UNIT 07

ELECTROCHEMISTRY

Student Learning Outcomes (SLOs)

After completing this lesson, the student will be able to:

- Define redox reactions as simultaneous oxidation and reduction in terms of transfer of oxygen, hydrogen, electrons, and changes in oxidation state.
- Use roman numerals to indicate oxidation number of an element in a compound.
- Identify oxidizing and reducing agents in redox reactions.
- Recognize that the oxidation number of elements in their free state is zero.
- Derive the formula of ionic compounds from ionic charges and oxidation numbers.
- Identify the oxidation number of a monoatomic ion is the same as the charge on the ion.
- Explain that the sum of the oxidation numbers in a neutral compound is zero.
- Explain that the sum of the oxidation numbers in an ion is equal to the charge on the ion.
- Identify the redox reaction by the colour changes involved when using acidified aqueous potassium manganate(VII) to (II) or aquous potassium iodide.
- Define corrosion and discuss methods to prevent it. (Some examples may include barrier method such using paint, galvanizing, electroplating, sacrificial protection such as using magnesium blocks in ships.

INTRODUCTION

Oxidation-reduction reactions (redox) are fundamental chemical processes that play a crucial role in a variety of natural phenomena and industrial applications. This chapter examines the commonalities between different redox reactions, such as the rusting of iron objects, the burning of fuel in car engines, forest fires, and the metabolism of food in human and animal bodies. In addition, the importance of redox reactions in the production of electricity in batteries, in the decolorization of substances with household bleaches, and in the industrial production of important metals and chemicals is discussed.

One of the most important application of electrochemistry is batteries and fuels cells, which use chemical energy to generate electrical energy.

7.1 OXIDATION AND REDUCTION

We can define redox reactions in terms of transfer of oxygen, hydrogen, and electrons. In redox reactions oxidation and reduction occur simultaneously.

7.1.1 Oxidation-Reduction in Terms of Loss or Gain of Oxygen, Hydrogen or Electrons

Oxidation is gain of oxygen and reduction is loss of oxygen

In steel mills iron ores, usually oxides of iron are converted to the pure metal commercially by the reaction with coke (carbon) in the blast furnace. The carbon first reacts with air to form carbon monoxide, which in turn reacts with iron oxide as follows.

$$Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$$

Which substance in this reaction is losing oxygen? Which substance is gaining oxygen?

CO is gaining oxygen, so undergoing oxidation.

Fe_2O_3 is losing oxygen, so undergoing reduction.

7.1.2 Oxidation-reduction in terms of transfer of hydrogen

Oxidation is loss of hydrogen and reduction is gain of hydrogen

Acetylene (C_2H_2) is commercially used for cutting and welding metals. When acetylene burns, it produces a very hot flame known as oxy-acetylene flame. Following reaction takes place when it burns.

$$2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O$$

Which substance is losing hydrogen? Which substance is gaining hydrogen?

C_2H_4 is losing hydrogen, so unregoing oxidation.

O_2 is gaining hydrogen, so undergoing reduction.

7.1.3 Oxidation and Reduction in Terms of Transfer of Electrons

Oxidation is loss of electrons and reduction is gain of electrons.

For example, consider the following reaction.

$$CuO + Zn \longrightarrow Cu + ZnO$$

Copper(II)oxide and zinc(II)oxides are both ionic compounds. Rewrite this equation as ionic equation.

$$Cu^{2+} + O^{2-} + Zn \longrightarrow Cu + Zn^{2+} + O^{2-}$$

Oxide ions are spectator ions, so net equation is;

$$Cu^{2+} + Zn \longrightarrow Cu + Zn^{2+}$$

Cu^{2+} ions are gaining 2 electrons to form Cu. Its oxidation state changes from +2 to zero. So Cu^{2+} are undergoing reduction. The oxidation number of Zn is increasing from zero to +2, so Zn is losing 2 electrons and undergoing oxidation.

- Oxidation is defined as the loss of hydrogen, gain of oxygen or loss of electrons.
- Reduction is defined as the gain of hydrogen, loss of oxygen or gain of electrons.

Example 7.1: Identifying the element undergoing oxidation in terms of transfer of oxygen or hydrogen

Following reaction occurs when you burn Sui gas.

$$CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O + heat$$

Identify the element undergoing oxidation.

Problem solving strategy:

Identify the substance that gains O-atoms or loses H-atoms.

Solution:

Since C in CH_4 loses H-atoms and combines with oxygen atoms, thus C atoms undergo oxidation. At the same time O-atoms combine with H-atoms to form H_2O, thus O-atoms undergo reduction

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Redox in photography

A photographic film is basically an emulsion of silver bromide, (AgBr) in gelatin. When the film is exposed to light, Silver bromide granules become activated. This activation depends on the intensity of the light falling upon them. When exposed film is placed in the developer solution that is actually a reducing agent. Hydroquinone which is a mild reducing agent is used as developer. In hydroquinone the activated granules of silver bromide are reduced to black metallic silver. Reduced silver atoms form image.

$$Ag^+ + 1e^- \longrightarrow Ag_{(s)}$$

Inactivated silver bromide is removed from the film by using a solvent called a fixer. Sodium thiosulphate is used for this purpose. The areas of the film exposed to the light appear darkest because they have the highest concentration of metallic silver. Thus photography involves oxidation-reduction reaction.

CONCEPT ASSESSMENT EXERCISE 7.1

Identify elements undergoing oxidation and reduction in the following reactions:

1. $N_2 + 3H_2 \longrightarrow 2NH_3$
2. $2H_2 + O_2 \longrightarrow 2H_2O$
3. $Fe_2O_3 + 3CO \longrightarrow 2Fe + 3CO_2$
4. $4Al + 3O_2 \longrightarrow 2Al_2O_3$

Example 7.2: Identifying the element oxidized or reduced in terms of transfer of electrons

In the following reaction identify which element is oxidized and which element is reduced

$$2Ca + O_2 \longrightarrow 2CaO$$

Problem solving strategy:

Ca being metal forms cation by losing electrons (oxidation) and oxygen being non-metal gains electrons (reduction) to form anion.

Solution:

Remember that Group IIA metals form M^{2+} cations, and that Group VIA non-metals form X^{2-} anions. This means in this reaction each Ca atom loses two electrons to form Ca^{2+}, so it is oxidized. Each oxygen atom gains two electrons to form O^{2-}, so it is reduced.

$$2Ca \longrightarrow 2Ca^{2+} + 2e^- \text{ (oxidation)}$$
$$O_2 + 4e^- \longrightarrow 2O^{2-} \text{ (reduction)}$$

CONCEPT ASSESSMENT EXERCISE 7.2

In the following reactions, identify which element is oxidized and which element is reduced in terms of electron transfer.

$$4Na + O_2 \longrightarrow 2Na_2O$$
$$2Al + 3Cl_2 \longrightarrow 2AlCl_3$$
$$Mg + Cl_2 \longrightarrow MgCl_2$$

7.2 OXIDATION STATES AND RULES FOR ASSIGNING OXIDATION STATES

7.2.1 Oxidation States

Oxidation state or oxidation number is defined as the number of charges an atom will have in a molecule or a compound.

The elements that show an increase in oxidation number are oxidized. The elements that show a decrease in oxidation number are reduced. Do you think H in HCl is oxidized and Cl is reduced? Comparison of oxidation and reduction processes is given in table 7.1.

Table 7.1: Process leading to oxidation and reduction	
Oxidation	**Reduction**
Gain of oxygen	Loss of oxygen
Loss of hydrogen	Gain of hydrogen
Loss of electrons	Gain of electrons
Increase in oxidation number	Decrease in oxidation number

7.2.2 Rules for Assigning Oxidation States or Numbers

1. The oxidation state of any uncombined or free elements is always zero e.g., oxidation state of Zn, Na, H in H_2, S in S_8 etc is zero.
2. In simple ions, oxidation state is same as their charge e.g., oxidation state of Na in Na^{1+} and Ca in Ca^{2+} is +1 and +2 respectively.
3. In a complex ion the sum of oxidation states of atoms is equal to the charge on their ion. e.g., in CO_3^{2-}, the sum of oxidation states of C and 3O atoms is -2. Similarly, in NH_4^{1+}, the sum of oxidation states of N and 4H atoms is +1.
4. The oxidation number of each of the atoms in a molecule or compound is counted separately and their algebraic sum is zero e.g., In HCl, the sum of oxidation states of H and Cl atoms is zero. Similarly in CO_2, the sum of oxidation states of one C and 2 oxygen atoms is zero.

Table 7.2 shows the oxidation states of some of the elements in binary compounds which rarely change.

Table 7.2: Oxidation states of some elements in binary compounds that rarely change

Elements	Oxidation State
Group-IA	+1
Group-IIA	+2
Group-IIIA	+3
H	+1 (except in metal hydrides where it is -1)
Group-VIIA	-1
O	-2(except peroxides and in OF_2)

Monoatomic ions and their oxidation numders

The oxidation number of a monatomic ion is equal to its charge. For example, Na^+ is formed after a Na atom has lost one electron to gain a +1 charge. So its oxidation number is +1. Similarly, a chlorine atom forms a Cl^- ion after gaining one electron to obtain a -1 charge. So its oxidation number is -1. The oxidation number of an atom is the number of electrons the atom has lost or gained. Because a monatomic ion is formed by the gain or loss of electrons from a single atom, its charge is equal to its oxidation number..

Polyatomic ions and their oxidation numders

In a polyatomic ion, the sum of the oxidation numbers of all the atoms is equal to the charge on the ion. For example, in CO_3^{2-} ion the oxidation numbers of carbon and oxygen are +4 and -2 respectively. So, the sum of the oxidation number of one carbon atom and three oxygen atoms would be $1(+4) + 3(-2) = -2$ which is the charge on the ion.

7.2.3 Determining the Oxidation Number of an Atom in a Compound

Let's see how to use rules discussed in section 7.2.2 to determine the oxidation number of an atom of an element in a compound.

Example 7.3: Determining oxidation number

A device called Breathalyzer is used by police to test a person's breath for alcohol. It contains an acidic solution of potassium dichromate $K_2Cr_2O_7$. It is a strong oxidizing agent. Determine oxidation state of Cr in it.

Problem Solving Strategy:

Use rules 1 to 4 and table 7.1 to get as many oxidation numbers as you can. Use rule 4 to get oxidation number that has not been assigned.

Solution:

1. The oxidation number of K is +1, since it belongs to Group-1A. There are 2 K atoms therefore, overall oxidation number for K is $2(+1) = +2$
2. There are 7 oxygen atoms, therefore overall oxidation state for O is $7(-2) = -14$
3. Suppose oxidation for Cr is x, since there are two Cr atoms, therefore, overall oxidation sate for Cr is 2x.
4. The sum of oxidation numbers must be zero.

$$+2 + 2x + (-14) = 0$$

$$2x - 12 = 0$$

$$2x = 12$$

$$x = +6$$

Thus oxidation state for Cr in $K_2Cr_2O_7$ is +6

Example 7.4: Determining oxidation state

Boric acid H_3BO_3 is used in eye wash. What is the oxidation state of B in this acid?

Problem solving strategy:

Use rules and table 7.2 to get the oxidation state of H and O- atoms. Use rule 4 to get the oxidation state of B.

Solution:

1. There are 3 H-atoms, therefore, overall oxidation state for H is

$$3(+1) = +3$$

2. There are 3 O-atoms, therefore, overall oxidation state for O is

$$3(-2) = -6$$

3. Suppose the oxidation state for B is x.
4. The total oxidation states for all the atoms must be zero.

$$+3 + x + (-6) = 0$$

$$+3 + x - 6 = 0$$

$$x - 3 = 0$$

$$x = 3$$

Thus the oxidation state for B in H_3BO_3 is + 3.

CONCEPT ASSESSMENT EXERCISE 3.1

One major problem of air pollution is the formation of acid rain. Air pollutants such as SO_2 and NO_2 combine with oxygen and water vapours in the air to form H_2SO_4 and HNO_3. These acids fall to the ground with the rain, making the rain acidic. Clouds can also absorb the acids and carry them hundreds of kilometers away from where the pollutants are released. Determine the oxidation number of N in NO_2 and HNO_3, S in SO_2 and H_2SO_4

Example 7.5: Determining the oxidation number of an element in an ion.

What is the oxidation number of C in carbonate ion, CO_3^{-2}

Problem Solving Strategy:

(a) Use rule that oxidation number of O is -2

(b) Use rule 3 to find oxidation state of C

Solution:

Oxidation State of one C-atom + Oxidation state of 3-O atoms = -2

Oxidation state of C-atom + 3(-2) = -2

Oxidation state of C-atom -6 = -2

Oxidation state of C-atom =-2+6

Oxidation state of C-atom = +4

Thus the oxidation of C in carbonate ion is +4

CONCEPT ASSESSMENT EXERCISE 7.4

Determine the oxidation state of

1. S in sulphate ion, SO_4^{-2}
2. P in phosphate ion, PO_4^{-3}
3. N in ammonium ion, NH_4^{+1}

7.3 FORMULA OF AN IONIC COMPOUND

To determine the formula of an ionic compound from the ionic charges and oxidation numbers of the constituent ions, you must determine the simplest whole-number ratio of cations (positively charged ions) to anions (negatively charged ions) that results in a neutral compound. Ionic compounds are electrically neutral, meaning that the total positive charge of the cations is equal to the total negative charge of the anions. Let's go through step by step to derive the formula:

Example: Calcium chloride

Step 1: Identify the ions in the compound and their charges. Consider, for example, the combination of calcium ions (Ca^{2+}) and chloride ions (Cl^-).

Step 2: Determine the ratio of charges needed to balance each other. In this case, calcium has a charge of +2 and chloride has a charge of -1. You need two chloride ions for every calcium ion to balance the charges.

Step 3: Write a formula with assignments that represents the relationship defined in Step 2. The formula for calcium chloride is $CaCl_2$.

Example: Magnesium oxide

Step 1: Identify the ions and their charges - magnesium ions (Mg^{2+}) and oxide ions (O^{2-})

Step 2: Determine the charge - the magnesium has a charge of 2 and the oxide has a charge of -2. Thus, one magnesium ion is required for each oxide ion.

Step 3: Write the formula - MgO.

Example: Aluminum sulphate

Step 1: Identify the ions and their charges—aluminum ions (Al^{3+}) and sulphate ions (SO_4^{2-})

Step 2: Determine the charge - aluminum has a charge of +3 and sulphate has a charge of -2. You need two aluminum ions for every three sulphate ions to balance the charges.

Step 3: Write the formula - $Al_2(SO_4)_3$.

It is important to remember that when writing a formula, parentheses must be used to add polyatomic ions when more than one is needed to balance the charges. By following these steps and understanding the charges on the ions, you can derive the formulas of various ionic compounds.

7.3.1 Use of Roman numerals as oxidation number

Roman numerals are used to indicate the oxidation states of elements in compounds, when a metal exhibits variable oxidations in compounds. For examples in transition metal compounds.

$CuSO_4$ is written as copper (II) sulphate.

$FeCl_2$ is written as iron (II) chloride.

$FeCl_3$ is written as iron (III) chloride

7.4 OXIDIZING AND REDUCING AGENTS

An oxidizing agent is a substance that causes another substance to oxidize by taking electrons from it. It is often called an electron acceptor because it accepts electrons during a reaction.

Oxidizing agents themselves are reduced in the process (they gain electrons). Examples of common oxidizing agents are oxygen (O_2), hydrogen peroxide (H_2O_2), chlorine (Cl_2) and potassium permanganate ($KMnO_4$). **A reducing agent is a substance that causes another substance to be reduced by donating electrons to it.** It is often called an electron donor because it loses electrons during the reaction. The reducing agents themselves are oxidized in the process (they lose electrons). Examples of common reducing agents include hydrogen gas (H_2), metal hydrides (such as $NaBH_4$), carbon monoxide (CO), and metals such as zinc (Zn) and aluminum (Al). it.

For example, in the reaction betw ..n sodium and chlorine to form sodium chloride.

$$2Na + Cl_2 \longrightarrow 2NaCl$$

Na is reducing agent as it is oxidized whereas Cl_2 is oxidizing agent as it is reduced.

Activity 7.1

Prepare solutions of ferrous sluphate($FeSO_4$) and potassium permanganate ($KMnO_4$) in separate beakers. Transfer about $10cm^3$ of ferrous sulphate solution in a test tube. Add about $10cm^3$ of dill. H_2SO_4 in it. Now add few drops of $KMnO_4$solution in the test tube. What happens?

$FeSO_4$reduces $KMnO_4$, so its purple colour is discharged. $KMnO_4$ oxidizes $FeSO_4$ in this reaction. $FeSO_4$ is reducing agent whereas $KMnO_4$is oxidizing agent

A color change during a chemical reaction may indicate a redox reaction.

Potassium permanganate (VII) is an oxidizing agent.

Often used to test for the presence of reducing agents.

When acidified $KMnO_4$ is added to the reducing agent, it changes from purple to colorless.

The above reaction discharges the purple colour of $KMnO_4$.

Therefore, the solution must contain a reducing agent that reduces MnO_4^{2-} ions (purple) to Mn^{2+} ions (colourless).

During this reaction, the oxidation number of Mn changes from +7 to +2.

Activity 7.2

Prepare an aqueous solution of potassium iodide and transfer it to a 10 cm^3 test tube. Add about 5 cm^3 of hydrogen peroxide to it. what's going on, the solution turns reddish-brown, indicating the formation of iodine.

Potassium iodide is a reducing agent. Often used to test for the presence of oxidizing agents. When a potassium iodide solution is added to an acidified hydrogen peroxide solution, the solution turns reddish-brown. The appearance of this color is due to the formation of iodine I_2.

KI is oxidized and H_2O_2 is reduced.

This redox reaction can be confirmed by a color change from colorless to reddish-brown.

$$H_2O_{2(l)} + 2KI_{(aq)} + 2H^+_{(aq)} \longrightarrow 2K^+_{(aq)} + I_{2(aq)} + 2H_2O$$

7.4.2 How can you identify oxidizing and reducing agents in a chemical reaction?

Consider the following reaction that takes place in the manufacture of steel.

$$Fe_2O_3 + 3CO \longrightarrow 2Fe \ +3CO_2$$

To identify the oxidizing and reducing agents, work out the oxidation states of all the elements involved in the reaction.

$$\overset{2(+3)\,(-2)3}{Fe_2O_3} + \overset{+2-2}{3CO} \longrightarrow \overset{0}{2Fe} + \overset{+4(-2)2}{3CO_2}$$

(i) Carbon is being oxidized because there is an increase in its oxidation state.

(ii) Fe is being reduced because there is a decrease in its oxidation state.

(iii) The reactant CO contains the C that is being oxidized, so CO is reducing agent.

(iv) The reactant Fe_2O_3 contains the Fe that is being reduced. So Fe_2O_3 is oxidizing agent.

Oxidizing or reducing agent is the whole molecule or formula unit and not the atom that has undergone change in oxidation number.

Example 7.6: Identifying the oxidizing and reducing agents

Tungsten is used to make filaments for electric bulbs because it has the highest melting point and high electrical resistance. This metal is obtained from tungsten (VI) oxide, WO_3 by reducing it with hydrogen gas.

$$WO_3 + 3H_2 \longrightarrow W + 3H_2O$$

Identify the oxidizing and reducing agents in this reaction.

Problem solving strategy:

Step 1: Workout the oxidation states of all the elements involved in the reaction.

Step 2: Note the element that is undergoing an increases in its oxidation state. Since it is being oxidized. The reactant that contains this element is reducing agent.

Step 3: Note the element that is undergoing a decrease in its oxidation state. Since it is being reduced. The reactant that contains this element is oxidizing agent.

Solution:

First assign oxidation numbers to each atom.

$$\overset{+6\,(-2)3}{WO_3} + \overset{0}{3H_2} \longrightarrow \overset{0}{W} + \overset{2(+1)\ -2}{3H_2O}$$

Because the oxidation number of W decreases, so WO_3 is an oxidizing agent. Similarly the oxidation number of H increases, therefore $H_2$is reducing agent.

CONCEPT ASSESSMENT EXERCISE 7.5

1. Identify oxidizing and reducing agents in the following reactions:
   a) $2S + Cl_2 \longrightarrow S_2Cl_2$
   b) $2Na + Br_2 \longrightarrow 2NaBr$
2. Differentiate between oxidizing and reducing agents.

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Redox in photography

Silver is very soft metal. Silver atoms have weak interactions and are loosely packed together. Silver tarnishes in air when it comes in contact with trace quantities of H_2S or SO_2 in the atmosphere or food such as eggs, that are rich in sulphur compounds. Silver tarnish is silver sulphide that gives silver blackish appearance. Due to this reason decorative and practical objects made of solid silver gradually turn black and lose shining appearance. Decorative and practical objects are plated with a thin layer of silver. Atoms in thin layers firmly adhere to the metal atoms of the object and form a durable layer. An article thickly plated with silver contains many layers of silver atoms. Such layers form soft covering. These layers gradually turn black.

7.5 CORROSION AND ITS PREVENTION

7.5.1 Corrosion

Corrosion is a natural electrochemical process that occurs when a metal reacts with its environment. **In this reaction a metal reacts with oxygen and moisture in the atmosphere.** Corrosion converts refined metals to the more stable metal oxides. It can cause significant damage to structures, vehicles, and equipments.

Most familiar example of corrosion is the formation of rust on iron. Oxygen and water are necessary for iron to rust. Corrosion is an oxidation-reduction reaction. A region of metal surface that has relatively less moisture, acts as anode. Will Fe oxidize in this region?

$$Fe\,(s) \longrightarrow Fe^{+2}(aq) + 2e^-$$

Another region on the surface of metal that has relatively more moisture acts as cathode. The electrons released in the oxidation process reduce atmospheric oxygen to hydroxyl ions.

$$O_2 + 2H_2O + 4e^- \longrightarrow 4OH^-$$

The Fe^{+2} ions formed at the anodic regions flow to the cathodic regions through the moisture on the surface. Here Fe^{+2} ions further react with oxygen to form rust, $Fe_2O_3.xH_2O$

7.5.2 Prevention of Corrosion

Corrosion is a widespread issue that affects industries, infrastructures, and everyday objects. Therefore, understanding corrosion and implementing preventive measures are crucial. Prevention of corrosion is an important way of conserving our natural resources. Following methods have been devised to protect metals from corrosion:

1. **Coating with paint:**

   Corrosion can be prevented by applying protective coating such as paint or epoxy creating a barrier between metal surface and its environment. Paint is cheap and can be applied easily. Paint is used to protect many everyday steel objects such as cars, trucks, trains, bikes, bridges etc. Paint also provides visual appeal.

2. Alloying:

The tendency of iron to oxidize can be greatly reduced by alloying it with other metals. For example, adding chromium to iron forms stainless steel, which is highly resistant to corrosion.

3. Coating with a thin layer of another metal:

Metals that readily corrode can be protected by coating with a thin layer of another metal that resists corrosion. This can be done by:

(a) Tinning

(b) Galvanizing

(c) Electroplating

(a) Tinning:

In the process of tin plating, clean iron sheet is dipped in a bath of molten tin. It is then passed through hot pair of rollers. Tin protects iron effectively, since, it is very stable.

(b) Galvanizing (Coating with Zinc):

The process of galvanizing consists of dipping a clean iron sheet in a hot zinc chloride bath and heating. After this sheet is rolled into zinc bath and cooled.

(c) Electroplating:

In electroplating an electrolytic process is used to deposit one metal on another metal.

4. Cathodic Protection:

Cathodic protection also called sacrificial protection is the process in which the metal to be protected from corrosion is made cathode and connected to an active metal such as magnesium or aluminum. These metals are more active than iron, they act as anode. The more active metal oxidizes itself and saves iron from corrosion. Cathodic protection is employed to prevent iron and steel structures such as pipes, tanks, oil rigs etc in the moist underground and marine environment. The large bars of magnesium are used to help protect the ship against rusting.

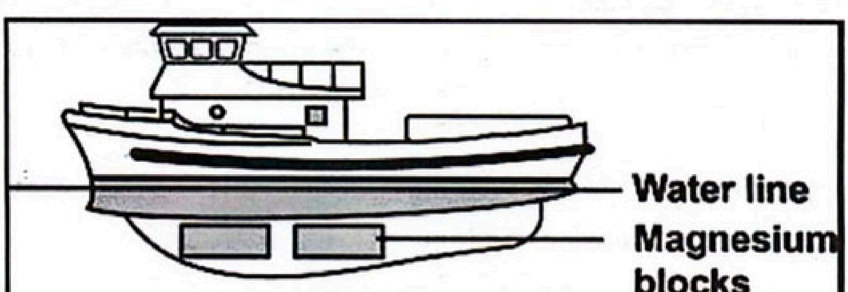

Figure 7.3: Cathodic Protection

KEY POINTS

- Oxidation is the gain of oxygen atom or loss of hydrogen atom or loss of electrons by a substance.
- Reduction is the loss of oxygen atom or gain of hydrogen atom or gain of electrons by a substance.
- Oxidation state or oxidation number is defined as the number of apparent charges that an atom will have in a molecule.
- The sum of oxidation state of all the atoms in a molecule of compound is zero.
- An oxidizing agent is the reactant containing the element that is reduced in a reaction.
- A reducing agent is the reactant containing the element that is oxidized in a reaction.
- Corrosion is the process in which a metal reacts with oxygen and moisture in the atmosphere.
- Electrolytic process used to deposit one metal on another metal is called electroplating.
- Cathodic protection is the process in which metal that is to be protected from corrosion is made cathode and is connected to metals such as magnesium or aluminum.

References for additional information

- B.Earl and LDR Wilford, Further Advanced Chemistry.
- B.Earl and LDR Wilford, Introducion to Advanced Chemistry.
- David E. Goldberg, Fundamental of Chemistry.
- Addison Wesley, Chemistry

REVIEW QUESTIONS

1. **Encircle the correct answer.**

(i) In which of the following changes, the nitrogen atom is reduced.

(a) N_2 to NO (b) N_2 to NO_2

(c) N_2 to NH_3 (d) N_2 to HNO_3

(ii) Which of the following changes reaction is an example of oxidation

(a) Chlorine molecule to chloride ion (b) Silver atoms to silver (I) ion

(c) Oxygen molecule to oxide ion (d) Iron (III) ion to iron(II) ion

(iii) Which of the following elements in the given reaction is reduced?

$$ZnO + H_2 \longrightarrow Zn + H_2O$$

(a) H_2 (b) ZnO

(c) Zn (d) O

(iv) Consider the following reaction:

$$H_2S + Cl_2 \longrightarrow 2HCl + S$$

In this reaction H_2S behaves as

(a) Reducing agent (b) Oxidizing agent

(c) Catalyst (d) Electrolyte

(v) The oxidation state of Cr in $K_2Cr_2O_7$ is

(a) +12 (b) +6

(c) +3 (d) -6

2. **Give short answer.**

(i) What is oxidation state?

(ii) What is the oxidation number of Cr in chromic acid (H_2CrO_4)?

(iii) Identify reducing agent in the following reaction

(iv) $CuO + H_2 \longrightarrow Cu + H_2O$

(v) Why tin plated steel is used to make food cans?

(vi) Explain one example from daily life which involves oxidation-reduction reaction?

3. Compare and contrast oxidation and reduction.

4. Define oxidation and reduction in terms of loss or gain of electrons.

5. Explain how food and beverage industries deal with corrosion.

6. State the substances which are oxidized or reduced. Give reason for your answer.

(a) $N_2 + 3H_2 \longrightarrow NH_3$

(b) $CO_2 + 2Mg \longrightarrow 2MgO + C$

(c) $Mg + H_2O \longrightarrow MgO + H_2$

(d) $H_2S + Cl_2 \longrightarrow 2HCl + S$

(e) $2NH_3 + 3CuO \longrightarrow 3Cu + N_2 + 3H_2O$

7. (a) Define oxidation number or oxidation state.

(b) Find the oxidation state of nitrogen in the following compounds.

(i) NO_2 (ii) N_2O (iii) N_2O_3 (iv) HNO_3

8. Find the oxidation state of S in the following compounds.

(a) H_2S (b) H_2SO_3 (c) $Na_2S_2O_3$

9. (a) Define oxidizing and reducing agents.

(b) Identify the oxidizing agents and reducing agents in the following reactions:

(i) $H_2S + Cl_2 \longrightarrow 2HCl + S$

(ii) $2FeCl_2 + Cl_2 \longrightarrow 2FeCl_3$

(iii) $2KI + Cl_2 \longrightarrow 2KCl + I_2$

(iv) $Mg + 2HCl \longrightarrow MgCl_2 + H_2$

10. Hydrogen peroxide reacts with silver oxide and lead(II) sulphide according to the following equations.

Is hydrogen peroxide an oxidizing or reducing agent in these reactions. Argue.

(i) $H_2O_2 + Ag_2O \longrightarrow 2Ag + H_2O + O_2$

(ii) $4H_2O_2 + PbS \longrightarrow PbSO_4 + 4H_2O$

PROJECT

Prepare a report about what type of chemical reaction is corrosion, giving suitable examples.