Chapter 16 & 17
Reaction Rates & Equilibrium Study Guide

Section 16.1 & 16.2 Reaction Rates and The Factors that affect Reaction Rates

Part A- Completion

Word Bank

~~Activation~~	~~Increasing~~	~~Rates~~	~~Temperature~~
~~Catalyst~~	~~Minimum~~	~~React~~	
~~Energy~~	~~Products~~	~~Slower~~	

Use the word bank above to check your understanding of the concepts and terms that are introduced in this section.

__1__ measure the speed of any change that occurs within a time interval. Collision theory states that particles __2__ when they collide, provided that they have enough __3__.

The rate at which a chemical reaction occurs is determined by an __4__ energy barrier. The activation energy is the __5__ energy that reactants must have to go to __6__. The higher the activation energy barrier, the __7__ the reaction. Chemists help reactants overcome the activation barrier in a number of ways. Two effective methods are to increase the __8__ at which the reaction is done or use a __9__. Rates of reaction can also be increased by __10__ the concentration of reactants.

1. Rates
2. React
3. Energy
4. activation
5. minimum
6. products
7. Slower
8. Temperature
9. Catalyst
10. Increasing

Part B- True-False

Classify each of these statements as always true (AT); sometimes true (ST); or never true (NT).

11. AT An increase in temperature will generally increase the rate of a reaction.
12. NT A catalyst is considered as a reactant in a chemical reaction.
13. AT The speed of a reaction can be increased by increasing reactant concentration or decreasing particle size.
14. AT An enzyme is a biological catalyst.

Part C- Matching

Match each description in Column B to the correct term in Column A

Column A	Column B
B 15. rate	a. synonym for an activated complex
D 16. collision theory	b. speed of a change that occurs over time
F 17. activation energy	c. substance that interferes with the action of a catalyst
A 18. transition state	d. Particles can react to form products when they collide, provided they have enough kinetic energy.
E 19. catalyst	e. substance that increases the rate of a reaction without being used up
C 20. inhibitor	f. minimum energy particles must have in order to react

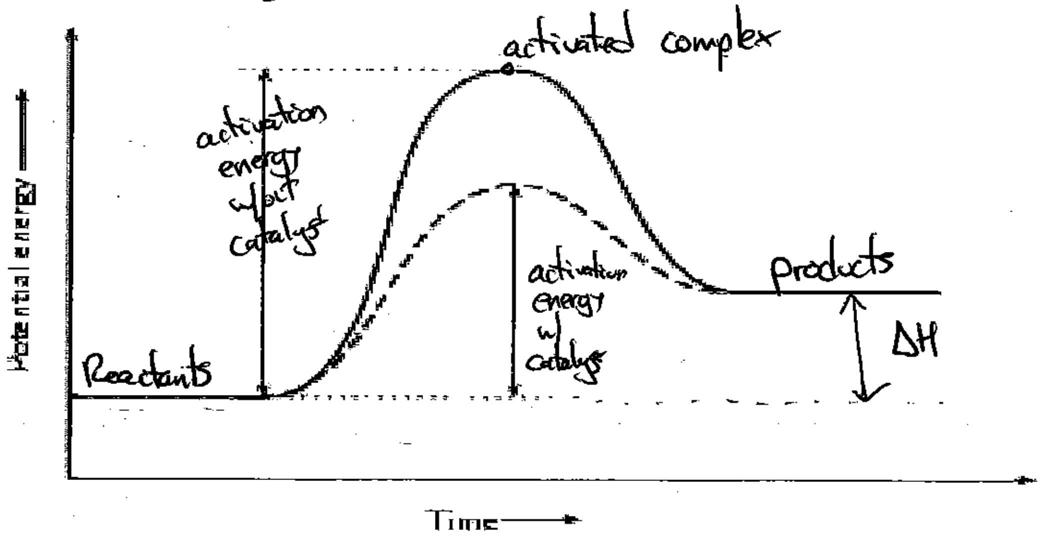
Part D- Questions & Problems

1. An ice machine can produce 120 kg of ice in 24 hours. Express the rate of ice production in kg/hr. In kg/day. In g/day. In kg/year.

$\frac{120\ kg}{24\ hr} = 5\ kg/hr$ $\frac{120\ kg}{day}$ $\frac{120{,}000\ g}{day}$ $\frac{120\ kg}{day}\left(\frac{365\ day}{yr}\right) = 43{,}800\ kg/yr$

2. Which of the following will increase the rate of a reaction?
   a. Increase particle size
   b. Increase temperature
   c. decrease concentration
   d. add a catalyst.

3. Label the following potential energy diagrams: (Label: reactants, products, activation energy without catalyst, activation energy with catalyst, activated complex, ΔH)



Is this reaction endothermic or exothermic? Explain.

endothermic - Energy Products is greater than Reactants

Section 17.1 Equilibrium: A State of Dynamic Balance

In your textbook, read about chemical equilibrium.

Complete each statement.

1. When a reaction results in almost complete conversion of reactants to products, chemists say the reaction goes to completion.
2. A reaction that can occur in both the forward and the reverse directions is called a(n) reversible.
3. Equilibrium is a state in which the forward and reverse reactions balance each other because they take place at equal rates.
4. At equilibrium, the concentrations of reactants and products are constant, but that does not mean that the amounts or concentrations are equal.
5. Equilibrium is a state of action, not one of inaction.

In your textbook, read about equilibrium expressions and constants.

For each statement below, write *true* or *false*.

True **6.** The law of chemical equilibrium states that at a given pressure, a chemical system may reach a state in which a particular ratio of reactant to product concentrations has a constant value.

True **7.** The equation $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ is an example of a homogeneous equilibrium.

True **8.** If an equilibrium constant has a value less than one, the reactants are favored at equilibrium.

False **9.** The value for K_{eq} is constant only at a specific volume.

False **10.** If the equilibrium constant for a reaction at 300 K is 49.7, the concentration of the reactants will be greater than the concentration of the products.

True **11.** A heterogeneous equilibrium means that reactants and products are present in more than one state.

True **12.** The product of the forward chemical reaction is HI, for the equilibrium expression:

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

With each of the following 4 equilibrium equations, write the equilibrium constant (K) expression for each. Also, determine if the equilibrium is heterogeneous or homogeneous.

a $4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g)$

$K = \frac{[NO]^4[H_2O]^6}{[NH_3]^4[O_2]^5}$

b $2NO_2(g) \rightleftharpoons N_2O_4(g)$

$K = \frac{[N_2O_4]}{[NO_2]^2}$

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

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

d $Ag^+(aq) + 2NH_3(aq) \rightleftharpoons Ag(NH_3)_2^+(aq)$

$K = \frac{[Ag(NH_3)_2^+]}{[Ag^+][NH_3]^2}$

Read the following table below. Based on the data in the table, put a star next to the reaction which contains the largest amount of product(s) at equilibrium. Put two stars next to the reaction which contains the largest amount of reactant(s) at equilibrium. Circle the number of the reaction(s) that have concentrations that represent the systems at equilibrium. For each system that is not at equilibrium, change the concentration of *only one* of the reactants or products do that the ratio represents the system at equilibrium.

Reaction	Concentrations	Equilibrium Constant (K_{eq})	
1. $2CH_4(g) \rightleftharpoons C_2H_2(g) + 3H_2(g)$	$[CH_4] = 0.500M$ $[C_2H_2] = 0.194M$ $[H_2] = 0.582M$ K=.153	0.153	$K = \frac{[C_2H_2][H_2]^3}{[CH_4]^2}$ = .153
2. $HCONH_2(g) \rightleftharpoons NH_3(g) + CO(g)$	$[HCONH_2] = 1.9 \times 10^{-2}M$ $[NH_3] = 0.30M$ $[CO] = 0.30M$ = 4.73	4.8	$K = \frac{[NH_3][CO]}{[HCONH_2]}$
3. $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ - change $[PCl_5]$ to .055M	$[PCl_5] = 0.30M$ $[PCl_3] = 0.45M$ $[Cl_2] = 0.22M$ = .33	1.8	$K = \frac{[PCl_3][Cl_2]}{[PCl_5]}$
4. $N_2O_4(g) \rightleftharpoons 2NO_2(g)$ ★★ Largest Amt of Reactants (K is smallest)	$[N_2O_4] = 0.754M$ $[NO_2] = 5.60 \times 10^{-2}M$ 4.16×10^{-3}	4.16×10^{-3}	$K = \frac{[NO_2]^2}{[N_2O_4]}$
5. $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$ ★ Largest Amt of Products (K is biggest)	$[H_2] = 0.110M$ $[I_2] = 0.500M$ $[HI] = 0.780M$ = 11.1	50.2	$K = \frac{[HI]^2}{[H_2][I_2]}$

change [HI] to 1.66 M

In your textbook, read about determining equilibrium constants.

A chemist did two experiments to determine the equilibrium constant for the reaction of sulfur dioxide with oxygen to form sulfur trioxide. Use the table showing the results of the experiments to answer the following questions.

$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$ at 873 K			
Experiment 1		**Experiment 2**	
Initial concentrations	Equilibrium concentrations	Initial concentrations	Equilibrium concentration
$[SO_2] = 2.00M$	$[SO_2] = 1.50M$	$[SO_2] = 0.500M$	$[SO_2] = 0.590M$
$[O_2] = 1.50M$	$[O_2] = 1.26M$	$[O_2] = 0M$	$[O_2] = 0.0450M$
$[SO_3] = 3.00M$	$[SO_3] = 3.50M$	$[SO_3] = 0.350M$	$[SO_3] = 0.260M$

13. Write the equation to calculate the equilibrium constant for the reaction.

$$K = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

14. Is this reaction an example of a homogeneous or heterogeneous equilibrium?

homogeneous (all gases)

15. Calculate the equilibrium constant from the data obtained in experiment 1.

$$K = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(3.5)^2}{(1.5)^2(1.26)} = 4.32$$

16. What is the equilibrium constant for the reaction in experiment 2?

$$K = \frac{[SO_3]^2}{[SO_2]^2[O_2]} = \frac{(.260)^2}{(.59)^2(.045)} = 4.32$$

17. Was it necessary to calculate both equilibrium constants? Why or why not?

No. K values are the same for the same reaction at the same temp - regardless of concentrations.

18. What does this experiment show about the initial concentrations of products and reactants in a reversible reaction?

The initial concentrations are irrelevant.

Section 17.2 Factors Affecting Chemical Equilibrium

In your textbook, read about Le Châtelier's Principle.

Answer the following questions.

1. What does Le Châtelier's Principle say?

When a stress is applied to a system in equilibrium, the system will "shift" to reduce the stress and maintain equilibrium.

2. What are three kinds of stresses that can be placed on a system?

(1) Addition or removal of reactants or products
(2) Addition or removal of heat (temp change)
(3) Addition or removal of pressure

For each reaction below, state the direction, left or right, in which the equilibrium will shift when the indicated substance is added. Identify one other way in which the reaction could be shifted in the same direction you indicated. (Hint: There may be more than one way to do this.)

3. Reaction: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$; NH_3 added

←

4. Reaction: $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$; H_2 added

→

5. Reaction: $CO(g) + H_2O \rightleftharpoons CO_2(g) + H_2(g)$; H_2O added

→

6. Reaction: $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$; SO_3 added

←

7. Reaction: $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$; SO_2 added

→

8. Reaction: $2NCl_3(g) \rightleftharpoons N_2(g) + 3Cl_2(g)$; NCl_3 added

→

For each reaction below, indicate in which direction the equilibrium shifts when the stated stress is applied to the system. Write *R* if the reaction shifts to the right, *L* if it shifts to the left, or *NC* if there is no change.

	Reaction	Stress
L	**20.** $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g) + heat$	temperature increase
NC	**21.** $CO(g) + Fe_3O_4(s) \rightleftharpoons CO_2(g) + 3FeO(s)$ (1) (1)	volume increase (P↓) - ~~less~~ more moles of G
R	**22.** $C_2H_2(g) + H_2O(g) \rightleftharpoons CH_3CHO(g) + heat$	temperature decrease
R	**23.** $2NO(g) + H_2(g) \rightleftharpoons N_2O(g) + H_2O(g) + heat$	volume decrease (P↑) less moles of G
L	**24.** $Heat + H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$	temperature decrease
NC	**25.** $H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g) + heat$	volume decrease (P↑) less moles of G

Summing up Le Chatelier's Principle

Changing Concentration

Add more reactants	**R → P**
Add more products	**R ← P**
Remove reactants	**R ← P**
Remove products	**R → P**

Changing Temperature

Increasing Temperature/ Adding Heat (endothermic reaction)	**R → P**
Increasing Temperature/ Adding Heat (exothermic reaction)	**R ← P**
Decreasing Temperature/ Removing Heat (endothermic rxn)	**R ← P**
Decreasing Temperature/ Removing Heat (exothermic rxn)	**R → P**

Changing Pressure

Generally, as the pressure increases, the reaction will shift in the direction of the least amount of gases. As the pressure decreases, the reaction will shift in the direction of the most amount of gases. **LOOK** at the number of moles of gases in both reactants and products.

For example: formation of water from hydrogen and oxygen: $2H_2 + O_2 \leftrightarrow 2H_2O$

3 moles **2 moles**

At high pressure: R → P (less moles of gas)
At low pressure: R ← P (more moles of gas)

Example: The Haber process (The formation of ammonia, an exothermic reaction)

$$\mathbf{N_2(g) + 3H_2(g) \leftrightarrow 2NH_3(g)} + \text{heat}$$

a. What would be the effect of raising the temperature on this reaction and why?

← raising temp will go away from heat

b. What would be the effect of increasing the pressure and why?

P↑ (less moles of Gas) →

c. What would happen if you decrease the concentration of NH_3?

→

d. What would happen if you add more N_2?

→

UNIT 5 — ASSIGNMENT 98

For these questions predict what changes will take place resulting from the changes in concentration, pressure, or heat. Le Chatelier's Principle will be helpful in making these predictions.

Le Chatelier's Principle: If a system at equilibrium is changed, the system will partially counteract that change.

Predict the result of these changes on the following equilibrium system(s):

1. Cu^{2+} (light blue) + $4NH_3$ (colorless) $\rightleftharpoons$ $Cu(NH_3)_4^{2+}$ (dark blue)
   a. Add Cu^{2+} solution $\rightarrow$
   b. Add NH_3 solution $\rightarrow$
   c. Remove NH_3 $\leftarrow$
   d. Add $Cu(NH_3)_4^{2+}$ $\leftarrow$
2. $2CO + O_2 \rightleftharpoons 2CO_2$ + heat
   a. Add CO_2 $\leftarrow$
   b. Remove O_2 $\leftarrow$
   c. Add CO $\rightarrow$
   d. Add heat $\leftarrow$
   e. Remove heat $\rightarrow$
3. $N_2 + 3H_2 \rightleftharpoons 2NH_3$ + heat
   a. Remove NH_3 $\rightarrow$
   b. Add H_2 $\rightarrow$
   c. Remove N_2 $\leftarrow$
   d. Add heat $\leftarrow$
   e. Remove heat (cool) $\rightarrow$
4. $Cr_2O_7^{2-}$ (orange) + OH^- (colorless) $\rightleftharpoons$ $2CrO_4^{2-}$ (yellow) + H^+ (colorless)
   a. Add H^+ $\leftarrow$
   b. Remove OH $\leftarrow$
   c. Add Ba^{2+} (reacts with CrO_4^{2-} but not with $Cr_2O_7^{2-}$) $\rightarrow$
   d. Remove $Cr_2O_7^{2-}$ $\leftarrow$