Answers to Equilibrium Problems

1. At 250°C the equilibrium constant for the following gaseous reaction is 0.041.

$$PCl_{5(g)} \leftrightarrow PCl_{3(g)} + Cl_{2(g)}$$

Calculate the concentrations of all of the substances present at equilibrium if 0.20 mol of PCl_5 are placed in a 4.0 L reaction vessel.

$$PCl_{5(g)} \leftrightarrow PCl_{3(g)} + Cl_{2(g)}$$

	PCl_5	PCl_3	Cl_2
[I]	0.050	0	0
[C]	-x	+x	+x
[E]	0.050 - x	x	x

$K_{eq} = \frac{[PCl_3][Cl_2]}{[PCl_5]}$

$0.041 = \frac{x^2}{0.050 - x}$

$0.041(0.050 - x) = x^2$

$0.00205 - 0.041x = x^2$

$x^2 + 0.041x - 0.00205 = 0$

$$x = \frac{-0.041 \pm \sqrt{0.041^2 - 4(-0.00205)}}{2}$$

X = 0.0292

Therefore, $[PCl_5] = 0.050 - 0.0292 = 0.021M$

$[PCl_3] = [Cl_2] = 0.0292M$

2. At 448°C the equilibrium constant for the following reaction is 50.0.

$$H_{2(g)} + I_{2(g)} \leftrightarrow 2HI_{(g)}$$

a) How many moles of HI will be present at equilibrium when 1.0 mol of $H_{2(g)}$ and 1.0 mol of $I_{2(g)}$ are allowed to reacted in a 1.0 L container?

	$H_{2(g)}$	$+ I_{2(g)}$ $\leftrightarrow$	$2HI_{(g)}$
[I]	1.0	1.0	0
[C]	-x	-x	+ 2x
[E]	1.0 - x	1.0 - x	2x

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

$50 = \frac{(2x)^2}{(1.0 - x)^2}$

$\sqrt{50} = \frac{2x}{1.0 - x}$

7.07(1.0 - x) = 2x

7.07 - 7.07x = 2x

7.07 = 9.07x

X = 0.78

Therefore, at equilibrium 2(0.78) mol/L x 1L = 1.6 mol of HI will be present.

b) How many moles of H_2 and I_2 remain unreacted?

1.0 - 0.78 = 0.22 mol/L

Therefore, the number of moles of unreacted H_2 and I_2 = 0.22mol/L x 1L = 0.22 mol

c) If the container was an open system and the reaction of H_2 and I_2 was complete (ie not an equilibrium reaction), how many moles of HI should be produced?

Since the reaction is a 1 : 1 : 2, 1.0 mol of H_2 should produce 2.0 mol HI.

d) What is the percent yield of the equilibrium mixture?

$\frac{1.6}{2.0}$ x 100% = 80%

3. A 1.0 L container contains 0.750 mol of CO and 0.275 mol of H_2O. After one hour, equilibrium is reached according the following equation:

$$CO_{(g)} + H_2O_{(g)} \leftrightarrow CO_{2(g)} + H_{2(g)}$$

Analysis shows that 0.25 mol of CO_2 is present. What is the equilibrium constant for the reaction?

$CO_{(g)} + H_2O_{(g)} \leftrightarrow CO_{2(g)} + H_{2(g)}$

	CO	H_2O	CO_2	H_2
[I]	0.750	0.275	0	0
[C]	- 0.25	- 0.25	+x	+ x
[E]	0.50	0.025	0.25	0.25

x = 0.25

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

$$= \frac{(0.25)^2}{(0.50)(0.025)}$$

$$= 5.0$$

4. Consider the equilibrium:

$$3I_{2(g)} + 6F_{2(g)} \leftrightarrow 2IF_{5(g)} + I_4F_{2(g)}$$

a) At a certain temperature, 3.0 mol of F_2 and 2.0 mol of I_2 are placed into a 10.0 L container. At equilibrium, the concentration of IF_5 is 0.020 mol/L. Calculate K_{eq} for the reaction. You will need to change 3.0 mol in 10.0L to concentration in mol/L.

	$3I_{2(g)}$ +	$6F_{2(g)}$ $\leftrightarrow$	$2IF_{5(g)}$	+ $I_4F_{2(g)}$
[I]	0.20	0.30	0	0
[C]	- 3x	- 6x	+2x	+x
[E]	0.20 - 3(0.010)	0.30 - 6(0.010)	0.020	0.010
	0.17	0.24		

2x = 0.020

X = 0.010

$$K_{eq} = \frac{[IF_5]^2\,[I_4F_2]}{[I_2]^3[F_2]^6}$$

$$= \frac{(0.020)^2\,(0.010)}{(0.17)^3\,(0.24)^6}$$

$$= 4.3$$

b) At a different temperature (this means that K_{eq} will be different than part a)), 6.0 mol of IF_5 and 8.0 mol of I_4F_2 are placed in a 10.0 L container. At equilibrium, 6.0mol of I_4F_2 are left. Calculate the K_{eq} for the new temperature. Again change moles to concentration.

NOTE: equilibrium starts on the right hand side

	$3I_{2(g)}$ +	$6F_{2(g)}$ $\leftrightarrow$	$2IF_{5(g)}$ +	$I_4F_{2(g)}$	
[I]	0	0	0.60	0.80	
[C]	+3x	+6x	-2x	-x	
[E]	3x	6x	0.60 - 2x	0.60	0.80 - x = 0.60
	3(0.20)	6(0.20)	0.60 - 2(0.20)		X = 0.20
	0.60	1.2	0.20	0.60	

$$K_{eq} = \frac{[IF_5]^2\,[I_4F_2]}{[I_2]^3[F_2]^6}$$

$$= \frac{(0.20)^2\,(0.60)}{(0.60)^3\,(1.2)^6}$$

$$= 0.037$$

Questions to consider:

Q? Did the new temperature cause the equilibrium to shift to the right or to the left?

Q? If the new temperature is higher, is the reaction as written exothermic or endothermic?

5. At a certain temperature, K_{eq} = 4.0 for the following reaction.

$$2HF_{(g)} \leftrightarrow H_{2(g)} + F_{2(g)}$$

Predict the direction in which the reaction will shift, if any, when the following amounts of substances are introduced into a 1.0 L container.

a) 3.0 mol of HF, 2.0 mol of H_2 and 4.0 mol of F_2

$$2HF_{(g)} \leftrightarrow H_{2(g)} + F_{2(g)}$$

[I] 3 2 4

$Q = \frac{[H_2][F_2]}{[HF]^2} = \frac{(2)(4)}{(3)^2} = 0.89$ This tells us that the reaction is NOT at equilibrium since if it was this would have given us the value of 4.0. The value needs to get bigger, so…….

The numerator needs to be larger and the denominator smaller => the reaction needs to shift to the right to reach equilibrium.

b) 0.20 mol of HF, 0.50 mol of H_2 and 0.60 mol of F_2

$Q = \frac{[H_2][F_2]}{[HF]^2} = \frac{(0.50)(0.60)}{(0.20)^2} = 7.5$ This tells us that the reaction is NOT at equilibrium.
The numerator needs to be smaller => the reaction needs to shift to the left to reach equilibrium.

c) 0.30 mol of HF, 1.8 mol of H_2 and 0.20 mol of F_2

$Q = \frac{[H_2][F_2]}{[HF]^2} = \frac{(1.8)(0.20)}{(0.30)^2} = 4.0$ This tells us that the reaction is at equilibrium.

6. The equilibrium constant for the following reaction is 7.0.

$$Br_{2(g)} + Cl_{2(g)} \leftrightarrow 2BrCl_{(g)}$$

If 0.080 mol of Br_2 and 0.60 mol of Cl_2 are placed into a 2.0 L container, what are the equilibrium concentrations for the reaction?

	$Br_{2(g)}$ +	$Cl_{2(g)}$ $\leftrightarrow$	$2BrCl_{(g)}$	$K_{eq} = 7.0$
[I]	0.040	0.30	0	
[C]	-x	-x	+2x	
[E]	0.040 – x	0.30 – x	2x	

$$7.0 = \frac{(2x)^2}{(0.040 - x)(0.30 - x)}$$

$7.0\,(0.012 - 0.30x - 0.04x + x^2) = 4x^2$

$7.0\,(0.012 - 0.34x + x^2) = 4x^2$

$0.084 - 2.38x + 7.0\,x^2 = 4x^2$

$3x^2 - 2.38x + 0.084 = 0$

$$x = \frac{2.38 \pm \sqrt{2.38^2 - 4(3)(0.084)}}{2(3)}$$

= 0.756 or 0.037 Note: 0.756 is too large – it would make $[Br_2]$ and $[Cl_2]$ negative

Therefore, $[Br_2] = 0.040 - 0.037 = 0.0030$ M

$[Cl_2] = 0.30 - 0.037 = 0.26$ M

$[BrCl] = 2(0.037) = 0.074$ M

7. At 425°C, the equilibrium constant is 1.82×10^{-2} for the reaction:

$$2HI_{(g)} \leftrightarrow H_{2(g)} + I_{2(g)}$$

Equilibrium is reached by adding HI to the reaction vessel.

a) What are the concentrations of H_2 and I_2 in equilibrium with 0.0100 mol/L HI?

	$2HI_{(g)}$ $\leftrightarrow$	$H_{2(g)}$ +	$I_{2(g)}$	$K_{eq} = 1.82 \times 10^{-2}$
[I]				
[C]				
[E]	0.0100	x	x	

$$1.82 \times 10^{-2} = \frac{x^2}{(0.0100)^2}$$

$X^2 = 1.82 \times 10^{-6}$

$X = 1.35 \times 10^{-3}$

Therefore, $[H_2] = [I_2] = 1.35 \times 10^{-3}$ M

b) What was the initial concentration of HI (i.e. before equilibrium was reached)?

	$2HI_{(g)}$ $\leftrightarrow$	$H_{2(g)}$ +	$I_{2(g)}$	$K_{eq} = 1.82 \times 10^{-2}$
[I]	y	0	0	
[C]	-2(0.00135)	+0.00135	+0.00135	
[E]	0.0100	+0.00135	0.00135	

$y - 2(0.00135) = 0.0100$

$y = 0.0127$

Therefore, the initial [HI] = 0.0127 M

c) What percent of HI reacted?

$$\% \text{ reacted} = \frac{0.00270}{0.0127} \times 100\% = 21.3\%$$

8. 1.00 mol of $CO_{(g)}$ and 1.00 mol $H_2O_{(g)}$ are placed in a 10.0 L container. At equilibrium, 0.665 mol of CO_2 and 0.665 mol of H_2 are present. The reaction proceeds as follows:

$$CO_{(g)} + H_2O_{(g)} \leftrightarrow CO_{2(g)} + H_{2(g)}$$

a) What are the equilibrium concentrations of all four gases?

	$CO_{(g)}$ +	$H_2O_{(g)}$ $\leftrightarrow$	$CO_{2(g)}$ +	$H_{2(g)}$
[I]	0.100	0.100	0	0
[C]	- 0.0665	- 0.0665	+ 0.0665	+ 0.0665
[E]	0.0335	0.0335	0.0665	0,0665

b) What is the value of K_{eq}?

$$K_{eq} = \frac{(0.0665)^2}{(0.0335)^2}$$

$= 3.94$

9. The reaction below is exothermic as written.

$$A_{(g)} + B_{(g)} \leftrightarrow C_{(g)} + \text{heat}$$

Assume that equilibrium has already been established. How would the concentration of C change with:

a) an increase in temperature? Shift left therefore [C] decreases
b) an increase in pressure? Shift right therefore [C] increases
c) an addition of A? Shift right therefore [C] increases
d) the addition of a catalyst? same
e) the removal of B? Shift left therefore [C] decreases
f) the removal of C? shift right but since C was removed, [C] won't be as high as initially, there decrease

How would the value of K_{eq} change with

g) an addition of A? same
h) an increase in temperature? Decrease because shifts left therefore [P] decreases and [R] increases
i) an addition of a catalyst? same

10. For each of the following equilibrium systems:

a) Write the equilibrium expression

b) State which direction the reaction would shift to reestablish equilibrium.

i) $2SO_{2(g)} + O_{2(g)} \leftrightarrow 2SO_{3(g)}$ (exothermic, temperature decrease)

$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$ Right

ii) $C_{(s)} + CO_{2(g)} \leftrightarrow 2CO_{(g)}$ (endothermic, increase in temperature)

$K_{eq} = \frac{[CO]^2}{[CO_2]}$ Right

ii) $N_2O_{4(g)} \leftrightarrow 2NO_{2(g)}$ (increased pressure)

$K_{eq} = \frac{[NO_2]^2}{[N_2O_4]}$ Left

iv) $CO_{(g)} + H_2O_{(g)} \leftrightarrow CO_{2(g)} + H_{2(g)}$ (decrease in pressure)

$K_{eq} = \frac{[CO_2][H_2]}{[CO][H_2O]}$ No change

v) $2NOBr_{(g)} \leftrightarrow 2NO_{(g)} + Br_{2(g)}$ (decrease pressure)

$K_{eq} = \frac{[NO]^2[Br_2]}{[NOBr]^2}$ Right

vi) $2O_{2(g)} + 3Fe_{(s)} + 4H_{2(g)} \leftrightarrow Fe_3O_{4(s)} + 4H_{2(g)}$ (add Fe)

$K_{eq} = \frac{[H_2]^4}{[O_2]^2[H_2]^4} = \frac{1}{[O_2]^2}$ No change (Fe is a solid; adding Fe does not increase the [Fe])

vii) $2SO_{2(g)} + O_{2(g)} \leftrightarrow 2SO_{3(g)}$ (add a catalyst)

$K_{eq} = \underline{[SO_3]^2}$ No change

$[SO_2]^2[O_2]$

viii) $CaCO_{3(s)} \leftrightarrow CaO_{(s)} + CO_{2(g)}$ (remove CO_2)

$K_{eq} = \frac{[CO_2]}{1}$ Right

ix) $N_{2(g)} + 3H_{2(g)} \leftrightarrow 2NH_{3(g)}$ (add $H_{2(g)}$)

$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$ Right

11. When at equilibrium, a reaction mixture contains: 0.20 mol H_2, 0.70 mol CO_2, 0.20 mol CO and 0.30 mol H_2O in a 1.0 L container. The reaction is as follows:

$$CO_{(g)} + H_2O_{(g)} \leftrightarrow CO_{2(g)} + H_{2(g)}$$

How many moles of CO_2 would have to be added to increase the amount of CO to 0.30 mol?

	$CO_{(g)}$ +	$H_2O_{(g)}$	$\leftrightarrow$	$CO_{2(g)}$ +	$H_{2(g)}$
[E]	0.20	0.30		0.70	0.20
Stress				+x	
[C]	+0.10	+0.10		-0.10	-0.10
New[E]	0.30	0.40		0.60 + x	0.10

$K_{eq} = \frac{(0.70)(0.20)}{(0.20)(0.30)} = 2.3$ (from first equilibrium)

$2.3 = \frac{(0.60 + x)(0.10)}{(0.30)(0.40)}$ (from new equilibrium)

$\frac{2.3(0.30)(0.40)}{0.10} = 0.60 + x$

X = 2.2

Therefore 2.2 moles of CO_2 need to be added to a 1 L container.