

## Answers to Equilibrium Problems

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



Calculate the concentrations of all of the substances present at equilibrium if 0.20 mol of  $\text{PCl}_5$  are placed in a 4.0 L reaction vessel.



[I]	0.050	0	0
[C]	-x	+x	+x
[E]	0.050 - x	x	x

$$K_{\text{eq}} = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]}$$

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

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

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

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

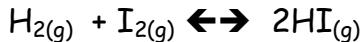
$$x = \frac{-0.041 \pm \sqrt{0.041^2 - 4(-0.00205)}}{2}$$

$$X = 0.0292$$

$$\text{Therefore, } [\text{PCl}_5] = 0.050 - 0.0292 = 0.021\text{M}$$

$$[\text{PCl}_3] = [\text{Cl}_2] = 0.0292\text{M}$$

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



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 react in a 1.0 L container?

$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$		
[I]	1.0	1.0
[C]	-x	-x
[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  $2(0.78) \text{ mol/L} \times 1\text{L} = 1.6 \text{ mol of HI will be present.}$

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

$$1.0 - 0.78 = 0.22 \text{ mol/L}$$

Therefore, the number of moles of unreacted  $H_2$  and  $I_2 = 0.22 \text{ mol/L} \times 1\text{L} = 0.22 \text{ mol}$

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?

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

3. A 1.0 L container contains 0.750 mol of CO and 0.275 mol of H<sub>2</sub>O. After one hour, equilibrium is reached according the following equation:



Analysis shows that 0.25 mol of CO<sub>2</sub> is present. What is the equilibrium constant for the reaction?



[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:



- a) At a certain temperature, 3.0 mol of F<sub>2</sub> and 2.0 mol of I<sub>2</sub> are placed into a 10.0 L container. At equilibrium, the concentration of IF<sub>5</sub> is 0.020 mol/L. Calculate K<sub>eq</sub> for the reaction. You will need to change 3.0 mol in 10.0L to concentration in mol/L.

3I <sub>2(g)</sub> +	6F <sub>2(g)</sub>	$\rightleftharpoons$	2IF <sub>5(g)</sub> + I <sub>4</sub> F <sub>2(g)</sub>
[I] 0.20	0.30		0 0
[C] - 3x	- 6x	+2x	+x
[E] 0.20 - 3(0.010) 0.17	0.30 - 6(0.010) 0.24	0.020	0.010

$$2x = 0.020$$

$$X = 0.010$$

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

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

$$= 4.3$$

- 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, 6.0 mol 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

	$3I_2(g)$	$\rightleftharpoons$	$2IF_5(g)$	$+ I_4F_2(g)$
[I]	0		0.60	0.80
[C]	$+3x$		$-2x$	$-x$
[E]	$3x$		$0.60 - 2x$	$0.60 - x$
	$3(0.20)$		$0.60 - 2(0.20)$	$0.60$
	0.60		0.20	0.60
	1.2			

$$0.80 - x = 0.60$$

$$X = 0.20$$

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

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

$$= 0.037$$

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.



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<sub>2</sub> and 4.0 mol of F<sub>2</sub>



$$[\text{I}] \quad 3 \quad \quad \quad 2 \quad \quad \quad 4$$

$$Q = \frac{[\text{H}_2][\text{F}_2]}{[\text{HF}]^2} = \frac{(2)(4)}{(3)^2} = 0.89 \quad \text{This tells us that the reaction is NOT at equilibrium since if it was this would have given us the value of 4.0. The value needs to get bigger, so.....}$$

The numerator needs to be larger and the denominator smaller  $\Rightarrow$  the reaction needs to shift to the right to reach equilibrium.

- b) 0.20 mol of HF, 0.50 mol of H<sub>2</sub> and 0.60 mol of F<sub>2</sub>

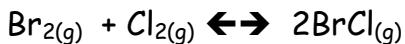
$$Q = \frac{[\text{H}_2][\text{F}_2]}{[\text{HF}]^2} = \frac{(0.50)(0.60)}{(0.20)^2} = 7.5 \quad \text{This tells us that the reaction is NOT at equilibrium.}$$

The numerator needs to be smaller  $\Rightarrow$  the reaction needs to shift to the left to reach equilibrium.

- c) 0.30 mol of HF, 1.8 mol of H<sub>2</sub> and 0.20 mol of F<sub>2</sub>

$$Q = \frac{[\text{H}_2][\text{F}_2]}{[\text{HF}]^2} = \frac{(1.8)(0.20)}{(0.30)^2} = 4.0 \quad \text{This tells us that the reaction is at equilibrium.}$$

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



If 0.080 mol of  $\text{Br}_2$  and 0.60 mol of  $\text{Cl}_2$  are placed into a 2.0 L container, what are the equilibrium concentrations for the reaction?

$\text{Br}_{2(g)}$	$\text{Cl}_{2(g)}$	$\rightleftharpoons$	$2\text{BrCl}_{(g)}$	$K_{eq} = 7.0$
[I] 0.040	0.30		0	
[C] -x	-x		+2x	
[E] 0.040 - x	0.30 - x		2x	

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

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

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

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

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

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

$$= 0.756 \text{ or } 0.037 \quad \text{Note: } 0.756 \text{ is too large - it would make } [\text{Br}_2] \text{ and } [\text{Cl}_2] \text{ negative}$$

$$\text{Therefore, } [\text{Br}_2] = 0.040 - 0.037 = 0.0030 \text{ M}$$

$$[\text{Cl}_2] = 0.30 - 0.037 = 0.26 \text{ M}$$

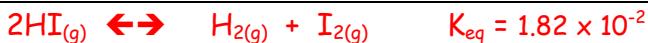
$$[\text{BrCl}] = 2(0.037) = 0.074 \text{ M}$$

7. At  $425^{\circ}\text{C}$ , the equilibrium constant is  $1.82 \times 10^{-2}$  for the reaction:



Equilibrium is reached by adding HI to the reaction vessel.

a) What are the concentrations of  $\text{H}_2$  and  $\text{I}_2$  in equilibrium with 0.0100 mol/L HI?



[I]		
[C]		
[E]	0.0100	x

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

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

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

$$\text{Therefore, } [\text{H}_2] = [\text{I}_2] = 1.35 \times 10^{-3} \text{ M}$$

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

$2\text{HI}_{(g)} \rightleftharpoons \text{H}_{2(g)} + \text{I}_{2(g)}$	$K_{\text{eq}} = 1.82 \times 10^{-2}$
[I]	y
[C]	-2(0.00135)
[E]	0.0100

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	0	0
	+0.00135	+0.00135
	0.00135	0.00135

$$y - 2(0.00135) = 0.0100$$

$$y = 0.0127$$

$$\text{Therefore, the initial } [\text{HI}] = 0.0127 \text{ M}$$

c) What percent of HI reacted?

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

8. 1.00 mol of  $\text{CO}_{(g)}$  and 1.00 mol  $\text{H}_2\text{O}_{(g)}$  are placed in a 10.0 L container. At equilibrium, 0.665 mol of  $\text{CO}_{2(g)}$  and 0.665 mol of  $\text{H}_2$  are present. The reaction proceeds as follows:



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

$\text{CO}_{(g)}$	$\text{H}_2\text{O}_{(g)}$	$\text{CO}_{2(g)}$	$\text{H}_2$
[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.



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

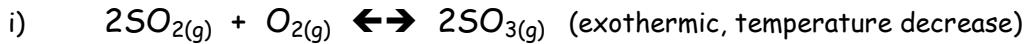
- a) an increase in temperature? Shift left therefore [C] decreases
- b) an increase in pressure? Shift right therefore [C] increases
- c) an addition of A? Shift right therefore [C] increases
- d) the addition of a catalyst? same
- e) the removal of B? Shift left therefore [C] decreases
- f) the removal 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:

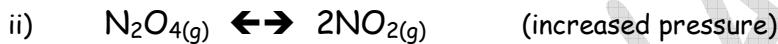
- Write the equilibrium expression
- State which direction the reaction would shift to reestablish equilibrium.



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



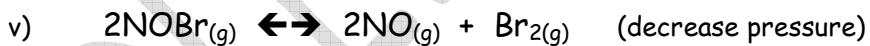
$$K_{eq} = \frac{[CO]^2}{[CO_2]} \quad \text{Right}$$



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



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



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

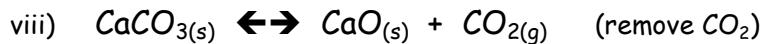


$$K_{eq} = \frac{[H_2]^4}{[O_2]^2 [H_2]^4} = \frac{1}{[O_2]^2} \quad \text{No change (Fe is a solid; adding Fe does not increase the [Fe])}$$



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

$$[SO_2]^2 [O_2]$$



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

Right



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

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:



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

	$CO_{(g)} + H_{2O(g)} \rightleftharpoons CO_{2(g)} + H_{2(g)}$			
[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.

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