$ ∴ The reverse reaction is favoured

[reactant] increases

[product] decreases

10. Set up an "ICE" table:

	$H_{2(g)}$ +	$I_{2(g)}$	qe $2\,HI_{(g)}$
Initial	6.90	2.40	0.00
Change	– 1.40	– 1.40	+ 2.80
Final	5.90	1.00	2.80

Use stoichiometry to determine change in H_2 and HI

5.50

Calculate Q:

$$Q = \frac{[HI]^2}{[H_2][I_2]} = \frac{(2.80)^2}{(5.50)(1.00)} = 1.43$$

Q < K, therefore the system is not at equilibrium.
The forward reaction is favoured.
$[H_2]$ and $[I_2]$ are decreasing while [HI] is increasing.