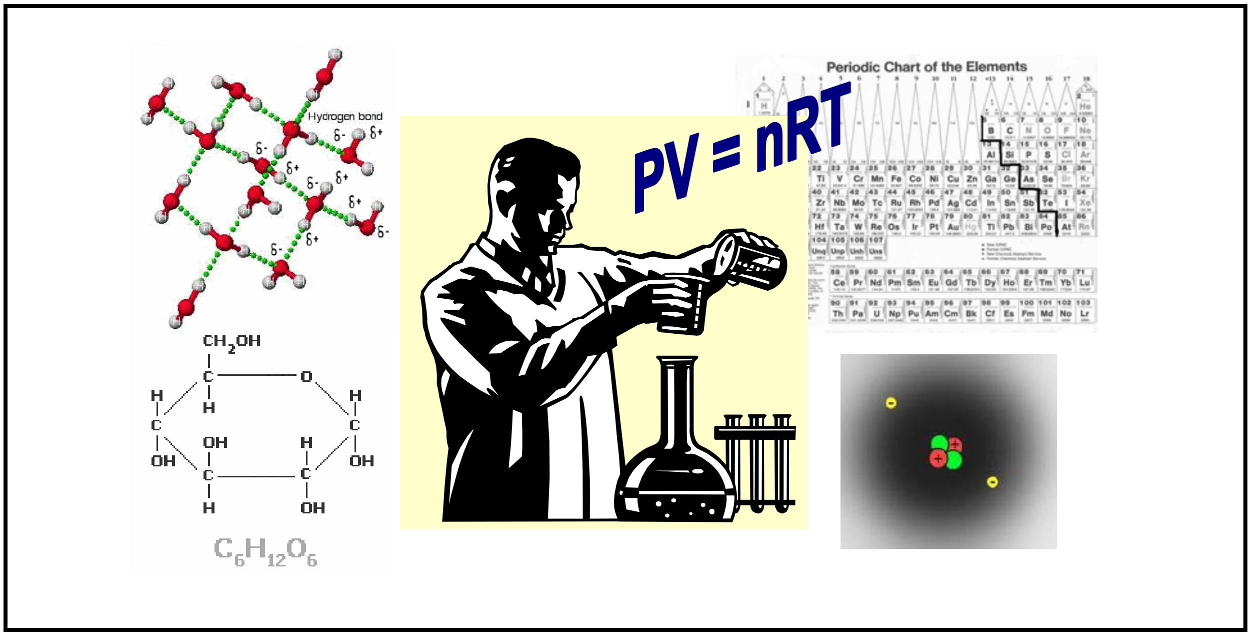
Project EASE

(Effective Alternative Secondary Education)

CHEMISTRY



MODULE 17

Reaction Rates and Equilibrium

BUREAU OF SECONDARY EDUCATION
Department of Education
DepEd Complex, Meralco Avenue
Pasig City

Module 17
Reaction Rates and Equilibrium

What this module is about

This module is all about the speed of chemical reactions and chemical equilibrium. Some reactions are very fast like the burning of wood while others are very slow, like the rusting of iron. Some reactions go to completion, that is, all the reactants are converted to products. However, upon forming the product some reactions will revert to reactants because of an equilibrium that exist. Understanding the factors affecting the rate or speed of chemical reactions and the chemical equilibrium will help you appreciate the chemical changes happening around you.

To make the discussion easy for you, the module is divided into three (3) lessons:

- **Lesson 1 – Factors Affecting Rates of Reaction**
- **Lesson 2 – Factors Affecting Chemical Equilibrium**
- **Lesson 3 – Applications of Chemical Equilibrium**

What you are expected to learn

After going through this module, you should be able to:

1. define and differentiate exothermic and endothermic reactions.
2. describe the factors affecting the rate of reactions.
3. explain the Le Chatelier's principle.
4. write an equilibrium expression.

How to learn from this module

Here are some pointers to remember as you go over this module.

1. Read and follow the instructions carefully.
2. Answer the pretest before reading the content of the module.

3. Take down notes and record points for clarification.
4. Always aim to get at least 70% of the total number of items given.
5. Be sure to answer the posttest at the end of the module.

What to do before (Pretest)

Take the pretest before proceeding to the lessons. Check your answers by referring to the answer key on the last page of the module.

I. Write **TRUE** if the statement is correct, otherwise write **FALSE.**

_______________ 1. The speed of chemical reactions can be altered by adding a catalyst.
_______________ 2. When a chemical reaction gives off heat, it is an exothermic reaction.
_______________ 3. The speed of chemical reaction increases as the particle size increases.
_______________ 4. When chemical equilibrium is achieved, there is no change in the concentration of reactants and products.
_______________ 5. Heating causes more products to be formed for an endothermic reaction.

II. Write whether the direction of the reaction will go **FORWARD, REVERSE** or **NO CHANGE**.

$$2\ A_{(g)} + B_{(g)} \leftrightarrow Q_{(g)} + \text{heat}$$

_______________ 1. Pressure increased
_______________ 2. Volume decreased
_______________ 3. 2A removed
_______________ 4. B added
_______________ 5. System was cooled.
_______________ 6. Catalyst was added.
_______________ 7. Q was collected.
_______________ 8. Both B and Q were removed; same amount.
_______________ 9. Temperature was increased.
_______________ 10. Both B and Q were added; same amount.

Key to answers on page 14.

Lesson 1. Factors Affecting Reaction Rates

When you burn wood, heat is produced. We use this heat to cook our food or boil water. The chemical reaction in this case is the burning of wood and since heat is given off, we say the reaction is **EXOTHERMIC**. So we define an exothermic reaction as a reaction where heat is evolved or given off.

When you mix urea and water, the container of the mixture feels cold. This is an example of **ENDOTHERMIC** reaction, a reaction wherein heat is absorbed. To represent an exothermic reaction, the value of change in enthalpy (ΔH) is negative while for the endothermic reaction, the value is positive. **Enthalpy** refers to the heat content measured at constant pressure.

Exothermic reaction: Reactants $\rightarrow$ Products

$$C + O_2 \rightarrow CO_2 \qquad \Delta H = -393.5 \text{ kJ/mol}$$

Endothermic reaction: Reactants $\rightarrow$ Products

$$N_2 + O_2 \rightarrow 2NO \qquad \Delta H = +180 \text{ kJ/mol}$$

Aside from being exothermic or endothermic, chemical reactions taking place around you are also sometimes very fast or very slow. Burning is a very fast chemical reaction while the rusting of iron is slow. If a certain chemical reaction is desirable or useful like the production of drugs to cure diseases, we want the chemical reaction to happen fast so we can produce more in a short period of time. Is there anything we can do to alter the rate or speed of chemical reaction?

Did you know that firecracker explosion is an example of an exothermic reaction…a very fast exothermic reaction!

Turn the page and do Activity 1.1 to observe the rate of dissolution of sugar.

What you will do

Activity 1.1 The Rate of Dissolving Rock and Powder Sugar

Materials:

Rock sugar	Two (2) drinking glasses (same size)
Powder sugar	Spoon
Water	Watch

Procedure:

1. Pour water up to the same level on two drinking glasses.

2. Take a pinch of sugar and drop it in one of the glasses. Do not stir. After one minute, stir the glass slowly until no sugar can be seen. Record the time it takes for the sugar to be dissolved completely.

3. Repeat procedure no. 2, this time use rock sugar. The size of your rock sugar should be that of a mongo bean.

Analysis:

How much time did it take the powder sugar to dissolve completely? How about the rock sugar? Why is this so?

The speed or rate of chemical reactions is altered or changed by the following factors:

1. Nature and size of the reactants
2. Concentration of the reactants
3. Temperature
4. Presence of Catalyst

Nature and Size of the Reactants

The activity demonstrated that particle size can increase or decrease the rate of reaction. Though no chemical reaction has taken place, the rate of dissolution of sugar is faster for the powder sugar than the rock sugar. Similarly, the wood we use to cook our food is chopped to a desirable size so it will burn easily. If we do not chop the wood, it will take more time for the wood to burn.

You also observe that some metals do not form rust while others do. Rusting is a chemical reaction wherein the metal reacts with oxygen present in the air. The rust formed is the product of that reaction. Gold does not rust but iron forms rust. Though both are metals, gold does not react with oxygen while iron does.

Some elements have different physical forms. An example is the crystalline form of diamond and graphite. Though the diamond and graphite are two different forms, both are made up of the same element - carbon. Carbon exhibits **ALLOTROPHY.** Another example of an element exhibiting allotrophy is phosphorus. One is color white and the other is red. Between the two, white is more reactive. It burns readily when exposed to air while the red phosphorus does not. Red phosphorus is less dangerous and can be stored for long periods of time without exploding. However, to store white phosphorus, it has to be submerged in water. Even though both are made up of the same element, their chemical behaviors are different.

Concentration: The Law of Mass Action

The rate of the reaction is also observed to be proportional to the concentration of the reacting system. This simply means that the greater the number of reacting components present in a container, the more reaction will occur in a given time. An example of this is the concentration of acids. If you have one (1) molar hydrochloric acid (HCl) and you drop a metal zinc, the gas produced from the reaction is slow. However, if you have a six (6) molar HCl, the reaction is very fast and you can see lots of bubbles forming. Six (6) molar HCl is corrosive as well. Great care has to be observed in handling this acid.

Temperature

What is the reason why foods stored in the refrigerator stay longer? The answer is temperature! Temperature can either increase or decrease the rate of reaction. The spoiling of food indicates that an undesirable chemical reaction has taken place. During summer, foods are easily spoiled. To avoid spoilage, these are stored inside the refrigerator.

In the same manner, desirable chemical reactions have to be carried out swiftly and to achieve this, temperature is elevated. Commercially, reaction mixtures are heated and then mixed. Extreme caution is required if the reaction is to be conducted at high temperature. However, some of the reactants are destroyed at high temperature. Instead of obtaining the desired valuable product, none is collected because the reactant has already decomposed to some worthless compound. What then can be done?

Presence of Catalyst

If the desired reaction has to be carried out swiftly, without elevating the temperature too much, **CATALYSTS** are used. Catalysts are substances that alter the rate of chemical reaction without being consumed. It simply means that catalysts take part in the reaction but after the reaction has been completed, they can be recovered and used again.

Catalysts alter the rate of reaction and therefore, these can increase or decrease the rate of a particular reaction. When a catalyst increases the rate of the chemical reaction, it is said to be a **POSITIVE CATALYST** or simply named **catalyst**. If it slows down the reaction it is called a **NEGATIVE CATALYST** or it is commonly called an **INHIBITOR**. How do catalysts accomplish this altering action? They do so by providing an alternative step or path for the reaction to occur; they make short-cuts to reach their destination.

Catalysts can be of metals, ions, acids or bases. Inside our body, we have catalysts. These are called **ENZYMES.** Enzymes do amazing tasks like converting the food we eat to energy, and maintaining and repairing our cells.

As an application of what we have learned, let us take a look at the formation of ammonia, NH_3. Ammonia is a very important industrial chemical and has many uses, one of which is in the manufacture of fertilizers. Ammonia is formed from the reaction of hydrogen gas, H_2, and nitrogen gas, N_2. The reaction is written this way:

Hydrogen gas + Nitrogen gas $\rightarrow$ Ammonia gas

$$3H_{2\,(g)} + N_{2\,(g)} \rightarrow 2NH_{3\,(g)}$$

The reaction is carried out at a higher pressure and at a slightly elevated temperature. Yet, it is still slow. The temperature cannot be elevated too much due to limitations imposed by chemical equilibrium. To overcome this problem, a catalyst is used. We will study the detail of this reaction when we go to chemical equilibrium.

What you will do

Self-Test 1.1

I. Classify whether the reaction is endothermic or exothermic.

_________________ 1. Ice making
_________________ 2. $H_2O_{(l)} \rightarrow H_2O_{(g)}$
_________________ 3. $H_2O_{(l)} \rightarrow H_2O_{(s)}$
_________________ 4. $CH_3OH_{(g)} \rightarrow CO_{(g)} + 2H_{2(g)}$ $\Delta H = 90.2$ KJ/mol
_________________ 5. $CH_2O_{(g)} + H_{2(g)} \rightarrow CH_3OH_{(g)}$ $\Delta H = -92.2$ KJ/mol

II. Indicate the factors that affect the rate of chemical reactions.

_________________ 1. Rusting of iron but not gold
_________________ 2. Spoiling of food at a warmer temperature
_________________ 3. Fermentation of sugars to alcohol using yeast
_________________ 4. Copper oxide speeds up the decomposition of H_2O_2 (hydrogen peroxide)
_________________ 5. Grinding of calcium carbonate before calcining

Key to answers on page 15.

Lesson 2. Factors Affecting Chemical Equilibrium

You can observe many fascinating events in a circus like a tiger jumping through a ring of fire or clowns juggling balls. The most daring act, however, is when a performer walks on a rope that is elevated several feet above the ground, crossing from one end to the other. You can see the rigid training he underwent just to perform this balancing act.

In chemistry, a balancing act also occurs. This balancing act is what we call chemical equilibrium. It is a dynamic state, not a static one. It is a state achieved when the rate of forming the product is the same as the rate of reverting the product back to its reactants.

Forward reaction: REACTANTS $\rightarrow$ PRODUCTS
Reverse reaction: PRODUCTS $\rightarrow$ REACTANTS

We simply represent this as

$$\text{REACTANTS} \leftrightarrow \text{PRODUCTS}$$

The important concept here is that when the reactants are converted to products, the products will revert to reactants. The speed of the forward reaction or product formation, is the same as the speed of the reverse reaction or reactant formation. If this is the case, how much reactants and products are there? Will the concentration of the reactants and product change after a long period of time since reverse reaction is happening?

The answer is that there is no tendency for the concentration of the reactants and products to change once they reached equilibrium. There is a particular numerical value for the ratio of the products and reactants upon reaching equilibrium and this is the **EQUILIBRIUM CONSTANT**, Kc, a characteristic property of the reaction system. The equilibrium constant is expressed as:

$$Kc = \frac{[\text{Products}]}{[\text{Reactants}]}$$

The products referred to is the concentration of products at equilibrium, raised to its stoichiometric coefficient. Likewise the reactant is the concentration of the reactants at equilibrium raised to its stoichiometric coefficient. As an illustration, we write the equilibrium expression for the given reaction as

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

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

In writing the equilibrium expression for a particular reaction, only **GASES** and **AQUEOUS SOLUTION** (aq) are included. The solid and liquid states for both the reactants and products are **NOT INCLUDED**.

So, if a certain chemical reaction is given, and you were asked to write its Kc, you present it as the ratio of products and reactants raised to its respective stoichiometric coefficient. An example is the decomposition of calcium carbonate to calcium oxide and carbon dioxide. Take note of the state of the reactants and products.

$$CaCO_{3\,(s)} \leftrightarrow CaO_{(s)} + CO_{2\,(g)}$$

We write the expression as

$$Kc = [CO_2]$$

Le Chatelier's Principle

If our chemical system is already at equilibrium, can we disturb it? Suppose we want to collect more of the products rather than just allow the products to go back as reactants, is there something we can do to shift the reaction towards the formation of products? The answer is yes and this is based on Le Chatelier's principle.

Le Chatelier, a french chemist, observed that when a system was already at equilibrium and a disturbance or stress was applied, the system reacted in such a way that stress was relieved and an equilibrium state was achieved. Picture it as a see-saw in a balanced position. If you add weight, the see-saw will adjust the added weight so that the balanced position can be achieved. If we can disturb the equilibrium thereby favoring the reaction we want, what are we going to adjust? What are the factors that can affect equilibrium? There are three things that we can do:

1. We add or remove one or more of the substances.

   The direction of the reaction will move to where the disturbance or stress is applied. Where will it be moved, to the left or to the right? Consider the reaction below:

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

   If we add more of either one of the substances in the reactant side, we are disturbing the equilibrium. There is now an increase in the concentration of the reactants and this disturbed the equilibrium. To relieve this disturbance or stress, the direction of the reaction will move to the right, forming more products. The change in the direction of the reaction therefore relieves the stress on the reactant side.

   If we remove HI in the product, this will favor the forward reaction. HI concentration has now decreased. To attain equilibrium, add HI, then the direction of the reaction will shift backward.

2. Increasing the pressure favors the lesser number of moles.

   For the same reaction, we see that the total number of moles of the reactants and product are the same. Pressure has no effect on this reaction system.

   $$\underset{\text{1mole}}{H_{2\,(g)}} + \underset{\text{1mole}}{I_{2\,(g)}} \leftrightarrow \underset{\text{2moles}}{2HI_{(g)}}$$

   However if we consider the formation of ammonia, there is now a difference in the total number of moles of the reactants and product. Previously, the reaction was written in one direction only, but this time it is a reversible reaction. If we increase the pressure, the direction will go to the right, its total number of moles is only two (2) while there are four (4) on the left.

$$\begin{array}{ccccc} 3H_{2\,(g)} & + & N_{2\,(g)} & \leftrightarrow & 2NH_{3\,(g)} \\ \text{3moles} & & \text{1mole} & & \text{2moles} \end{array}$$

The relation of pressure and volume are inversely proportional. If we increase the pressure, the volume will decrease. If volume is increased, pressure will decrease and this will shift the direction of the reaction to the left.

3. Temperature

Picture the exothermic reaction wherein the heat produced is part of the product while for an endothermic reaction, heat is one of the reactants. For an exothermic reaction, increasing the temperature will direct the reaction towards the left. More reactants will be formed instead of the product. An endothermic reaction will then favor the production of more reactants if we elevate the temperature.

Adding some substances that are not included in the reaction system will have no effect on the equilibrium, That is, these cannot affect or change the direction of the reaction despite their presence. As long as the substance being added does not react with either one of the reactants or products, there is no effect. A catalyst also has no effect since both the forward and reverse reactions are altered. It only hastens the attainment of equilibrium.

What you will do

Activity 2.1

From the given reaction $4NH_{3(g)} + 5O_{2(g)} \rightleftarrows 4NO_{(g)} + 6H_2O_{(g)}$ $\Delta H = 950$ KJ. Write **True** if statement is correct, otherwise write **False**.

_______________1. The reaction is exothermic.
_______________2. Heating will cause forward reaction to occur.
_______________3. Pressure has no effect.
_______________4. The reaction will go forward if volume is increased.
_______________5. Adding ammonia will cause forward reaction to occur.

Key to answers on page 15.

What you will do
Self-Test 2.1

Try to check how much you have learned in this lesson by answering the following questions.

I. Write the equilibrium expression for the following reactions

1. $Ni_{(s)} + 4CO_{(g)} \leftrightarrow Ni(CO)_{4(g)}$
2. $H_2O_{(l)} \leftrightarrow H_2O_{(g)}$
3. $NiO_{(s)} + CO_{(g)} \leftrightarrow CO_{2(g)} + Ni_{(s)}$
4. $H_{2(g)} + I_{2(g)} \leftrightarrow 2HI_{(g)}$
5. $3H_{2(g)} + N_{2(g)} \leftrightarrow 2NH_{3(g)}$

II. Consider the exothermic reaction

$$2SO_{2(g)} + O_{2(g)} \leftrightarrow 2SO_{3(g)}$$

Write $\rightarrow$ if the reaction will go forward, $\leftarrow$ for reverse and $\leftrightarrow$ for no change in the direction of the reaction.

________________ 1. Adding oxygen
________________ 2. Removing sulfur dioxide
________________ 3. Removing sulfur trioxide
________________ 4. Adding helium gas
________________ 5. Adding catalyst

Key to answers on page 15.

Lesson 3. Applications of Chemical Equilibrium

In this lesson, we will study some of the wonderful applications of chemical equilibrium. The principles of chemical equilibrium are evident from the production of industrial chemicals, the acid-balance in lakes and streams, to the maintenance of our blood pH.

Haber Process: Ammonia synthesis

We have encountered the synthesis of ammonia from previous lessons. This time, we will take a closer look at this reaction. The complete equation is written as

$$3H_{2(g)} + N_{2(g)} \leftrightarrow 2NH_{3(g)} \qquad \Delta H = -92 \text{ kJ/mol}$$

To produce ammonia economically, we have to run the reaction in such a way that more ammonia can be obtained in a short period of time. To accomplish this, a catalyst is used to reach equilibrium faster and temperature is elevated to increase the rate of reaction.

However, since the reaction is exothermic, elevating the temperature too much will go in the reverse direction. To balance the need to accelerate the reaction and lessen the formation of reactants, very high pressure is employed - about 300atm, and the temperature is slightly increased to about 500^oC. The catalyst used is usually iron oxide. Ammonia is collected periodically and the unreacted gases are recycled to the reactor.

Acid-Balance in Lakes

Unpolluted rainwater has a pH of about 5.6. This is due to the carbon dioxide, CO_2 present in air. The reactions are given below.

$$CO_{2\,g)} + H_2O_{(g)} \leftrightarrow H_2CO_{3(aq)}$$
$$H_2CO_{3(aq)} \leftrightarrow H^+_{(aq)} + HCO_3{}^-_{(aq)}$$

The presence of dissolved CO_2 makes our rain slightly acidic. Does this mean that the acid rain causing environmental problems is due to carbon dioxide? The answer is no. Acid rain is caused by sulfur dioxide, SO_2, produced from the combustion of coal.

The use of coal as fuel has posted some problems due to the production of sulfur dioxide. Sulfur content of coal ranges from 3% to 5%. Combustion of coal and the sulfur contained therein is given below

Combustion of coal: $C_xH_y + O_2 \rightarrow XCO_2 + YH_2O$

where x represents the number of carbon atoms
y represents the number of hydrogen atoms

Combustion of sulfur: $S + O_2 \leftrightarrow H^+_{(aq)} + SO_2$

Similarly for SO_2,

$$SO_{2(g)} + H_2O_{(g)} \leftrightarrow H_2SO_{3(aq)}$$
$$H_2SO_{3(aq)} \leftrightarrow H^+_{(aq)} + HSO_3{}^-_{(aq)}$$

H_2SO_3 is a stronger acid than H_2CO_3. About 0.12 ppmv (parts per million by volume) of SO_2 present in air will produce a pH of 4.30 in water whereas 350 ppmv of CO_2 will have a pH of 5.6.

In lakes surrounded by rocks, acid rain is counteracted. The acid-balance is provided by the carbonates present in rocks. Limestone and rocks contain calcium carbonate, $CaCO_3$. We can represent the reaction as

$$CaCO_{3(s)} + 2H^+_{(aq)} \leftrightarrow Ca^+_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

Maintenance of Blood pH

The pH of blood is between 7.35 and 7.45. A value beyond this range indicates illness and death may occur. Blood has many functions. One of its vital roles is to supply oxygen gas to vital organs and tissues and transport carbon dioxide from tissues and organs back to the lungs. Nutrients are likewise delivered by the blood and waste products are carried out in return.

If the pH of blood is below 7.35, the condition is called **ACIDOSIS.** If the pH of the blood is higher than 7.45, the condition is called **ALKALOSIS**. The protection of lake from acid rain is very much similar to what is happening in the blood. The part of the complex reaction is:

$$CO_{2(g)} + H_2O_{(g)} \leftrightarrow H_2CO_{3(aq)} \qquad \textit{reaction 1}$$
$$H_2CO_{3(aq)} \leftrightarrow H^+_{(aq)} + HCO_3^-{}_{(aq)} \qquad \textit{reaction 2}$$

Little H_2CO_3 is formed so we can write the reaction simply as

$$CO_{2\,(g)} + H_2O_{\,(g)} \leftrightarrow H^+_{(aq)} + HCO_3^-{}_{(aq)} \qquad \textit{reaction 3}$$

Carbon dioxide enhances the release of oxygen gas from the blood and the HCO_3^- produced is carried to the lungs. In the lungs, the reverse of reaction 2 occurs and it is followed by the reverse of reaction 1. Carbon dioxide is then exhaled out of the lungs.

Take a look at *reaction 3*. If carbon dioxide cannot escape from the blood, higher concentration of CO_2 shifts the reaction towards the right, therefore more H+ is produced making the blood acidic. Pneumonia is one sickness that causes acidosis. Breathing is impaired, which leads to build up of CO_2.

What you will do
Activity 3.1

Choose the correct answer.

________________1. ($CaCO_3$, SO_2) counteracts acids in lakes.
________________2. If the pH of blood is 7.45, the condition is called (acidosis, alkalosis).
________________3. Blood supplies (O_2, CO_2) to vital organs and tissues.
________________4. (H_2SO_3, H_2CO_3) is a stronger acid.
________________5. Pneumonia causes (acidosis, alkalosis).

Key to answers on page 15.

What you will do

Self-Test 3.1

The following statements are false. Change the underlined word to make it true.

________________1. In combustion reactions, water is the reactant.
________________2. Ammonia synthesis is called Ostuald Process.
________________3. Copper oxide is used as catalyst in ammonia synthesis.
________________4. Carbon dioxide makes the raid acidic.
________________5. High concentration of O_2 causes acidosis.

Key to answers on page 15.

Let's Summarize

1. **Exothermic reaction** is the type of reaction wherein heat is given off.
2. **Endothermic reaction** is the type of reaction wherein heat is absorbed.
3. **Law of mass action** states that the rate of reaction is directly proportional to the concentration of the reacting system.
4. **Catalysts** are substances that alter the rate of reaction.
5. **Positive catalysts** increase the rate of the reaction.
6. **Inhibitor** or negative catalysts decrease the rate of the reaction.
7. **Enzymes** are biological catalysts inside our body.
8. **Chemical equilibrium** is a state achieved when the rate of the forward and reverse reaction are equal.
9. **Le Chatelier** states that when a system at equilibrium has been disturbed or stress has been applied, the system will respond in such a way to relieve the disturbance or the stress.
10. **Equilibrium constant, Kc,** is the characteristic property of the reacting system. It is written as the ratio of equilibrium concentrations of products and reactants raised to its respective stoichiometric coefficients.
11. **Acidosis** is a condition wherein the pH of blood is below 7.35.
12. **Alkalosis** is a condition wherein the pH of blood is above 7.45.

Posttest

I. Write **TRUE** if the statement is correct, otherwise write **FALSE**.

_______________ 1. Catalysts are used up during chemical reaction.
_______________ 2. Catalysts have an effect on the direction of the reaction when equilibrium is reached.
_______________ 3. Acidosis is due to high concentration of CO_2 in the blood.
_______________ 4. Catalyst increases the rate of chemical reaction.
_______________ 5. Burning is an exothermic reaction.

II. For the reaction $3H_{2(g)} + N_{2(g)} \leftrightarrow 2NH_{3(g)}$ $\Delta H = -92$ kJ/mol
Cite five (5) ways on how you will increase the formation of ammonia

1. ____________________
2. ____________________
3. ____________________
4. ____________________
5. ____________________

III. Write the equilibrium expression for the following reactions

_______________ 1. $2N_2O_{(g)} + 3O_{2(g)} \leftrightarrow 2N_2O_{4(g)}$
_______________ 2. $CO_{2(g)} \leftrightarrow CO_{2(s)}$
_______________ 3. $N_{2(g)} + O_{2(g)} \leftrightarrow 2NO_{(g)}$
_______________ 4. $2Cr^{3+}_{(aq)} + Fe_{(s)} \leftrightarrow 2Cr^{2+}_{(aq)} + Fe^{2+}_{(aq)}$
_______________ 5. $2NaHCO_{3(s)} \leftrightarrow Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{(g)}$

Key to answers on page 16.

Key to Answers

Pretest

I. TRUE or FALSE

1. TRUE
2. TRUE
3. TRUE
4. TRUE
5. TRUE

II. FORWARD, REVERSE or NO CHANGE

1. FORWARD
2. FORWARD
3. REVERSE
4. FORWARD
5. FORWARD
6. NO CHANGE
7. FORWARD
8. NO CHANGE
9. NO CHANGE
10. NO CHANGE

Lesson 1

Self-Test 1.1

I.
1. Endothermic
2. Exothermic
3. Endothermic
4. Endothermic
5. Exothermic

II.
1. Nature of reactants
2. Temperature
3. Catalyst
4. Catalyst
5. Size of reactants

Lesson 2

Activity 2.1

1. True
2. False
3. False
4. True
5. True

Self-Test 2.1

I. Write the equilibrium expression

1. $Kc = \dfrac{[Ni(CO)_4]}{[CO_2]}$
2. $Kc = [H_2O]$
3. $Kc = \dfrac{[CO_2]}{[CO]}$
4. $Kc = \dfrac{[HI]^2}{[H_2][I_2]}$
5. $Kc = \dfrac{[NH_3]^2}{[H_2]^3[N_2]}$

II. Forward (→), reverse (←) or no change (↔)

1. →
2. ←
3. →
4. ↔
5. ↔

Lesson 3

Activity 3.1

1. $CaCO_3$
2. Alkalosis
3. O_2
4. H_2SO_3
5. Acidosis

Self-Test 3.1

1. Oxygen
2. Haber
3. Iron oxide
4. SO_2
5. CO_2

Posttest

I. TRUE or FALSE

1. FALSE
2. FALSE
3. TRUE
4. TRUE
5. TRUE

II. Five (5) ways to increase formation of NH_3

1. Add H_2
2. Add N_2
3. Remove NH_3
4. Increase pressure/decrease volume
5. Decrease temperature

III. Writing equilibrium expression

1. $Kc = \dfrac{[N_2O_4]^2}{[N_2O]^2\,[O_2]^3}$
2. $Kc = \dfrac{1}{[CO_2]}$
3. $Kc = \dfrac{[NO]^2}{[CO]}$
4. $Kc = \dfrac{[Cr^{2+}]^2\,[Fe^{2+}]}{[Cr^{3+}]^2}$
5. $Kc = [CO_2]\,[H_2O]$

References

Petrucci, R. (1982). *General chemistry: Principles and applications*. (3^{rd} ed.) USA: McMillan Publishing Co.

Bunce, N. (1993). *Introduction to environmental chemistry*. Canada: Wuerz Publishing Ltd.

Bettleheim, F.A., Brown, W.H. & March, J. (2004). *Introduction to organic and biochemistry*. (5^{th} ed.) USA: Brooks/Cole Publishing Co.

Denniston, K.J., Topping, J.J. & Caret, R.L. (2004). *General, organic and biochemistry: Principles and applications*. (4^{th} ed.) USA: McGrawHill Publishing Co.

Manahan, S.E. (1991). *Environmental chemistry*. (5^{th} ed.) USA: Lewis Publisher.