Key

Equilibrium (Ch 18)

Basics (pg 552-559)

1. What does the term "dynamic equilibrium" mean?

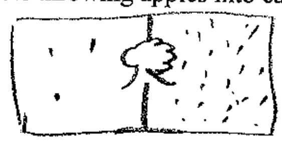
Forward & reverse rxns keep happening even after equilibrium is reached

2. What does it mean for a reaction to be at equilibrium?

Rate of forward & reverse reactions are equal

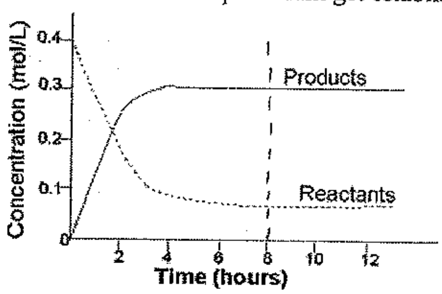
3. How does a reaction at equilibrium compare to two neighbors throwing apples into each other's yard?

See notes for full description



K > 1

4. At what time does equilibrium get established?



~8 hr

5. What is the equilibrium expression? What goes into it? What doesn't? How are coefficients incorporated?

↳ fraction comparing equilibrium concentrations of products over reactants

↳ no solids or liquids

↳ coefficients ⇒ exponents

6. What is the equilibrium constant? What does its value imply about a reaction when it is at equilibrium?

K ; If K > 1, @ eq., [Products] > [Reactants]

If K < 1, @ eq., [Products] < [Reactants]

7. How can you change a reaction's equilibrium constant?

Δ Temp

8. Consider the following reactions at some temperature:

$2NOCl\ (g) \leftrightarrow 2NO\ (g) + Cl_2\ (g)$ $K = 1.6 \times 10^{-5}$ K < 1, so more reactants @ eq.

$2NO\ (g) \leftrightarrow N_2\ (g) + O_2 (g)$ $K = 1 \times 10^{31}$ K > 1, so more products @ eq.

Each reaction is done in separate containers, and each is allowed to reach equilibrium. Describe the relative amounts of reactants and products in each container.

9. Consider the following generic reaction:

$2A_2B\ (g) \leftrightarrow 2A_2\ (g) + B_2\ (g)$

Some molecules of A_2B are placed in a 1.0 L container. As time passes, several snapshots of the reaction mixture are taken as illustrated below.

Which illustration is the first to represent an equilibrium mixture? Explain.

Eq

The 2nd & 3rd pics are identical, which means concentrations are staying steady

How many molecules of A_2B were initially placed in the container?

6; there are 6 B's

10. Write the expression for the equilibrium constant, K, for the reactions below:

a. $N_2\ (g) + 3H_2\ (g) \rightleftharpoons 2NH_3\ (g)$

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

b. $2KClO_3\ (s) \rightleftharpoons 2KCl\ (s) + 3O_2\ (g)$

$$K = [O_2]^3$$

c. $H_2O\ (l) \rightleftharpoons H^+\ (aq) + OH^-\ (aq)$

$$K = [H^+][OH^-]$$

d. $2CO\ (g) + O_2\ (g) \rightleftharpoons 2CO_2\ (g)$

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

e. $Li_2CO_3\ (s) \rightleftharpoons 2Li^+\ (aq) + CO_3^{2-}\ (aq)$

$$K = [Li^+]^2[CO_3^{2-}]$$

11. Using the equilibrium constant expressions you determined in the previous problem, calculate K when:

a. $[NH_3] = 0.010$ M, $[N_2] = 0.020$ M, $[H_2] = 0.020$ M

$K = 625$

b. $[O_2] = 0.050$ M

$K = 0.000125$

c. $[H^+] = 1x10^{-8}$ M, $[OH^-] = 1x10^{-6}$ M

1×10^{-14}

d. $[CO] = 2.0$ M, $[O_2] = 1.5$ M, $[CO_2] = 3.0$ M

$K = 1.5$

e. $[Li^+] = 0.2$ M, $[CO_3^{2-}] = 0.1$ M

$K = 0.004$

12. $PCl_5(g) \rightarrow PCl_3(g) + Cl_2(g)$. What is the equilibrium constant if the equilibrium concentrations are as follows: PCl_5 is 0.0096 mol/L, PCl_3 is 0.0247 mol/L and Cl_2 is 0.0247 mol/L? Ans: 0.064

$$K = \frac{[PCl_3][Cl_2]}{[PCl_5]} = \frac{(0.0247)(0.0247)}{(0.0096)} = 0.064$$

13. At 1000°C, a 1.00 L container has an equilibrium mixture consisting of 0.102 mol of ammonia, 1.03 mol of nitrogen, and 1.62 mol of hydrogen. Calculate the K_{eq} for the equilibrium system.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Ans: 0.00238

$$K = \frac{[NH_3]^2}{[N_2][H_2]^3} = \frac{\left(\frac{0.102}{1}\right)^2}{\left(\frac{1.03}{1}\right)\left(\frac{1.62}{1}\right)^3} = 0.00238$$

14. At a given temperature, the K_{eq} for the reaction $2HI(g) \rightarrow H_2(g) + I_2(g)$ is 1.40×10^{-2}. If the concentration of both H_2 and I_2 at equilibrium are 2.00×10^{-4}M, find the concentration of HI.

$$1.40\times10^{-2} = \frac{(2\times10^{-4})(2\times10^{-4})}{[HI]^2} \qquad [HI] = \sqrt{\frac{(2\times10^{-4})^2}{(1.40\times10^{-2})}} = 0.00169\,M$$

Ans: 0.00169M

15. Acetic acid dissociates in water. If $K_{eq} = 1.80 \times 10^{-5}$ and the equilibrium concentrations of acetic acid is 0.09986M, what is the concentration of $H^+(aq)$ and $C_2H_3O_2^-(aq)$? Ans: 0.00134M

$$HC_2H_3O_2(aq) \rightarrow H^+(aq) + C_2H_3O_2^-(aq)$$

$$1.80\times10^{-5} = \frac{(x)(x)}{0.09986} \qquad \sqrt{x^2} = \sqrt{1.797\times10^{-6}}$$

$$x = 0.00134\,M = [H^+] = [C_2H_3O_2^-]$$

16. At 60.2°C the equilibrium constant for the reaction $N_2O_4(g) \rightarrow 2NO_2(g)$ is 8.75×10^{-2}. At equilibrium at this temperature a vessel contains N_2O_4 at a concentration of 1.72×10^{-2}M. What concentration of NO_2 does it contain?

Ans: 0.0388M

17. At equilibrium, K for the decomposition of HI(g) was found to be 1.07×10^{-5}. The equilibrium concentration of HI(g) was found to be 0.129M. Calculate the concentration of I_2 at equilibrium.

Ans: 0.000422M

18. Solve for the X's

[HI]	[H_2]	[I_2]	K_{eq}	
1.78	0.172	0.172	X	Ans: 0.00934
X	0.242	0.242	0.217	Ans: 0.519
0.78	0.112	X	2.06×10^{-2}	Ans: 0.112

ICE Tables –reactions not initially at equilibrium (not covered in our textbook)

19. What is an ICE table? What are they used for?

Initial, Change, Equilibrium ↳ when a reaction is not initially at equilibrium

20. For the system:
$CO_{2\,(g)} + H_{2\,(g)} \leftrightarrow CO_{(g)} + H_2O_{(g)}$

$K_{eq} = 0.64$ at 900K

	$CO_2(g)$ +	$H_2(g)$ ⇌	$CO(g)$ +	$H_2O(g)$
I	0.1	0.1	0	0
C	-x	-x	+x	+x
E	0.1-x	0.1-x	x	x

If we start with 0.100mol/L of CO_2 and H_2, what are the concentrations of the reactants and products at equilibrium?

Answer: $[CO] = [H_2O] = 0.044M$, $[CO_2] = [H_2] = 0.056M$

$[H_2O] = [CO] = 0.044\,M$

$$0.64 = \frac{(x)(x)}{(0.1-x)(0.1-x)}$$

$$\sqrt{0.64} = \sqrt{\frac{x^2}{(0.1-x)^2}}$$

$$0.8 = \frac{x}{0.1-x}$$

$0.08 - 0.8x = x$

$0.08 = 1.8x$

$0.044 = x$

$0.1 - 0.044 = 0.056$

$[CO_2] = [H_2] = 0.056\,M$

21. For the system:
$2HI_{(g)} \leftrightarrow H_{2\,(g)} + I_{2\,(g)}$

$K_{eq} = 0.016$

	$2HI(g)$ ⇌	$H_2(g)$ +	$I_2(g)$
I	0.096	0.01	0.01
C	-2x	+x	+x
E	0.096-2x	0.01+x	0.01+x

If we start with 0.010 mol/L of H_2 and I_2 and 0.096 mol/L of HI, what are their concentrations at equilibrium?

Answer: $[H_2] = [I_2] = 0.012M$, $[HI] = 0.092M$

$$0.016 = \frac{(0.01+x)(0.01+x)}{(0.096-2x)^2}$$

$$\sqrt{0.016} = \sqrt{\frac{(0.01+x)^2}{(0.096-2x)^2}}$$

$$0.126 = \frac{0.01+x}{0.096-2x}$$

$0.0121 - 0.252x = 0.01 + x$

$0.0021 = 1.252x$

$0.0017 = x$

$[H_2] = [I_2] = 0.01 + 0.0017 = 0.012\,M$

$[HI] = 0.096 - 2(0.0017) = 0.092\,M$

22. At 650°C, the reaction below has a K_{eq} value of 0.771. If 2.00 mol of both hydrogen and carbon dioxide are placed in a 4.00L container and allowed to react, what will be the equilibrium concentrations of all four gases?

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

Answer: $[CO] = [H_2O] = 0.234M$, $[CO_2] = [H_2] = 0.266M$

$H_2 + CO_2 \rightleftharpoons CO + H_2O$

	H_2	CO_2	CO	H_2O
I	0.5	0.5	0	0
C	$-x$	$-x$	$+x$	$+x$
E	$.5-x$	$.5-x$	x	x

$\sqrt{0.771} = \sqrt{\frac{x^2}{(.5-x)^2}}$

$0.878 = \frac{x}{.5-x}$

$0.439 - 0.878x = x$

$0.439 = 1.878x$

$0.234 = x$

$[H_2O] = [CO] = 0.234\,M$

$0.5 - x = 0.266$

$[H_2] = [CO_2] = 0.266\,M$

23. Carbonyl bromide, $COBr_2$, can be formed by reacting CO with Br_2. The equation for the reaction is:

$CO_{(g)} + Br_{2\,(g)} \leftrightarrow COBr_{2\,(g)}$ $K_{eq} = 5.26$

	CO	Br_2	$COBr_2$
I	0.08	0.06	0.004
C	$-x$	$-x$	$+x$
E	$.08-x$	$.06-x$	$.004+x$

A mixture of 0.400 mol CO, 0.300 mol Br_2, and 0.0200 mol $COBr_2$ is sealed in a 5.00L flask. Calculate equilibrium concentrations for all gases.

Answer: $[COBr_2] = 0.0167M$, $[CO] = 0.0673M$, $[Br_2] = 0.0473M$

Use the formula: $\frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$

$5.26 = \frac{0.004+x}{(0.08-x)(0.06-x)}$

Quadratic

$x = 0.013$

24. 0.463 mol/L of HI(g) is placed in a 1.00 L reaction vessel. The temperature is raised to 300°C and maintained until equilibrium is established. At equilibrium, the vessel contains 0.119 mol/L of hydrogen gas and 0.119 mol/L iodine gas. Are the reactants or the products favored? Justify your answer.

$2HI \rightleftharpoons H_2 + I_2$

	$2HI$	H_2	I_2
I	0.463	0	0
C	$-2x$	$+x$	$+x$
E	$0.463-2x$	x	x

$[H_2] = [I_2] = 0.119\,M$

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

$K < 1$

$[HI] = 0.225\,M$

$x = 0.119$

$K = \frac{(.119)(.119)}{(.225)^2} = 0.28$ so, reactants are favored

$2BrCl \rightleftharpoons Br_2 + Cl_2$
I .337 0 0
C $-2x$ $+x$ $+x$
E $.337-2x$ x x

25. An equilibrium is established when bromine monochloride, BrCl(g), decomposes to form its elements at 200°C. If the initial concentration of BrCl(g) is 0.337 mol/L, and if the equilibrium concentration of chlorine gas, $Cl_2(g)$, is measured to be 0.0215 mol/L. Are the reactants or products favored?

$X = 0.0215$

$$K = \frac{(x)(x)}{(.337-2x)^2}$$

$K = 0.0053$ $K < 1$

Reactants

26. At high temperatures, nitrogen and oxygen gases react to produce nitrogen monoxide. Calculate the equilibrium concentration of nitrogen monoxide if the equilibrium concentrations of oxygen and nitrogen are 0.457 mol/L and K is 1.3×10^{-4} at 1800 K.

$N_2 + O_2 \rightleftharpoons 2NO$
I ? ? 0
C $-x$ $-x$ $+2x$
E 0.457 0.457 $2x$

$$K = \frac{(NO)^2}{(N_2)(O_2)}$$

$$1.3 \times 10^{-4} = \frac{(2x)^2}{(.457)^2}$$

$X = 0.0026$ M

$[NO] = 2(.0026) = 0.0052M$

27. Initial concentrations of $H_{2(g)}$ and $Cl_{2(g)}$ in a flask were 0.576 mol/L. The mixture was allowed to reach equilibrium at 45°C. The equilibrium concentration of HCl(g) was found to be 0.356 mol/L. Calculate the equilibrium concentrations of the reactants.

$H_2 + Cl_2 \rightleftharpoons 2HCl$
0.576 0.576 0
$-x$ $-x$ $+2x$
$.576-x$ $.576-x$ $2x$

$2x = 0.356$
$x = 0.178$

$0.576 - x = 0.398$ M $= [H_2] = [Cl_2]$

28. Consider the following equilibrium:

$2NO_{2(g)} \rightleftharpoons N_2O_{4(g)}$

If 2.00 moles of NO_2 are placed in a 1.00 L flask and allowed to react. At equilibrium 1.80 moles NO_2 are present. Calculate the K_{eq}.

$2NO_2 \rightleftharpoons N_2O_4$
2 0
$-2x$ $+x$
$2-2x$ x

$2-2x = 1.8$
$X = 0.1$

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

$$K = \frac{(0.1)}{(2-2(0.1))^2}$$

$K = 0.0309$

★29. Consider the equilibrium:

$I_{2\,(g)} + Cl_{2\,(g)} \rightleftharpoons 2ICl_{(g)}$ Keq= 10.0

The same number of moles of I_2 and Cl_2 are placed in a 1.0L flask and allowed to reach equilibrium. If the equilibrium concentration of ICl is 0.040 M, calculate the initial number of moles of I_2 and Cl_2.

$I_2 + Cl_2 \rightleftharpoons 2ICl$
? ? 0
$-x$ $-x$ $+2x$
$?-x$ $?-x$ $2x$

$2x = 0.04$
$x = 0.02$

$$10 = \frac{(2x)^2}{(?-x)(?-x)}$$

$$10 = \frac{(2(0.02))^2}{(?-0.02)^2}$$

$? = 0.0326$ M $= [I_2] = [Cl_2]$

★30. Consider the equilibrium: $2ICl_{(g)} \rightleftharpoons I_{2\,(g)} + Cl_{2\,(g)}$ Keq= 10.0
If x moles of ICl were placed in a 5.0 L container at 10 °C and if an equilibrium concentration of I_2 was found to be 0.60 M, calculate the number of moles ICl initially present.

$2ICl \rightleftharpoons I_2 + Cl_2$
? 0 0
$-2x$ $+x$ $+x$
$?-2x$ x x

$X = 0.6$

$$10 = \frac{(x)(x)}{(?-2x)^2}$$

$$\sqrt{10} = \sqrt{\frac{x^2}{(?-2x)^2}}$$

0.6

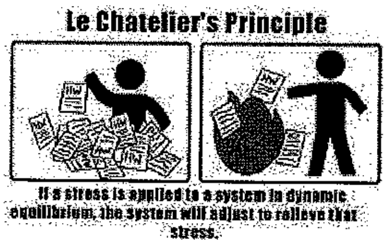
$$3.16 = \frac{x}{?-2x}$$

$? = 1.39$ M ICl

$1.39 = \frac{x\ mol}{5L}$

$X = 6.95$ mol ICl

LeChâtelier's Principle (pg 562-568)



31. Ammonia is produced commercially by the Haber Process: $N_2 (g) + 3H_2 (g) \rightleftharpoons 2NH_3 (g)$
The formation of ammonia is favored by
   a. An increase in pressure
   b. A decrease in pressure
   c. Removal of N_2 (g)
   d. Removal of H_2 (g)
32. Given the reaction at equilibrium: $2SO_2 (g) + O_2 (g) \rightleftharpoons 2SO_3 (g)$ + heat
What change will shift the equilibrium to the right?
   a. Increase temp
   b. Increase pressure
   c. Decrease amount of SO_2
   d. Decrease amount of O_2
33. Which system at equilibrium will be the least affected by a change in pressure?
   a. $N_2 (g) + 3H_2 (g) \rightleftharpoons 2NH_3 (g)$
   b. $2S (s) + 3O_2 (g) \rightleftharpoons 2SO_3 (g)$
   c. $AgCl (s) \rightleftharpoons Ag^+ (aq) + Cl^- (aq)$
   d. $2HgO (s) \rightleftharpoons 2Hg (l) + O_2 (g)$
34. Given the system at equilibrium: $CO_2 (g) \rightleftharpoons CO_2 (aq)$
As the pressure on the system increases, the solubility of CO_2 (g) (hint: Henry's Law)
   a. Decreases
   b. Increases
   c. Stays the same
35. Given the reaction at equilibrium: $2SO_2 (g) + O_2 (g) \rightleftharpoons 2SO_3 (g)$ + heat
When the pressure on the system is increased, the concentration of SO_3 will
   a. Decrease
   b. Increase
   c. Stay the same
36. Given the reaction at equilibrium: $2A (g) + 3B (g) \rightleftharpoons A_2B_3 (g)$ + heat
Which change will not affect the equilibrium concentrations of A (g), B (g), and A_2B_3 (g)?
   a. Add more A (g)
   b. Add a catalyst
   c. Increase the temp
   d. Increase the pressure

37. The addition of a catalyst to a system at equilibrium will increase the rate of
   a. The forward reaction only
   b. The reverse reaction only
   c. Both the forward and reverse reactions
   d. Neither the forward nor reverse reaction
38. For a given system at equilibrium, lowering the temperature will always
   a. ~~Increase the rate of the reaction~~
   b. ~~Increase the concentration of the products~~
   c. Favor an exothermic reaction
   d. Favor an endothermic reaction
39. Consider the reaction

$Fe^{3+} (aq) + SCN^- (aq) \leftrightarrow FeSCN^{2+} (aq)$

How will the equilibrium position shift if

   a. $AgNO_3$ (aq) is added *AgSCN is insoluble Left
   b. NaOH (aq) is added *$Fe(OH)_3$ is insoluble Left
   c. $Fe(NO_3)_3$ (aq) is added Right
   d. Water is added, doubling the volume of the solution? Decreases all concentrations

Orange Yellow

40. $Cr_2O_7^{2-}$ (aq) + H_2O (l) ↔ $2CrO_4^{2-}$ (aq) + $2H^+$ (aq)

Dichromate = $Cr_2O_7^{2-}$ Chromate = CrO_4^{2-}

Explain why orange dichromate solutions turn yellow when NaOH is added.

OH^- reacts with H^+, which uses up H^+

41. Novelty devices for predicting rain contain cobalt (II) chloride and are based on the following equilibrium:
$CoCl_2$ (s) + $6H_2O$ ~~(l)~~ (g) ↔ $CoCl_2{\cdot}6H_2O$ (s)
purple pink

What color will such a device be when rain is imminent? Why?

Pink; H_2O in the air shifts it right

42. N_2 (g)+ $3H_2$ (g) ↔ $2NH_3$ (g) + heat

At 300°C, K = 0.00434
At 500°C, K = 0.0000145
At 600°C, K = 0.00000225

hotter = smaller K = more reactants

Is this reaction exothermic or endothermic? Explain. Remember, $K = \frac{[products]}{[reactants]}$

43. Use the following reaction to fill in the table below: N_2 (g) + $3H_2$ (g) ⇌ $2NH_3$ (g) + heat

Stress	Equilibrium Shift	$[N_2]$	$[H_2]$	$[NH_3]$	K
1. Add N_2	Right	—	Decreases	Increases	Remains the same
2. Add H_2	R	↓	—	↑	—
3. Add NH_3	L	↑	↑	—	—
4. Remove N_2	L	—	↑	↓	—
5. Remove H_2	L	↑	—	↓	—
6. Remove NH_3	R	↓	↓	—	—
7. Increase Temp	L	↑	↑	↓	↓
8. Decrease Temp	R	↓	↓	↑	↑
9. Increase Pressure	R	↓	↓	↑	—
10. Decrease Pressure	L	↑	↑	↓	—

44. Use the following reaction to fill in the table below: 12.6 kcal + H_2 (g) + I_2 (g) $\rightleftharpoons$ 2HI (g)

Stress	Equilibrium Shift	$[H_2]$	$[I_2]$	[HI]	K
1. Add H_2	Right	—	Decreases	Increases	Remains the same
2. Add I_2	R	↓	—	↑	—
3. Add HI	L	↑	↑	—	—
4. Remove H_2	L	—	↑	↓	—
5. Remove I_2	L	↑	—	↓	—
6. Remove HI	R	↓	↓	—	—
7. Increase Temp	R	↓	↓	↑	↑
8. Decrease Temp	L	↑	↑	↓	↓
9. Increase Pressure	—	—	—	—	—
10. Decrease Pressure	—	—	—	—	—

45. Use the following reaction to fill in the table below: NaOH (s) $\rightleftharpoons$ Na^+ (aq) + OH^- (aq) + 10.6 kcal

Stress	Equilibrium Shift	Amount of NaOH (s)	$[Na^+]$	$[OH^-]$	K
1. Add NaOH (s)	—	—	—	—	—
2. Add NaCl (aq) (adds Na^+)	L	↑	—	↓	—
3. Add KOH (aq) (adds OH^-)	L	↑	↓	—	—
4. Add HCl (aq) (removes OH^-)	R	↓	↑	—	—
5. Increase Temp	L	↑	↓	↓	↓
6. Decrease Temp	R	↓	↑	↑	↑
7. Increase Pressure	—	—	—	—	—
8. Decrease Pressure	—	—	—	—	—