LeChatelier's Principle Practice Sheet

KEY

1. What three types of changes shift the position of a chemical equilibrium?
2. For each of the following chemical systems at equilibrium, use Le Châtelier's principle to predict the effect of the change imposed on the chemical system. Indicate the direction in which the equilibrium is expected to shift. For each example, sketch the graph of concentrations versus time, plotted from just before the change to the established new equilibrium.
   (a) $H_2O(l)$ + energy ⇄ $H_2O(g)$ P
   The container is heated.
   (b) $H_2O(l)$ ⇄ $H^+(aq) + OH^-(aq)$ R
   A few crystals of NaOH(s) are added to the container.
   (c) $CaCO_3(s)$ + energy ⇄ $CaO(s) + CO_2(g)$ P
   $CO_2(g)$ is removed from the container.
   (d) $CH_3COOH(aq)$ ⇄ $H^+(aq) + CH_3COO^-(aq)$ P
   A few drops of pure $CH_3COOH(l)$ are added to the system.
3. Much methanol is produced industrially by the exothermic reaction $CO(g) + 2H_2(g)$ ⇄ $CH_3OH(l)$, carried out at high pressure (5–10 MPa) and temperature (250 °C) in the presence of several catalyst substances. Methanol is less flammable than gasoline, and so it is a safer fuel. It is the fuel used in open-wheel Champ Car racing, and also in the Indianapolis 500.
   (a) State, in terms of forward and reverse reaction rates, why using a very high pressure of the reactant gases is economically desirable for the manufacturer. High pressure causes shift to products
   (b) State in which direction a high temperature will shift this reaction equilibrium. reactants
   (c) Explain why using a high temperature is desirable, in terms of the time required for the reaction to reach equilibrium. high temp means more collisions ∴ faster reactions

1. The following equation represents part of the industrial production of nitric acid. Predict the direction of the equilibrium shift for each of the following changes. Explain any shift in terms of the changes in forward and reverse reaction rates.

   $4NH_3(g) + 5O_2(g)$ ⇄ $4NO(g) + 6H_2O(g)$ + energy

   (a) $O_2(g)$ is added to the system. P
   (b) The temperature of the system is increased. R
   (c) NO(g) is removed from the system. P
   (d) The pressure of the system is increased by decreasing the volume. R
2. The following chemical equilibrium system is part of the Haber process for the production of ammonia.

   $N_2(g) + 3H_2(g)$ ⇄ $2NH_3(g)$ + energy

   Suppose you are a chemical process engineer. Use Le Châtelier's principle to predict five specific changes that you might impose on the equilibrium system to increase the yield of ammonia.
3. In a solution of copper(II) chloride, the following equilibrium exists:

1) ↑ $[N_2]$
2) ↑ $[H_2]$
3) ↓ $[NH_3]$
4) ↑ P
5) ↓ T

6. Chloromethane (methyl chloride) is manufactured by "chlorinating" methane. For this reaction system at equilibrium, explain the effect of each of the imposed changes on the position of reaction equilibrium.

   $CH_4(g) + Cl_2(g)$ ⇄ $CH_3Cl(g) + HCl(g)$ $\Delta_r H$ is negative

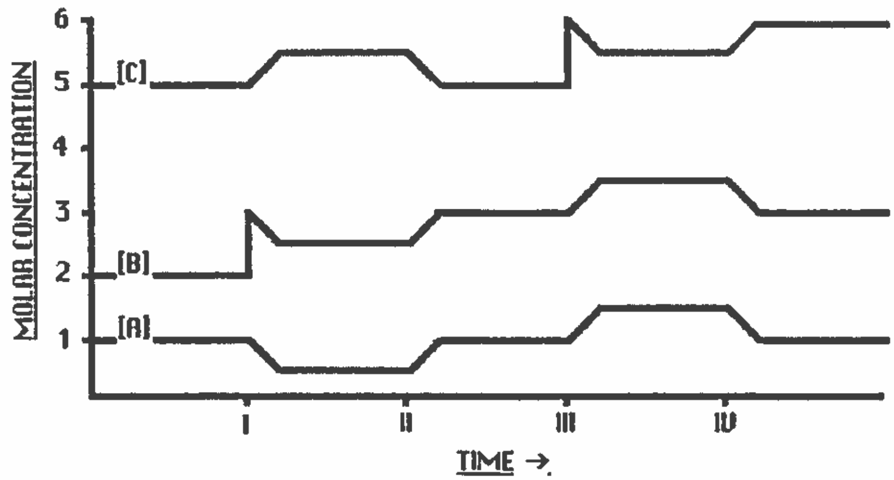
   (a) More methane is injected into the reaction vessel. P
   (b) The container volume is increased. no change
   (c) The temperature is lowered. P
   (d) A catalyst is introduced into the system. no change
7. Ethyne (acetylene) is manufactured by a high-temperature combustion of methane, using a large excess of methane. For this endothermic reaction system at equilibrium, explain the effect of each of the imposed changes on the value of the equilibrium constant.

   $6CH_4(g) + O_2(g)$ ⇄ $2C_2H_2(g) + 10H_2(g) + 2CO(g)$

   (a) More methane is injected into the reaction vessel. P
   (b) The container volume is decreased. R
   (c) The temperature is lowered. R
   (d) A catalyst is introduced into the system. no change

16. Given the following graph showing the concentrations of species A, B and C, state what changes in **temperature** or **concentration** are responsible for each of the shifts shown on the graph. The equilibrium equation is:

$$A_{(g)} + B_{(g)} \rightleftarrows C_{(g)} \qquad \Delta H = -65\ kJ$$



a) At time I, the [B] was increased

b) At time II, the temp was increased

c) At time III, the [C] was increased

d) At time IV, the temp was decreased

12. Given the following equilibrium system, state which way the equilibrium will shift when the changes below are made:

$$2C_2H_{6(g)} + 7O_{2(g)} \rightleftarrows 4CO_{2(g)} + 6H_2O_{(g)} + heat$$

a) The *volume* of the container is halved.................... Answer R

b) The *temperature* is decreased .... Answer P

c) CO_2 is added to the container................................. Answer R

d) The *total pressure* is increased .... Answer R

e) O_2 gas is removed from the system .... Answer R

f) Neon gas is added to increase the total pressure ..... Answer R

h) A *catalyst* is added.. Answer no chang

WWW WEB Activity

Simulation—Writing Equilibrium Expressions

This simulation allows you to select a reaction type and the initial reactant concentrations, which the program uses to plot the resulting equilibrium graph. You will then be guided through a series of questions.

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Section 15.1 Questions

1. Write a balanced equation with integer coefficients and the expression of the equilibrium law for each of the following reaction systems at fixed temperature.
   (a) Hydrogen gas reacts with chlorine gas to produce hydrogen chloride gas in the industrial process that eventually produces hydrochloric acid.
   (b) In the Haber process (Chapter 8), nitrogen reacts with hydrogen to produce ammonia gas.
   (c) At some time in the future, industry and consumers may make more extensive use of the combustion of hydrogen as an energy source.
   (d) When aqueous ammonia is added to an aqueous nickel(II) ion solution, the $Ni(NH_3)_6^{2+}(aq)$ complex ion is formed (**Figure 11**).

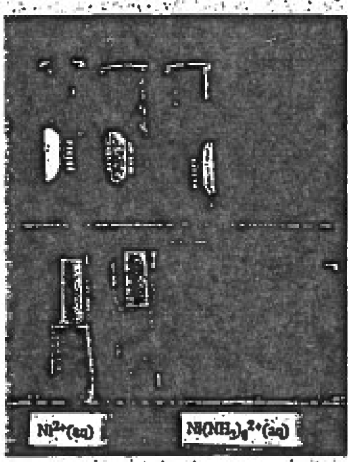

Figure 11
A $Ni^{2+}(aq)$ solution is green. Ammonia reacts with the nickel(II) ion to form the intensely blue hexaammine-nickel(II) ion, $Ni(NH_3)_6^{2+}(aq)$.

   (e) In the Solvay process for making washing soda (Chapter 7), one reaction involves heating solid calcium carbonate (limestone) to produce solid calcium oxide (quicklime) and carbon dioxide.
   (f) In Investigation 15.1, aqueous solutions of sodium sulfate and calcium chloride are mixed. (Remember to use a net ionic equation.)
   (g) In a sealed can of soda, carbonic acid, $H_2CO_3(aq)$, decomposes to liquid water and carbon dioxide gas.
2. You can apply the empirical and theoretical concepts of equilibrium to many different chemical reaction systems. Use the generalizations from your study of organic chemistry to predict the position of equilibrium for bromine placed in a reaction container with ethylene at a high temperature.
3. Interpret the graph in **Figure 12** to answer the questions about the reaction. Hydrogen and iodine were placed in a reaction vessel, which was then sealed, and heated to 450 °C.

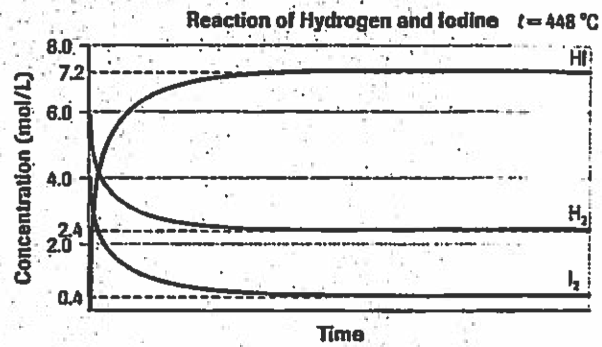

Figure 12
The progress of a hydrogen–iodine reaction

   (a) All three substances are gases. If the container has a volume of 2.00 L, what chemical amount of each substance was present initially?
   (b) What chemical amount of hydrogen iodide had formed at equilibrium? (Create an ICE table.)
   (c) Describe the rate at which hydrogen is reacting from the moment the reactants are mixed to the time when equilibrium has been established, in terms of collision–reaction theory.
4. For each of the following, write the chemical reaction equation with appropriate equilibrium arrow, as shown in **Table 3** (page 680).
   (a) The Haber process is used to manufacture ammonia fertilizer from hydrogen and nitrogen gases. Under less-than-desirable conditions, only an 11% yield of ammonia is obtained at equilibrium.
   (b) A mixture of carbon monoxide and hydrogen, known as water gas, is used as a supplementary fuel in many large industries. At high temperatures, the reaction of coke and steam forms an equilibrium mixture in which the products (carbon monoxide and hydrogen gases) are favoured. (Assume that coke is pure carbon.)
   (c) Because of the cost of silver, many high school science departments recover silver metal from waste solutions containing silver compounds or silver ions. A quantitative reaction of waste silver ion solutions with

copper metal results in the production of silver metal and copper(II) ions.

(d) One step in the industrial process used to manufacture sulfuric acid is the production of sulfur trioxide from sulfur dioxide and oxygen gases. Under certain conditions, the reaction produces a 65% yield of products.

5. Write the expression of the equilibrium law for the hydrogen–iodine–hydrogen iodide system at 448 °C. Using the evidence for System 1 as reported in Table 1 on page 679, calculate the value of the equilibrium constant.

6. In the Haber process for synthesizing ammonia gas from nitrogen and hydrogen, the value of K_c is 6.0×10^{-2} for the reaction at 500 °C. In a sealed container at equilibrium at 500 °C, the concentrations of $H_2(g)$ and of $N_2(g)$ are measured to be 0.50 mol/L and 1.50 mol/L, respectively. Write the equilibrium law expression and calculate the equilibrium concentration of $NH_3(g)$.

7. At a certain constant (very high) temperature, 1.00 mol of HBr(g) is introduced into a 2.00 L container. Decomposition of this gas to hydrogen and bromine gases quickly establishes an equilibrium, at which point the amount concentration of HBr(g) is measured to be 0.100 mol/L.
   (a) Write a balanced equation for the reaction.
   (b) Write the equilibrium law expression.
   (c) Calculate the chemical amount of HBr(g) present at equilibrium.
   (d) Calculate the chemical amount of HBr(g) that has reacted to form $H_2(g)$ and $Br_2(g)$ products when equilibrium is established.
   (e) Calculate the chemical amounts of $H_2(g)$ and $Br_2(g)$ that have been produced, and, thus, are present, when equilibrium is established.
   (f) Calculate the amount concentration of all substances present at equilibrium.
   (g) Calculate K_c for this reaction at this temperature.

8. To a heated reaction vessel with a volume of 1.00 L, a lab technician adds 6.23 mmol $H_2(g)$, 4.14 mmol of $I_2(g)$, and 22.40 mmol of HI(g). At equilibrium, a spectrophotometer is used to determine that the concentration of iodine vapour is 2.58 mmol/L. Construct an ICE table and find K_c for the reaction system
   $H_2(g) + I_2(g) \rightleftarrows 2\,HI(g)$.

9. Consider the system
   $CO_2(g) + H_2(g) \rightleftarrows CO(g) + H_2O(g)$
   Initially, 0.25 mol of water and 0.20 mol of carbon monoxide are placed in a 1.0 L reaction vessel. At equilibrium, spectroscopic evidence shows that 0.10 mol of carbon dioxide is present. Construct an ICE table and find K_c for this system.

10. Consider the system
    $2\,HBr(g) \rightleftarrows H_2(g) + Br_2(g)$
    Initially, 0.25 mol of hydrogen and 0.25 mol of bromine are placed into a 500 mL electrically heated reaction vessel. K_c for the reaction at the temperature used is 0.020.
    (a) Find the concentrations of the substances at equilibrium.
    (b) Calculate the chemical amount of each substance present at equilibrium.

11. Explain briefly how atomic theory, kinetic molecular theory, collision–reaction theory, and the concepts of reaction rate and reversible reactions are all necessary to explain chemical reaction equilibrium observations.

Extension

12. In a *very* long-term sense, Earth may be considered a closed system. One equilibrium of concern to scientists is the same one involved in carbonation of soft drinks, on a vastly larger scale. Scientists believe that over time, the carbon dioxide gas in the atmosphere should be in equilibrium with carbon dioxide dissolved in the oceans. They also know that the concentration of $CO_2(g)$ in the atmosphere has been increased significantly (by about 20%) in the last century, which, they believe, is mostly due to the burning of fossil fuels. Concerns about the consequences of global warming make it imperative that scientists improve their theories about the various cycles, processes, and equilibria involving this greenhouse gas. Research and summarize currently accepted theory about carbon dioxide dissolved in the oceans, and list some other cycles and systems involving reaction or production of $CO_2(g)$.

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1a. $1\,H_{2(g)} + 1\,Cl_{2(g)} \rightleftarrows 2\,HCl_{(g)}$ $\quad k_c = \dfrac{[HCl_{(g)}]^2}{[H_{2(g)}][Cl_{2(g)}]}$

b. $1\,N_{2(g)} + 3\,H_{2(g)} \rightleftharpoons 2\,NH_{3(g)}$ $\quad k_c = \dfrac{[NH_{3(g)}]^2}{[N_{2(g)}][H_2]^3}$

c. $2H_{2(g)} + 1O_{2(g)} \rightleftharpoons 2\,H_2O_{(g)}$ $\quad k_c = \dfrac{[H_2O_{(g)}]^2}{[H_{2(g)}]^2[O_{2(g)}]}$

d. $6\,NH_{3(aq)} + Ni^{2+}_{(aq)} \rightleftharpoons Ni(NH_3)^{2+}_{6(aq)}$ $\quad k_c = \dfrac{[Ni(NH_3)_6^{2+}]}{[NH_{3(aq)}]^6[Ni^{2+}_{(aq)}]}$

e. $CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$ $\quad k_c = [CO_2]$

f. $Na_2SO_{4(aq)} + CaCl_{2(aq)} \rightarrow 2NaCl_{(aq)} + CaSO_{4(s)}$

$Ca^{2+}_{(aq)} + SO_4{}^{2-}_{(aq)} \rightarrow CaSO_{4(s)}$ $\quad k_c = \dfrac{1}{[Ca^{2+}_{(aq)}][SO_4{}^{2-}_{(aq)}]}$

g. $H_2CO_{3(aq)} \longrightarrow H_2O_{(l)} + CO_{2(g)}$ $\quad k_c = \dfrac{[CO_{2(g)}]}{[H_2CO_{3(aq)}]}$

Rz

3a. $n_{HI} = cV$
$= (0.0\ mol/L)(2.00\ L)$
$= 0.0\ mol$

$n_{H_2} = cV$
$= (6.0\ mol/L)(2.00\ L)$
$= 12\ mol$

$n_{I_2} = cV$
$= (4.4\ mol/L)(2.00\ L)$
$= 8\ mol$

b. $n_{HI} = cV$ @ eq.
$= (7.2\ mol/L)(2.00\ L)$
$= 14.4\ mol$

c. H_2 is initially reacting quickly b/c there is lots of it present, but as it reacts there is less of it and the rate at which it reacts slows down b/c there are fewer collisions between H_2 + I_2 reacting

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4. a. $N_{2(g)} + 3H_{2(g)} \overset{11\%}{\rightleftharpoons} 2NH_{3(g)}$

b. $C_{(s)} + H_2O_{(g)} \overset{>50\%}{\rightleftharpoons} CO_{(g)} + H_{2(g)}$

c. $2Ag^+_{(aq)} + Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$

d. $2SO_{2(g)} + 1O_{2(g)} \overset{65\%}{\rightleftharpoons} 2SO_{3(g)}$

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6. $3H_{2(g)} + 1N_{2(g)} \rightleftharpoons 2NH_{3(g)}$

$$k_c = \frac{[NH_3]^2}{[H_2]^3[N_2]} \qquad 0.060 = \frac{[NH_3]^2}{(0.50)^3(1.50)}$$

$[NH_3] = 0.11\ mol/L$

7. $2HBr_{(g)} \rightleftharpoons H_{2(g)} + Br_{2(g)}$

	$2HBr$	H_2	Br_2
I	1.00 mol	0	0
C	− 0.800 mol (d)	+0.400 mol	+ 0.400 mol (e)
E	0.200 mol (e)	0.400 mol	0.400 mol

(b) $k_c = \dfrac{[Br_2][H_2]}{[HBr]^2}$

(c) $n = cV$
$= (0.100\ mol/L)(2.00\ L)$
$= 0.200\ mol$

(g) $k_c = \dfrac{(0.200\ mol/L)(0.200\ mol/L)}{(0.100\ mol/L)^2}$
$= 4.00$

(f) $C_{HBr} = \dfrac{0.200\ mol}{2L} = 0.100\ mol/L$

$C_{H_2} = \dfrac{0.400\ mol}{2.00\ L} = 0.200\ mol/L$ ← same for Br_2

8. $H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$

	H_2	I_2	$2HI$
I	6.23 mol/L	4.14 mol/L	22.40 mol/L
C	-1.56	-1.56	+3.12
E	4.67	2.58 mol/L	25.52

$$k_c = \frac{[HI]^2}{[H_2][I_2]} = 54.05\ldots$$

$= \boxed{54.1}$

9. $CO_{2(g)} + H_{2(g)} \rightleftharpoons CO_{(g)} + H_2O_{(g)}$

	CO_2	H_2	CO	H_2O
I	0	0	0.20 mol/L	0.25 mol/L
C	+0.10	+0.10	-0.10	-0.10
E	0.10 mol/L	0.10	0.10	0.15

$k_c =$