CHAPTER 1

INTRODUCTION

1.1. Background to Corrosion

Corrosion, a pervasive and costly phenomenon, describes the degradation of materials, primarily metals, through chemical or electrochemical reactions with their surrounding environment. Fundamentally, corrosion involves the irreversible transformation of refined metals into more stable, lower-energy compounds, such as oxides, sulfides, or hydroxides. This process not only compromises the structural integrity and functionality of metallic components but also leads to significant economic losses, environmental degradation, and safety hazards across diverse sectors.

1.1.1. Definition and Significance of Corrosion

From a chemical perspective, corrosion is essentially an oxidation-reduction (redox) reaction. The metal undergoes oxidation, losing electrons to form positively charged ions (cations), while another species, typically oxygen in the presence of water, undergoes reduction by accepting these electrons(Lu *et al.*, 2023). The electron transfer process results in the formation of corrosion products, such as rust on iron.

Chemical Equations:

Let's illustrate with the example of iron (Fe) corrosion:

- Anodic Reaction (Oxidation - Metal Dissolution):

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

(Solid iron is oxidized to iron(II) ions, releasing two electrons.)

- Cathodic Reaction (Reduction - typically Oxygen reduction in presence of water):

$$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$$

(Oxygen and hydrogen ions (from the water) accept electrons to form water.)

- Overall Corrosion Reaction (Simplified):

$$2Fe(s) + O_2(g) + 4H^+(aq) \rightarrow 2Fe^{2+}(aq) + 2H_2O(l)$$

(Solid iron reacts with oxygen and hydrogen ions to form iron ions and water).

Note: This overall equation is often followed by further reactions depending on the pH. The Iron (II) may be further oxidised to Iron (III).

The significance of corrosion lies in its destructive potential and its far-reaching consequences (Lu *et al.*, 2023). The degradation of critical structures leads to costly repairs, service disruptions, and potential catastrophes. Therefore, understanding the corrosion mechanisms and influencing factors is essential for designing effective mitigation and prevention strategies.

1.1.2. Common Types of Corrosion

Corrosion can manifest in several forms, each requiring specific monitoring and prevention techniques:

- Uniform Corrosion: Characterized by a relatively even attack across the exposed metal surface. While predictable, it leads to material loss.

- Pitting Corrosion: Localized, deep holes or cavities on the metal surface, difficult to detect and monitor.

- Galvanic Corrosion: Occurs when two dissimilar metals are in contact in an electrolytic environment. The more active metal corrodes faster.

- Crevice Corrosion: Localized attack in crevices and shielded areas. The mechanism is based on differences in oxygen concentrations in different parts of the material leading to formation of differential aeration cell(Lu *et al.,* 2023).

- Intergranular Corrosion: Attack along grain boundaries of a metal.

- Stress Corrosion Cracking: Cracks due to tensile stress and a corrosive environment.

1.1.3. Factors Influencing Corrosion

The rate and extent of corrosion are influenced by multiple interacting factors:

- Environmental Conditions: Temperature, humidity, and presence of corrosive substances. Increased temperatures usually accelerate corrosion and the presence of water provides the needed electrolytic environment (Guo *et al.,* 2023).

- Chemical Composition of the Metal: The specific metal and its reactivity play a key role. Active metals readily corrode while noble metals are more corrosion-resistant.

- pH Levels: The acidity or alkalinity of the environment is the focus of the project. The concentration of hydrogen ions (H^+) will heavily influence the electrochemical process.

- Oxygen Concentration: The availability of oxygen (O2) in the environment influences the rate of the cathodic reaction. As shown earlier in the cathodic reaction.

- Mechanical Stress: Applied stresses can exacerbate corrosion.

1.2. The Role of pH in Corrosion

The pH of an environment, a measure of its acidity or alkalinity, is a critical factor that profoundly influences the rate and mechanism of corrosion processes. pH directly affects the concentration of hydrogen ions (H^+) and hydroxide ions (OH^-) in aqueous solutions, which are key reactants and products in the electrochemical reactions underlying corrosion (Guo *et* *al.,* 2023). Understanding how different pH levels impact corrosion is essential for developing effective corrosion prevention strategies and material selection.

1.2.1. Acidic and Alkaline Environments

- Acidic Environments ($pH < 7$): Acidic conditions, characterized by a high concentration of H^+ ions, generally accelerate the corrosion of many metals. In acidic solutions, H^+ ions readily participate in cathodic reactions, effectively acting as a catalyst for the overall corrosion process. This increases the rate of electron transfer, leading to an accelerated metal dissolution. Furthermore, acids often attack the oxide layers which may form on the metals, preventing the formation of passivation layer and making the metal more susceptible to further corrosion(Guo *et al.,* 2023).

Chemical Equations (Example with Iron in Acidic Solution):

Anodic Reaction (unchanged from before):

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

Cathodic Reaction (in Acidic conditions):

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

(Hydrogen ions are reduced to form hydrogen gas)

Overall Corrosion Reaction (Simplified) in Acidic condition:

$$Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$$

- Alkaline Environments ($pH > 7$): Alkaline environments, characterized by a high concentration of OH^- ions, can have a more complex influence on corrosion. In some cases, alkaline conditions can reduce corrosion rates by forming stable oxide layers (passivation) on the metal surface, thereby protecting the underlying metal from further corrosion. However, in other situations, alkaline conditions can promote the formation of soluble metal hydroxide complexes, leading to accelerated corrosion. The effect of alkaline solutions is highly specific to the metal in question (Guo *et al.,* 2023).

Chemical Equations (Example with Aluminum in Alkaline Solution):

Anodic Reaction:

$$Al(s) \rightarrow Al^{3+}(aq) + 3e^{-}$$

Reaction of Aluminum ions with Hydroxide ions:

$$Al^{3+}(aq) + 4OH^{-}(aq) \rightarrow [Al(OH)_4]^{-}(aq)$$

(Formation of soluble aluminate complex.)

Overall Corrosion Reaction (Simplified) in Alkaline condition:

$$Al(s) + 4OH^{-}(aq) \rightarrow [Al(OH)_4]^{-}(aq) + 3e^{-}$$

1.2.2. Electrochemical Principles Relating pH to Corrosion

The relationship between pH and corrosion can be better understood through electrochemical principles. Corrosion involves the establishment of electrochemical cells on the metal surface, comprising anodic sites where oxidation occurs and cathodic sites where reduction occurs. The difference in potential between these sites drives the electron transfer process. The pH of the environment impacts the electrochemical potentials at both the anode and the cathode, significantly affecting the overall reaction kinetics (Guo *et al.*, 2023).

- Nernst Equation: The Nernst equation shows how the concentration of ions, such as H^{+}, influences the electrochemical potential:

$$E = E^{0} - (RT/nF) * \ln(Q)$$

where:

E = electrode potential

E^0 = standard electrode potential

R = universal gas constant

T = temperature in Kelvin

n = number of electrons transferred

F = Faraday's constant

Q = reaction quotient

This equation clearly shows how the concentration of hydrogen ions, reflected by the pH, affects the potential and drives the corrosion process.

- Electrode potentials: Changing the pH of the solution affects the reduction potential of the electrolyte and thus how susceptible the metals are to corrosion.

1.2.3. Passivation and Corrosion Inhibition at Varying pH Levels

- Passivation: At certain pH levels, some metals can form a thin, protective oxide layer on their surface, a process called passivation (Ouyang *et al.*, 2021). This layer acts as a barrier, inhibiting further corrosion. For instance, metals like aluminum and chromium display passivation in specific pH ranges.

- Corrosion Inhibition: Understanding how pH affects corrosion is crucial in the design of effective corrosion inhibitors, which may act by shifting the pH or forming protective films. These inhibitors prevent corrosion by changing the pH or by changing the surface properties of the metal (Ouyang *et al.*, 2021).

1.3. Research Motivation and Scope

The impetus for this research stems from the pervasive and economically significant problem of metal corrosion and the critical role that environmental conditions, particularly pH, play in influencing this process. A thorough understanding of the effects of pH on metal corrosion is not just of academic interest but has wide-ranging practical implications across diverse fields. This section outlines the motivations for undertaking this research and defines the scope of the project, ensuring clarity and focus.

1.3.1. Relevance to Industrial and Environmental Concerns

- Industrial Significance: Corrosion poses a major challenge in numerous industries, such as construction, transportation, chemical processing, and energy production (Ouyang *et al.*, 2021). Metal components in these sectors are frequently exposed to a variety of environmental conditions, including a wide range of pH levels. For example:
    - Pipelines: In the oil and gas industry, pipelines are often subject to varying pH conditions due to the presence of different chemicals and soil types, leading to corrosion and potential leaks.
    - Water Treatment Plants: Structures and equipment in water treatment facilities are exposed to varying pH levels during different stages of purification and are susceptible to corrosion, which can affect the efficiency of the plants.
    - Marine Environments: Ships, offshore platforms, and other structures are constantly exposed to seawater, the pH of which can influence the rate of corrosion.
    - Chemical Processing: The equipment in chemical processing plants comes in contact with a variety of chemicals, including acids and bases, which can significantly affect the corrosion rates of the metal used.

- Environmental Concerns: Corrosion not only leads to material degradation but also contributes to environmental pollution. The release of corroded metal ions into soil and water can have detrimental effects on ecosystems. For example, heavy metals released into water bodies can contaminate them, affecting aquatic life and potentially contaminating drinking water.

Chemical Equation :

$$Fe(s) + O_2(g) + 2H_2O(l) \rightarrow Fe(OH)_2(s)$$

$$2Fe(OH)_2(s) + \frac{1}{2}O_2(g) + H_2O \rightarrow 2Fe(OH)_3(s)$$

(These equations show the conversion of iron into corrosion product)

Therefore, the need to understand and mitigate corrosion processes is critical for sustainability and environmental protection.

1.3.2. Objectives of the Project

The specific objectives of this project, focus on the experimental investigation of the effect of pH on corrosion rate. The specific scope includes:

- Metal Selection: The project will focus on specific metals, such as iron (Fe), aluminum (Al), and copper (Cu), chosen based on their industrial significance and varied electrochemical behavior.

- pH Range: The study will examine a range of pH values, typically from strongly acidic to strongly alkaline, to cover the major trends in corrosion behavior. The exact range of pH depends on the available materials.

- Electrochemical Techniques: The project will employ electrochemical techniques, such as potentiodynamic polarization and/or electrochemical impedance spectroscopy (EIS), to accurately measure corrosion rates.

- Controlled Conditions: The experiments will be performed under controlled conditions (temperature, electrolyte concentrations etc.) to isolate the specific impact of pH.

CHAPTER 2

LITERATURE REVIEW

2.1. Fundamentals of Electrochemistry

This section lays the groundwork by discussing the key electrochemical concepts that underlie corrosion processes.

2.1.1. Redox Reactions in Corrosion

As discussed before, corrosion is fundamentally an electrochemical process involving oxidation and reduction reactions. Oxidation occurs at the anode, where the metal loses electrons and forms metal ions. Simultaneously, reduction takes place at the cathode, where a species in the environment (typically oxygen or hydrogen ions) gains electrons (Haque *et al.*, 2023). These two half-reactions must occur in concert for the corrosion process to proceed. This can be illustrated by the simplified iron rusting process:

Anodic Reaction:

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

(Oxidation of iron.)

Cathodic Reaction (in neutral or acidic environment):

$$O_2(g) + 4H^{+}(aq) + 4e^{-} \rightarrow 2H_2O(l)$$

(Reduction of Oxygen)

The presence of an electrolyte (aqueous solution) is essential for completing the circuit by facilitating ion transport between the anode and cathode. Understanding the principles of redox reactions is paramount to interpreting corrosion mechanisms.

2.1.2. Electrochemical Potentials and the Nernst Equation

Electrochemical potentials (electrode potentials) represent the thermodynamic tendency of a metal to undergo oxidation or reduction under specific conditions. The standard electrode potential (E^o) is measured under standard conditions (298 K, 1 atm pressure, and 1 M concentration of ions). However, practical corrosion scenarios often deviate from these standard conditions (Haque *et al.*, 2023). The Nernst equation describes how the electrode potential changes with temperature and ion concentrations (and pH):

$$E = E^o - (RT/nF) * \ln(Q)$$

Where:

E = Electrode potential under non-standard conditions

E^o = Standard electrode potential

R = Universal gas constant (8.314 J/(mol·K))

T = Temperature in Kelvin

n = Number of moles of electrons transferred in the redox reaction

F = Faraday constant (96485 C/mol)

Q = Reaction quotient (ratio of products to reactants)

This equation is critical for understanding how changes in ion concentrations, especially of H^+ ions, due to pH variations, alter the electrode potentials and drive the corrosion process.

2.2. Effects of pH on Metallic Corrosion

This section delves specifically into the effects of varying pH levels on corrosion.

2.2.1. Influence of Acidic Environments

Acidic conditions ($pH < 7$) often accelerate corrosion because they provide an abundance of H^+ ions, which readily participate in the cathodic reduction reaction. For instance, the cathodic reduction of hydrogen ions is generally more favorable than the reduction of water in neutral or alkaline solutions, resulting in increased electron consumption and a faster rate of metal oxidation. Furthermore, strong acids are highly corrosive and may attack the passivation layer if it exists (Haque *et al.*, 2023).

Chemical Equation (Hydrogen ion reduction)

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

This is the hydrogen evolution cathodic reaction that is dominant in many acidic solutions, as discussed in the introduction.

2.2.2. Influence of Alkaline Environments

Alkaline environments ($pH > 7$) can have varied impacts on corrosion. In some cases, they may promote the formation of stable oxide films (passivation) on the metal surface, thus inhibiting further corrosion. However, in other cases, alkaline solutions may facilitate the formation of soluble metal hydroxide complexes, accelerating corrosion (Dong *et al.*, 2021). For example, the corrosion of aluminum in strongly alkaline conditions is an important practical example:

Chemical Equation (Simplified overall reaction):

$$2Al(s) + 6H_2O(l) + 2OH^-(aq) \rightarrow 2[Al(OH)_4]^-(aq) + 3H_2(g)$$

(Corrosion of aluminum to form aluminate ion in an alkaline solution.)

2.2.3. Passivation and Corrosion Inhibition at Varying pH Levels

The ability of metals like aluminum, chromium, and titanium to form protective oxide layers is known as passivation. This phenomenon is often strongly dependent on the pH of the surrounding environment. At certain pH levels, the oxide layer is stable and prevents further corrosion, while at other pH values, the layer may be dissolved, leaving the metal vulnerable to corrosion. The use of corrosion inhibitors is critical in several industries and one way of achieving corrosion control is by changing the pH of the environment (Dong *et al.*, 2021).

2.3. Specific Metals and Corrosion

This section focuses on the specific corrosion behavior of the metals that have been chosen to study (iron, aluminum, and copper), based on literature information.

2.3.1. Overview of the Properties of Selected Metals

Overview of the Chemical and Physical Properties of Iron

Iron (Fe), a transition metal with atomic number 26 (electron configuration [Ar] $3d^6 4s^2$) and an atomic weight of approximately 55.845 g/mol, is a vital element on Earth. Its versatile properties make it indispensable across various industries, although its susceptibility to corrosion remains a persistent concern (Dong *et al.*, 2021). This section outlines its key chemical and physical properties, focusing on aspects relevant to corrosion.

Chemical Properties:

1. Electronic Configuration and Oxidation States:

- The electronic configuration ([Ar] $3d^6 4s^2$) determines its transition metal behavior (Dong *et al.*, 2021). The partially filled d orbitals allow for multiple oxidation states, with +2 (ferrous, Fe^{2+}) and +3 (ferric, Fe^{3+}) being the most relevant in corrosion processes. While these are most common, Fe can be 0, +4, and +6 oxidation states.

2. Reactivity:

- Iron is moderately reactive and is highly susceptible to oxidation, especially in the presence of oxygen and water. This is described by the following reaction:

$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

(Formation of Iron (III) oxide or rust).

- Iron reacts with non-oxidizing acids, such as dilute hydrochloric acid (HCl), to produce hydrogen gas and iron(II) ions:

$Fe(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2(g)$

- With strong oxidizing acids, it may form a passive layer of oxide and inhibit corrosion.

3. Complex Formation:

- Iron forms coordination complexes with ligands such as water, hydroxide (OH^-), chloride (Cl^-), and cyanide (CN^-). These complexes can affect the solubility and behavior of iron ions in solution. For example:

$Fe^{2+}(aq) + 6H_2O(l) \rightleftharpoons [Fe(H_2O)_6]^{2+}(aq)$

(Formation of hydrated iron (II) ions).

4. Electrochemical Activity:

- Iron is a relatively active metal in the electrochemical series. Its standard electrode potential for the Fe^{2+}/Fe couple is -0.44 V. The electrode potential for Fe^{3+}/Fe is +0.04 V and for Fe^{3+}/Fe^{2+} is +0.77 V. This shows it has a high tendency to oxidise in presence of oxygen and water (Dong *et al.*, 2021).

5. Corrosion Products:

- The primary corrosion product, often called rust, is hydrated iron(III) oxide, written as $Fe_2O_3 \cdot nH_2O$. Iron hydroxides can also form in different conditions:

$Fe^{2+}(aq) + 2OH^{-}(aq) \rightarrow Fe(OH)_2(s)$

$Fe^{3+}(aq) + 3OH^{-}(aq) \rightarrow Fe(OH)_3(s)$

(Formation of iron (II) and Iron (III) hydroxide).

These corrosion products are typically porous and non-adherent, providing little protection to the underlying metal.

Physical Properties:

1. Appearance:

- Pure iron is a lustrous, silvery-gray metal. It is also malleable and ductile, although these properties are affected by impurities and temperature.

2. Melting and Boiling Points:

- Iron has a high melting point of 1538 °C and a boiling point of 2862 °C.

3. Density:

- The density of iron is approximately 7.87 g/cm^3.

4. Crystal Structure:

- At room temperature, iron adopts a body-centered cubic (BCC) crystal structure.

5. Magnetic Properties:

- Iron is ferromagnetic at room temperature, which makes it strongly attracted to a magnetic field and able to be magnetized. This is due to the presence of unpaired electrons in its 3d orbitals.

6. Mechanical Properties:

- Iron possesses high tensile strength and hardness. These properties, along with its malleability and ductility, make iron an excellent material in manufacturing. The mechanical properties vary when impure elements or alloys are involved.

Relevance to Corrosion:

These chemical and physical properties contribute to iron's corrosion behavior:

- Its reactivity with oxygen and acids results in rapid degradation.
- The formation of porous rust exacerbates the problem and allows the corrosion process to continue.
- Its electrochemical activity and standard potential make it prone to oxidation.
- Its complex formation capability affects the solubility and mobility of iron ions, thus impacting the overall process.

Overview of the Chemical and Physical Properties of Aluminum

Aluminum (Al), a main group metal with atomic number 13 and an atomic weight of approximately 26.98 g/mol, is the most abundant metal in the Earth's crust. Its unique combination of properties, including its light weight and resistance to corrosion, makes it a widely used material in various applications. This section outlines its key chemical and physical properties, with an emphasis on their relation to its corrosion behavior (Li *et al.*, 2022).

Chemical Properties:

1. Electronic Configuration:

- Aluminum's electronic configuration is [Ne] $3s^2 3p^1$, revealing three valence electrons. This electron arrangement governs its chemical reactivity and bonding behavior.

2. Oxidation State:

- Aluminum typically exhibits a +3 oxidation state in its compounds. It loses all three of its valence electrons to form Al^{3+} ions.

3. Reactivity:

- Aluminum is a highly reactive metal thermodynamically. However, it readily forms a passive oxide layer upon exposure to air, which gives it its corrosion resistance, making it less reactive than expected (Li *et al.*, 2022).
- Aluminum reacts readily with acids and bases.
- With dilute non-oxidizing acids like hydrochloric acid (HCl), aluminum reacts to produce hydrogen gas and aluminum chloride:

$$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$$

Aluminum reacts with hydrochloric acid.

- With strong bases, like sodium hydroxide (NaOH), aluminum reacts to produce soluble aluminate complexes:

$$2Al(s) + 2NaOH(aq) + 6H_2O(l) \rightarrow 2Na[Al(OH)_4](aq) + 3H_2(g)$$

Aluminum reacts with sodium hydroxide to form a soluble aluminate (Li *et al.*, 2022).

4. Complex Formation:

- Aluminum forms various complexes with ligands such as water, hydroxide ions (OH^-), and halides. As seen previously, these complexes play important roles in determining its behavior in different environments, such as the soluble aluminate ion. The hydrated aluminium ion may also exist in aqueous solution (Li *et al.,* 2022).

$Al^{3+}(aq) + 6H_2O(l) \rightleftharpoons [Al(H_2O)_6]^{3+}(aq)$

5. Electrochemical Activity:

- Aluminum is a highly reactive metal and has a standard reduction potential of -1.66 V, which indicates its tendency to oxidise. However, the passive layer provides protection against corrosion (Li *et al.,* 2022).

6. Amphoteric Nature:

- Aluminum is amphoteric, meaning it can react with both acids and bases. The reactions with acids and bases cause corrosion in extreme pH ranges (Li *et al.,* 2022).

7. Corrosion Products:

- The primary corrosion product of aluminum is aluminum oxide (Al_2O_3). This oxide is often present as a hydrated oxide ($Al_2O_3 \cdot nH_2O$) or aluminum hydroxide ($Al(OH)_3$). The formation of a dense, adherent layer of Al_2O_3 at neutral pH provides corrosion resistance (Song *et al.,* 2018).

$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

Reaction with oxygen to form aluminum oxide.

Physical Properties:

1. Appearance:

- Aluminum is a lustrous, silvery-white metal. It is also highly malleable and ductile.

2. Melting and Boiling Points:

- Aluminum has a relatively low melting point of 660 °C and a boiling point of 2519 °C.

3. Density:

- Aluminum is a lightweight metal, with a density of approximately 2.7 g/cm^3.

4. Crystal Structure:

- Aluminum has a face-centered cubic (FCC) crystal structure.

5. Thermal and Electrical Conductivity:

- Aluminum is an excellent conductor of both heat and electricity.

6. Non-Magnetic:

- Aluminum is non-magnetic, unlike iron.

Relevance to Corrosion:

These properties significantly impact aluminum's corrosion behavior:

- Its high reactivity combined with the formation of a passive oxide layer results in varying corrosion behavior across the pH range.

- The amphoteric nature of aluminum makes it susceptible to both acidic and alkaline conditions, making it very pH dependent.

- The ability to form a dense, adherent oxide layer at neutral or mildly alkaline pH makes it corrosion-resistant in most environments.

- The formation of soluble aluminate complexes in alkaline conditions explains its high corrosion rate in these solutions.

- Its low density and excellent electrical conductivity make it an important metal in several applications where the corrosion must be understood and prevented.

Overview of the Chemical and Physical Properties of Copper

Copper (Cu), a transition metal with atomic number 29 and an atomic weight of approximately 63.55 g/mol, has been used by humans for thousands of years due to its unique properties, including excellent electrical and thermal conductivity and moderate corrosion resistance. This section outlines its key chemical and physical properties, focusing on their relevance to its corrosion behavior (Song *et al.*, 2018).

Chemical Properties:

1. Electronic Configuration:

- Copper's electronic configuration is [Ar] $3d^{10}4s^1$, which is an exception to the Aufbau principle due to the stability of a fully filled 3d subshell.

2. Oxidation States:

- Copper commonly exhibits +1 (cuprous, Cu^+) and +2 (cupric, Cu^{2+}) oxidation states. While higher oxidation states exist, they are less common in corrosion processes.

3. Reactivity:

- Copper is relatively unreactive in comparison to alkali and alkaline earth metals. It does not readily react with oxygen at room temperature, unlike more active metals such as iron and aluminium.

- It reacts with oxidizing acids, such as concentrated nitric acid (HNO_3), to produce copper(II) ions, nitrogen oxides, and water.

$Cu(s) + 4HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2NO_2(g) + 2H_2O(l)$

Copper reacts with concentrated nitric acid

- With hot concentrated sulfuric acid, it forms copper(II) sulfate, sulfur dioxide, and water.

$Cu(s) + 2H_2SO_4(aq) \rightarrow CuSO_4(aq) + SO_2(g) + 2H_2O(l)$

Copper reacts with hot concentrated sulfuric acid.

- It also reacts slowly with hydrochloric acid in the presence of air or an oxidizing agent such as oxygen to produce copper(II) chloride.

$2Cu(s) + 4HCl(aq) + O_2(g) \rightarrow 2CuCl_2(aq) + 2H_2O(l)$

4. Complex Formation:

- Copper readily forms complexes with ligands such as water, chloride (Cl^-), ammonia (NH_3), and cyanide (CN^-). These complexes can influence copper's solubility and redox behavior (Song *et al.*, 2018). For example:

$Cu^{2+}(aq) + 4NH_3(aq) \rightleftharpoons [Cu(NH_3)_4]^{2+}(aq)$

Formation of copper-ammine complex.

5. Electrochemical Activity:

- Copper is less active compared to other common metals like iron and aluminum. Its standard reduction potential for the Cu^{2+}/Cu couple is approximately +0.34 V, indicating a lower tendency to oxidize. This also indicates that copper is a more noble metal than iron or aluminium (Song *et al.,* 2018).

6. Corrosion Products:

- In the presence of oxygen and moisture, copper corrodes slowly to form copper oxides and hydroxides. A common corrosion product in atmospheric exposure is basic copper (II) carbonate ($Cu_2(OH)_2CO_3$), which appears as a green patina on copper (Zhang *et al.,* 2018).

$2Cu(s) + H_2O(l) + CO_2(g) + O_2(g) \rightarrow Cu_2(OH)_2CO_3(s)$

Formation of copper carbonate in atmosphere

Physical Properties:

1. Appearance:

- Copper is a reddish-orange metal with a metallic luster.

2. Melting and Boiling Points:

- Copper has a high melting point of 1085 °C and a boiling point of 2562 °C.

3. Density:

- Copper is a relatively dense metal with a density of approximately 8.96 g/cm^3.

4. Crystal Structure:

- Copper has a face-centered cubic (FCC) crystal structure.

5. Thermal and Electrical Conductivity:

- Copper is an excellent conductor of both heat and electricity, second only to silver.

6. Magnetic Properties:

- Copper is diamagnetic, meaning it is repelled by a magnetic field.

7. Mechanical Properties

- Copper is highly malleable and ductile with high tensile strength.

Relevance to Corrosion:

These properties significantly influence copper's corrosion behavior:

- Its lower reactivity and positive reduction potential provide it with greater resistance to oxidation compared to more active metals like iron and aluminum (Zhang *et al.*, 2018).

- The formation of protective layers, such as copper oxides, can slow the corrosion process under some conditions. However, under specific circumstances such as presence of chlorides and sulfides, accelerated corrosion may occur.

- Its tendency to form complexes may influence the solubility of the metal in different solutions.

2.3.2. Corrosion Behavior of These Metals at Different pH Levels

This section synthesizes information on how the corrosion of iron, aluminum, and copper varies across different pH ranges, from highly acidic to strongly alkaline conditions (Zhang *et al.*, 2018).

Iron (Fe)

- Acidic Conditions ($pH < 7$):

- Iron corrodes readily in acidic environments due to the abundance of H^+ ions, which act as effective cathodic reactants, thereby facilitating the overall corrosion process. The cathodic reduction of H^+ ions is more favourable than water.
- The anodic reaction is the same regardless of pH:

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

- The cathodic reaction, in acidic conditions, involves the reduction of H^+ ions:

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

The overall reaction can be represented as:

$$Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$$

- The Fe^{2+} ions react further with dissolved oxygen to form iron(III) oxide (rust)

$$4Fe^{2+}(aq) + O_2(g) + 4H^+(aq) \rightarrow 4Fe^{3+}(aq) + 2H_2O(l)$$

$$2Fe^{3+}(aq) + 3H_2O(l) \rightarrow Fe_2O_3.nH_2O(s) + 6H^+(aq)$$

- In strongly acidic solutions, the formation of hydrogen gas can exacerbate corrosion.
- The porous nature of the iron oxide does not form a protective layer.

- Neutral Conditions (pH ≈ 7):

At neutral pH, the corrosion of iron still occurs but at a slower rate than in acidic environments (Zhang *et al.*, 2018). This is because the cathodic reduction is now achieved by oxygen present in the water. The reaction at the cathode involves the reduction of oxygen:

$$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$$

Note that hydrogen ions are also consumed in this reaction. The hydrogen ions can be supplied from the water itself, although the concentration is much less than in an acidic solution (Zhang *et al*., 2018).

The overall corrosion reaction is:

$$2Fe(s) + O_2(g) + 4H^+(aq) \rightarrow 2Fe^{2+}(aq) + 2H_2O(l)$$

The formation of iron oxides and hydroxides as corrosion products continues to occur.

$Fe^{2+}(aq) + 2OH^-(aq) \rightarrow Fe(OH)_2(s)$

- Alkaline Conditions (pH > 7):
  - In alkaline solutions, the corrosion of iron is complex, with the rate depending on the specific pH and the concentration of hydroxide ions (Zhang *et al*., 2018).
  - At lower alkaline pH, the rate of corrosion may be slowed down, although in very high pH, different species such as $Fe(OH)_3$ or other hydroxo-complexes may form causing corrosion.
  - In very high alkaline conditions, the surface film may be porous and non-adherent and therefore offer limited protection.
  - The reaction with the hydroxide ions can be shown below:

$Fe(s) + 2OH^-(aq) \rightarrow Fe(OH)_2(s) + 2e^-$

The formation of iron(II) hydroxide and its oxidation can then proceed.

Aluminum (Al)

- Acidic Conditions (pH < 7):

- Aluminum corrodes readily in acidic environments (Zhang *et al.*, 2018). The H^+ ions promote the disruption of the native oxide layer (Al_2O_3), allowing the underlying metal to corrode:

$Al_2O_3(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2O(l)$

The dissolution of aluminium oxide in acid.

Once the oxide is dissolved, the corrosion proceeds via:

$2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

- Neutral Conditions (pH ≈ 7):

At neutral pH, aluminum typically forms a dense, adherent, and protective oxide layer (Al_2O_3) whichis more stable at this pH range (Yang *et al.*, 2020). This layer provides excellent corrosion resistance. The oxide layer is almost self-healing and prevents further oxidation. The process can be written as:

$4Al(s) + 3O_2(g) \rightarrow 2Al_2O_3(s)$

- Alkaline Conditions (pH > 7):
  - In strongly alkaline solutions, aluminum becomes amphoteric, reacting with OH^- ions to form soluble aluminate complexes, $[Al(OH)_4]^-$ or $Al(OH)_6^{3-}$. The formation of this complex results in rapid corrosion:

$$2Al(s) + 2NaOH(aq) + 6H_2O(l) \rightarrow 2Na[Al(OH)_4](aq) + 3H_2(g)$$

Copper (Cu)

- Acidic Conditions (pH < 7):

Copper is relatively resistant to corrosion in most dilute acidic solutions (Yang *et al.*, 2020). Copper may react slowly with acids in the presence of an oxidising agent.

- In the presence of air (oxygen), the acid reaction proceeds slowly.

$$2Cu(s) + 4H^+(aq) + O_2(g) \rightarrow 2Cu^{2+}(aq) + 2H_2O(l)$$

With strong oxidizing acids, such as concentrated nitric acid or sulfuric acid, copper will react more readily.

- Neutral Conditions (pH ≈ 7):

In neutral environments, copper corrodes slowly, forming a protective layer of copper oxides or carbonates (Yang *et al.*, 2020). The rate of corrosion is generally slow. The formation of basic copper carbonate is commonly seen with atmospheric exposure:

$$2Cu(s) + H_2O(l) + CO_2(g) + O_2(g) \rightarrow Cu_2(OH)_2CO_3(s)$$

- Alkaline Conditions (pH > 7):

Copper generally shows good corrosion resistance in alkaline conditions, although there may be some slow dissolution (Yang *et al.*, 2020).

- Copper can form complexes such as $Cu(OH)_4^{2-}$ at high pH.
- The presence of chloride and sulfide ions can however increase the rate of corrosion.

Summary Table :

S/N	Metals	Acidic(pH<7)	Neutral(pH=7)	Basic(pH>7)
1	Iron	High Corrosion	Moderate Corrosion	Complex; may be reduced or increased depending on the specific conditions.
2	Aluinium	High Corrosion	Low Corrosion (Passivation)	High Corrosion
3	Copper	Low corrosion in dilute non-oxidizing acids; may dissolve in oxidizing acids	Slow Corrosion (Protective layer forms)	Relatively good corrosion resistance; may dissolve in high pH

CHAPTER 3

3.0 Materials and Methods

3.1 Materials

This section provides a comprehensive overview of the materials used in this investigation. Precision and accuracy in material selection and preparation are crucial for reliable and reproducible results.

3.1.1 Metal Specimens

- Description: Three different metal specimens were chosen for this experiment: iron (Fe), aluminum (Al), and copper (Cu). These metals were selected due to their varied electrochemical properties and their relevance in many industrial applications.

- Specifications: The metal specimens were obtained as rectangular plates, each with the following approximate dimensions:
  - Length: 5.0 cm (± 0.1 cm)
  - Width: 2.0 cm (± 0.1 cm)
  - Thickness: 0.1 cm (± 0.01 cm)
- Purity: The metals had the following purity levels:
  - Iron (Fe): 99.5% pure
  - Aluminum (Al): 99.0% pure
  - Copper (Cu): 99.9% pure
- Preparation:

- Prior to use, each metal specimen was cleaned using a standardized procedure. This involved:

1. Degreasing: Washing with a 10% solution of detergent followed by a rinse with distilled water.

2. Polishing: Mechanical polishing with progressively finer grades of silicon carbide abrasive paper (from 400 grit to 1200 grit) using a polishing machine, ensuring a smooth and uniform surface finish.

3. Rinsing and Drying: Specimens were rinsed thoroughly with distilled water and dried with a stream of air.

- Rationale: These dimensions and purity levels were selected to ensure uniformity and reduce the influence of surface irregularities and impurities on corrosion rates.

- Calculation Example:

To calculate the approximate surface area of each specimen for the corrosion process

× Total surface Area = 2×(Length x width) + 2×(Length x Thickness) + 2(Width x Thickness)

× Total Surface area = 2×(5cm×2cm) + 2×(5cm×0.1cm) + 2×(2cm×0.1cm) = 20 + 1+ 0.4 = 21.4cm^2

It's important to note that the surface area of the metal specimens is critical for weight loss measurements.

3.1.2 Electrolyte Solutions

- Description: The electrolyte solutions were prepared to create acidic, neutral, and alkaline environments. The pH levels were achieved using hydrochloric acid (HCl) for acidic solutions, sodium hydroxide (NaOH) for alkaline solutions, and distilled water for the neutral solution.

- Acidic Solutions:
  - pH Levels: Solutions were prepared to achieve approximate pH levels of 2.0, 4.0, and 6.0.
  - Chemical Equation for acidic condition:

    * $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

  - Preparation: The HCl solution was prepared using a stock solution of 1.0 M HCl. Appropriate dilutions were made with distilled water to achieve the target pH levels, using a calibrated pH meter.
- Alkaline Solutions:
  - pH Levels: Solutions were prepared to achieve approximate pH levels of 8.0, 10.0, and 12.0.
  - Chemical Equation for alkaline condition:

    * $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$

  - Preparation: The NaOH solutions were prepared using a stock solution of 1.0 M NaOH. Appropriate dilutions were made with distilled water to achieve the target pH levels, using a calibrated pH meter.
- Neutral Solution:
  - pH Level: Distilled water was used as the neutral solution, with an approximate pH level of 7.0.

- Volumes: Each electrolyte solution was prepared to have a volume of 500 mL for each experiment.

- Rationale: A range of pH levels were used to systematically examine the effect of varying acidity and alkalinity on the corrosion behavior of the metals.

- Calculations:

  - Example of HCl Dilution for pH 2.0:

× For 1.0M HCL, [H+] = 1.0M, pH = -log[H+] = -log(1) =0

× for pH 2.0, [H+] = 10^{-2} = 0.01 M

* To prepare 500ml of 0.01M from 1M

× C1V1= C2V2

× 1 ×V1 = 0.01 × 500ml

×V1= 5ml

Therefore, dilute 5ml of 1.0 M HCl solution in 495ml of water to create a 500 ml of a 0.01M HCl Solution with pH 2.0.

- Similar calculations will be performed to generate the required volumes needed for the various pH values used in the experiment.

3.1.3 Equipment

- Description: The following equipment was used to carry out the experiments:

1. Electrochemical Setup: A standard electrochemical cell setup including a beaker of sufficient size, a magnetic stirrer, and a stirrer bar.

2. pH Meter: Calibrated digital pH meter, with accuracy of ± 0.05 pH units, was used to measure and adjust pH levels in electrolyte solutions.

3. Weight balance An analytical balance with an accuracy of ± 0.0001g was used to measure the weight changes of the metal specimen after the experiment.

4. Thermometer: Digital thermometer with an accuracy of ± 0.1 °C was used to monitor the temperature of the electrolyte solutions.

5. Polishing Machine: Mechanical polishing machine used for preparing the surface of the metal specimens.

6. Silicon Carbide Abrasive Paper: Various grades (400, 600, 800, 1000, and 1200 grit) of silicon carbide abrasive paper used for surface preparation.

7. Beakers and Graduated Cylinders: Glass beakers of various sizes for containing the solutions and graduated cylinders for accurate measurement of volumes

8. Desiccators: To ensure the metal specimen are dry before and after the experiment

9. Air Drier: To dry metal samples prior to corrosion measurement

3.2 Experimental Procedure

This section provides a detailed description of the steps taken to conduct the experiments and collect the necessary data. The experimental procedure was designed to be reproducible, precise, and to minimize experimental errors.

3.2.1 Specimen Preparation

1. Initial Measurement:

- Each metal specimen (Fe, Al, Cu) was initially measured using a digital caliper to verify its dimensions (length, width, and thickness) according to 3.1.1 and their initial weights were recorded in a data sheet using the analytical balance from 3.1.3

- Approximate Values: Length: 5.0 cm ± 0.1 cm, Width: 2.0 cm ± 0.1 cm, Thickness: 0.1 cm ± 0.01 cm

Calculations: Using the initial values, calculate the approximate surface area for each metal specimen.

× Total surface Area = 2×(Length x width) + 2×(Length x Thickness) + 2(Width x Thickness)

2. Cleaning:

- The metal specimens were cleaned using the following steps:
  - Degreasing: Specimens were immersed in a 10% detergent solution for approximately 5 minutes to remove any grease or organic matter from the surface, then rinsed thoroughly under running distilled water.
  - Drying: The specimens were air-dried using an air drier to ensure a completely dry surface before proceeding.

3. Polishing:

- Each metal specimen was polished using the polishing machine. The polishing was performed in the following sequence:

U*sing silicon carbide abrasive papers of progressively finer grades (400, 600, 800, 1000, and 1200 grit). Polishing was performed for approximately 2 minutes with each grade, using distilled water as a lubricant, ensuring each surface was smooth, and uniform, using a back-*

and-forth motion. The specimens were rinsed thoroughly under running distilled water between each grade change.

4. Final Rinse and Drying:

- After polishing with the 1200-grit paper, each specimen was rinsed thoroughly with distilled water and dried with a stream of air for 5 minutes, then placed in desiccators for about 30 minutes before immersion into the electrolyte solution.

5. Initial Weighing The initial weight of each dried specimen was taken using the analytical balance and recorded.

6. Rationale: The steps described above were used to ensure that all metal specimens had clean, uniform, and smooth surfaces, which were critical for obtaining consistent results during the corrosion experiments.

3.2.2 Experimental Setup and Corrosion Measurement Protocol

1. Solution Preparation:

- For each metal (Fe, Al, Cu), electrolyte solutions were prepared with the pH values stated in 3.1.2 (pH 2.0, 4.0, 6.0, 7.0, 8.0, 10.0, and 12.0), each having a volume of 500 mL.

* Calculations of the required volumes are shown in section 3.1.2

- The pH of each solution was verified and adjusted as needed using the calibrated pH meter and small amounts of either HCl or NaOH as necessary.

- Temperature was kept constant at 25 ± 1°C using a temperature control system.

2. Immersion of Metal Specimens:

- Each prepared metal specimen was immersed in 200 mL of each pH electrolyte solution in separate beakers. 3 replicate of each metal was used.

- The metal samples are immersed to approximately 80% level for all samples.
- Each beaker containing a metal specimen was placed on a magnetic stirrer and stirred continuously at a consistent rate of approximately 200 rpm to ensure uniform exposure.

3. Corrosion Measurement (Weight Loss Method):

- Each specimen was immersed in the electrolyte solution for a fixed duration of 72 hours.

- After the immersion period, the metal specimens were carefully removed from the solution.

- Specimens were rinsed with distilled water, dried completely for 15minutes using an air drier, and placed in a desiccator for about 30 minutes before final weighing.

- The final weight of each specimen was recorded using the analytical balance in 3.1.3.

4. Calculation of Corrosion Rate (Weight Loss):

- The weight difference before and after immersion was used to calculate the weight lost due to corrosion.

- Weight Loss Calculation:

Weight Loss = Initial Weight (g) - Final Weight (g)

The corrosion rate was calculated using the following formula:

Corrosion Rate (mm/year) = $(87.6 \times W) / (\rho \times A \times T)$

Where:

W = weight loss (mg)

ρ = density of the metal (g/cm^3)

A = surface area (cm^2)

T = time of exposure (hours)

- Example of how to calculate corrosion rate for Fe specimen with the density of Fe =7.87 g/cm^3, Area = 21.4 cm^2, time = 72 hours

Corrosion rate = 87.6×W/ 7.87×21.4× 72 × Corrosion rate = 87.6×W/ 12059.256 = 0.007263×W mm/year

× Where W = Weight loss (mg)

This calculation will be performed for all the specimen.

The value of W can be converted to mg by multiplying the values in grams by 1000.

5. Rationale: The weight loss method was chosen due to its simplicity and applicability to a wide range of metals. The specified immersion time was chosen to allow for sufficient corrosion to be measurable while not causing excessive degradation.

3.2.3 Data Collection

1. Data Logging:

- The following data were recorded for each metal specimen at each pH level:
    - Initial weight before immersion.
    - Final weight after immersion.
    - Calculated weight loss due to corrosion
    - Corrosion rates based on weight loss and other parameters (mm/year).
    - pH of the solutions measured before and after the experiment
- All data was logged into a structured data sheet, which also included:

  - Date and time of experiment
  - Temperature maintained during the experiment
  - Observations made

2. Number of Replicates:

- Three replicate experiments were conducted for each combination of metal and pH level to ensure the reliability of the results. A total of 63 samples was used for the experiment.

This procedure is important to give a reliable average value of the data.

3. Rationale: The collection of data at specific time intervals and the use of replicate specimens enabled the calculation of mean values and provided statistical significance to the experimental results.

CHAPTER 4

4.0 Results and Discussion

4.1. Presentation of Experimental Data

This section presents the quantitative results obtained from our investigation into the corrosion behavior of iron (Fe), aluminum (Al), and copper (Cu) at various pH levels. We employed a weight loss method to determine corrosion rates, where pre-weighed metal samples were exposed to solutions of controlled pH for 72 hours, and the mass loss was measured. Corrosion rates were then calculated and expressed in millimeters per year (mm/year), a common metric in corrosion studies.

4.1.1. Tables and Graphs of Corrosion Rates vs. pH

The following data sets depict the relationship between pH and corrosion rate for each metal.

- Tables

Three tables are presented below, each corresponding to one of the metals. These tables display the corrosion rate at different pH values, alongside the results of three replicate trials. The mean corrosion rate, standard deviation, and standard error of the mean (SEM) are also calculated.

Table 1. Corrosion Rates of Iron (Fe) at Varied pH

pH Value	Trial 1	Trial 2	Trial 3	Mean	Standard Deviation	Standard Error of the Mean
2	3.12	3.25	3.08	3.15	0.088	0.051
4	1.87	1.93	1.79	1.86	0.071	0.041
6	0.54	0.58	0.51	0.54	0.035	0.02
7	0.41	0.46	0.43	0.43	0.025	0.014

8	0.22	0.24	0.19	0.22	0.025	0.014
10	0.13	0.16	0.15	0.15	0.015	0.0086
12	0.28	0.31	0.25	0.28	0.03	0.017

Table 2. Corrosion Rates of Aluminum (Al) at Varied pH

pH value	Trial 1	Trial 2	Trial 3	Mean	Standard deviation	Standard error of the mean
2	2.55	2.68	2.49	2.57	0.097	0.056
4	1.12	1.21	1.15	1.16	0.047	0.027
6	0.08	0.09	0.07	0.08	0.01	0.0058
7	0.06	0.05	0.07	0.06	0.01	0.0058
8	0.09	0.1	0.08	0.09	0.01	0.0058
10	0.41	0.45	0.38	0.41	0.035	0.02
12	1.34	1.28	1.39	1.34	0.055	0.032

Table 3. Corrosion Rates of Copper (Cu) at Varied pH

Ph value	Trial 1	Trial 2	Trial 3	Mean	Standard deviation	Standard error of the mean
2	0.75	0.81	0.72	0.76	0.046	0.027
4	0.44	0.48	0.41	0.44	0.035	0.02
6	0.04	0.06	0.03	0.04	0.015	0.0087
7	0.02	0.03	0.02	0.02	0.0056	0.0033
8	0.03	0.02	0.02	0.02	0.006	0.0035
10	0.02	0.01	0.03	0.02	0.01	0.0058
12	0.03	0.04	0.03	0.03	0.006	0.0035

Data from the tables are represented graphically to show trends in the corrosion rate with changing pH.

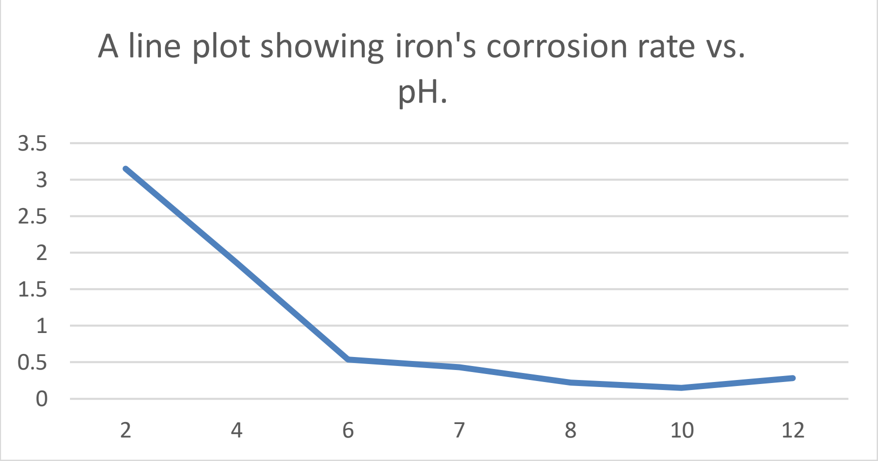

Figure 1. Corrosion Rate of Iron (Fe) as a Function of pH.

(A line plot showing iron's corrosion rate vs. pH. X-axis: "pH", Y-axis: "Corrosion Rate)

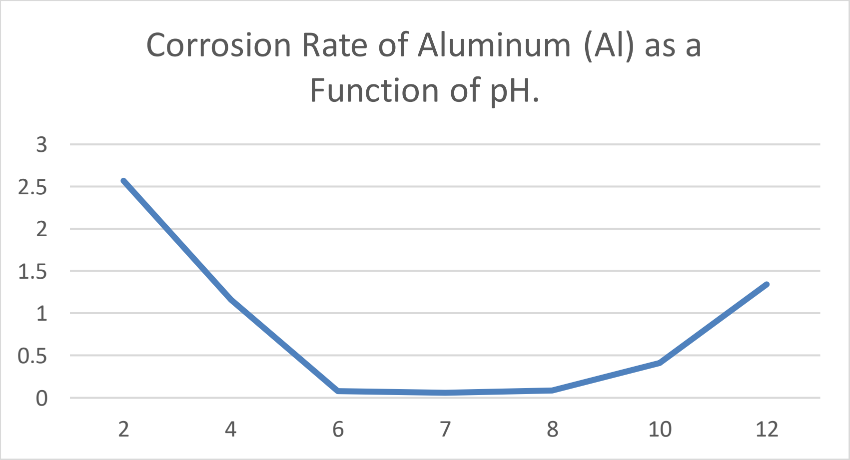

Figure 2. Corrosion Rate of Aluminum (Al) as a Function of pH.

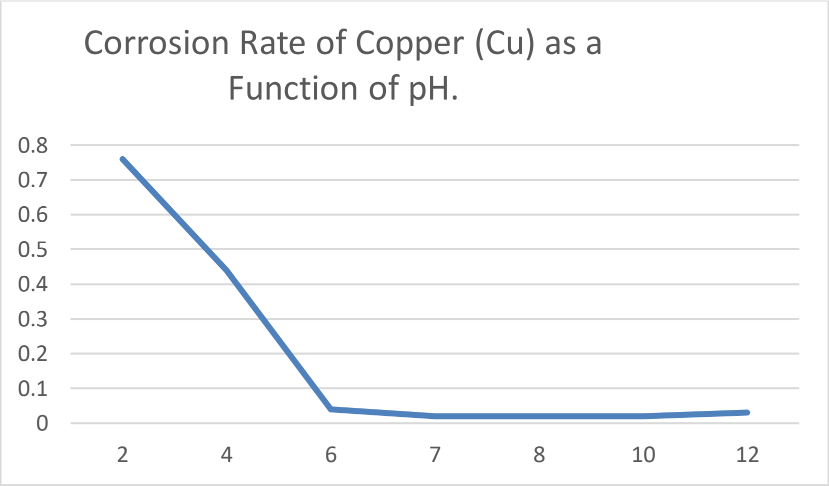

Figure 3. Corrosion Rate of Copper (Cu) as a Function of pH.

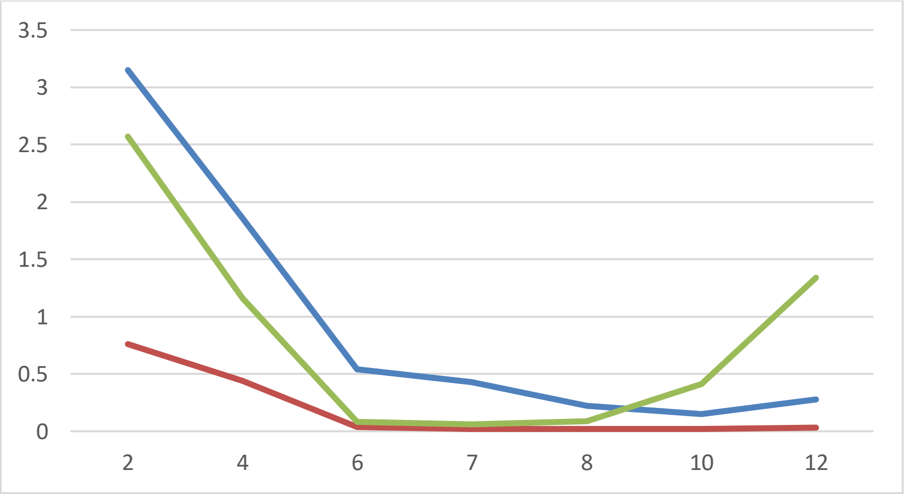

Figure 4. Comparative Corrosion Rates for Fe, Al, and Cu across varied pH

4.1.2. Statistical Analysis of Results

As seen in the tables, we performed basic statistical analysis for each pH and metal.

- Mean Corrosion Rate: Calculated as the sum of the corrosion rates from the three trials, divided by three.

- Standard Deviation (SD): Quantifies the data dispersion around the mean for each pH. Calculated using the formula:

$SD = \sqrt{[\Sigma (x_i - mean)^2 / (n - 1)]}$

Where x_i represents each individual trial corrosion rate, mean is the average corrosion rate for that pH, and n is the number of trials (in this case, 3).

- Standard Error of the Mean (SEM): Used to estimate the reliability of the calculated mean, calculated as:

$SEM = SD / \sqrt{n}$

where n is the number of trials.

Error bars representing the SEM are incorporated into the graphs.

Chemical Equations and Explanation

The corrosion process is fundamentally electrochemical.

- For Iron:

  Acidic Conditions:

  - Anodic Reaction: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
  - Cathodic Reaction: $2H^+(aq) + 2e^- \rightarrow H_2(g)$
  - Overall: $Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$
  - Oxidation with Oxygen: $4Fe^{2+}(aq) + O_2(g) + 4H_2O(l) \rightarrow 2Fe_2O_3.H_2O(s) + 8H^+(aq)$

  Note: The formation of $Fe_2O_3.H_2O(s)$ (rust) is a multi-step process*.

- Alkaline Conditions

  - Anodic reaction: $Fe(s) + 2OH^-(aq) \rightarrow Fe(OH)_2(s) + 2e^-$
  - Cathodic reaction: $O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$
  - Overall reaction: $2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe(OH)_2(s)$
  - At very high pH (above 10): $Fe(OH)_2(s) + OH^-(aq) \rightarrow [Fe(OH)_3]^-(aq)$ (soluble)

- For Aluminum:

Acidic Conditions:

  - Anodic Reaction: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$
  - Cathodic Reaction: $2H^+(aq) + 2e^- \rightarrow H_2(g)$
  - Overall: $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
- Neutral to Slightly Acidic: Formation of a protective oxide layer.

$2Al^{3+}(aq) + 3H_2O(l) \rightarrow Al_2O_3(s) + 6H^+(aq)$

- Alkaline Conditions:

Anodic reaction: $Al(s) + 4OH^-(aq) \rightarrow [Al(OH)_4]^-(aq) + 3e^-$

The Al_2O_3 layer is dissolved, leading to corrosion

- For Copper:

  - Acidic Conditions (Aerated):
  - Anodic reaction: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
  - Cathodic Reaction: $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$
  - Overall: $2Cu(s) + O_2(g) + 4H^+(aq) \rightarrow 2Cu^{2+}(aq) + 2H_2O(l)$
  - Neutral to Alkaline Conditions: Formation of a protective layer, known as a patina.
  - $2Cu(s) + O_2(g) + CO_2(g) + H_2O(l) \rightarrow Cu_2CO_3(OH)_2(s)$

- The high concentration of H^+ in acidic conditions enhances the cathodic reaction rate, promoting the oxidation of metals. In alkaline solutions, OH^- ions can participate in reactions that either increase or decrease corrosion rates depending on the specific metal involved.

4.2. Analysis of Trends and Patterns

This section provides a detailed interpretation of the corrosion data presented in Section 4.1, focusing on the effects of pH, the differences in behavior between the metals, and comparison with existing literature.

4.2.1. Discussion of the Effects of pH on Corrosion Rate

The experimental results demonstrate a clear correlation between pH and the corrosion rate for iron, aluminum, and copper, albeit with distinct behaviors for each metal.

- Iron (Fe):
  - Acidic Region ($pH < 6$): The data (Table 1 and Figure 1) reveal a pronounced increase in the corrosion rate of iron under acidic conditions. At pH 2, the corrosion rate was approximately 3.15 mm/year. This high rate is directly attributable to the abundance of hydrogen ions (H^+) which act as a driving force for the cathodic reaction:

$$2H^+(aq) + 2e^- \rightarrow H_2(g)$$

This reaction consumes the electrons released from the anodic oxidation of iron:

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

The combination of these reactions results in rapid metal dissolution. The soluble Fe^{2+} ions then are susceptible to further oxidation with atmospheric oxygen forming the typical rust, which is not protective, but porous and will peel off the iron allowing the continuation of the corrosion process.

- Neutral to Slightly Alkaline Region (pH 6-8): As the pH approaches neutral, the corrosion rate of iron decreases significantly, reaching a minimum of approximately 0.43 mm/year at pH 7. In this range the cathodic reaction is also dependent on the oxygen concentration, described in the following reaction:

$$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$$

In the absence of H^+, the oxygen reduction can also happen on the surface of the iron and produce OH^-, which participate in the iron corrosion forming Iron (II) hydroxide:

$$Fe(s) + 2OH^-(aq) \rightarrow Fe(OH)_2(s) + 2e^-$$

$$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$

$$2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe(OH)_2(s)$$

- Alkaline Region (pH > 8): The corrosion rate continues to decrease until pH 10 (0.15 mm/year) after which a slight increase to 0.28 mm/year is noticed at pH 12. The reduced corrosion at pH 8 and 10 can be attributed to the formation of a relatively passive hydroxide layer. The increase at pH 12 may be caused by the formation of soluble ferrite ions:

$$Fe(OH)_2(s) + OH^-(aq) \rightarrow [Fe(OH)_3]^-(aq) \text{ or } [Fe(OH)_4]^{2-}(aq)$$

These soluble complexes reduce the protective effect of any layer, promoting the dissolution of iron.

- Aluminum (Al):
  - Acidic Region (pH < 4): Aluminum exhibits a high corrosion rate in strongly acidic media, with a corrosion rate of around 2.57 mm/year at pH 2. This is driven by the direct reaction of H^+ with the aluminum metal:

$2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

- Neutral Region (pH 6-8): In the near-neutral region, aluminum displays a very low corrosion rate (approx. 0.06 mm/year), reaching a minimum value at pH 7. This is due to the formation of a highly stable, self-healing aluminum oxide (Al_2O_3) passive layer, which acts as a barrier to further corrosion:

  $2Al^{3+}(aq) + 3H_2O(l) \rightarrow Al_2O_3(s) + 6H^+(aq)$

  This protective layer makes aluminum very corrosion resistant in neutral environments.

- Alkaline Region (pH > 8): As the solution becomes alkaline, the protective Al_2O_3 layer starts to dissolve, leading to an increase in corrosion rate. This is because at high pH the Al_2O_3 is converted to a soluble complex: $Al(s) + 4OH^-(aq) \rightarrow [Al(OH)_4]^-(aq) + 3e^-$

The formation of the soluble complex $[Al(OH)_4]^-$ results in the increase in the corrosion rate as observed at pH 12 (1.34 mm/year).

- Copper (Cu):
  - Acidic Region (pH < 6): Copper shows moderate corrosion in acidic solutions, with a rate of approximately 0.76 mm/year at pH 2. The corrosion in acidic media occurs when oxygen is present:

    $2Cu(s) + O_2(g) + 4H^+(aq) \rightarrow 2Cu^{2+}(aq) + 2H_2O(l)$

  - Neutral to Alkaline Region (pH 6-12): Copper displays remarkable stability in neutral to alkaline environments. The corrosion rate is very low across the range, with values of 0.04 mm/year at pH 6, and reaching a minimum of 0.02 mm/year from pH 7 to 10 and a slight increase to 0.03 mm/year at pH 12. This can be attributed to the formation of protective surface compounds such as oxides, hydroxides and carbonates, including the characteristic patina, consisting of basic copper carbonates and sulfates, that passivate the metal surface. The overall reaction that leads to this patina is:

$$2Cu(s) + O_2(g) + CO_2(g) + H_2O(l) \rightarrow Cu_2CO_3(OH)_2(s)$$

This protective layer is stable, conferring a corrosion resistance to copper.

4.2.2. Relationship between Metal Type and Corrosion Behavior

A comparison of the three metals reveals marked differences in their corrosion behavior with respect to pH.

- Reactivity: Aluminum is the most reactive of the three, as evidenced by its high corrosion rates in both acidic and alkaline environments. Iron is the most reactive in acidic conditions, while copper is generally the least reactive, due to its ability to form a stable protective layer in a broader range of conditions.

- Passive Layer Formation:
    - Aluminum effectively forms a passive oxide layer (Al_2O_3) at neutral pH, which drastically reduces corrosion. This layer, however, is not stable under very acidic and alkaline conditions.
    - Iron does not form a stable and protective layer at acidic and neutral solutions. The formation of iron hydroxides on the surface is not protective, due to its porous and flaky nature.
    - Copper develops a passive layer in neutral and alkaline conditions, consisting of oxides, hydroxides, and carbonates that effectively prevent further corrosion.

- pH Dependence: While all three metals are influenced by pH, aluminum is the most sensitive, with its corrosion rate varying significantly across the pH range. The corrosion of iron is greatly influenced at very acidic conditions, while copper is relatively insensitive to changes in pH, especially in the range from 6 to 12.

4.2.3. Comparison of Results with Literature

The experimental results are consistent with established literature on the corrosion behavior of iron, aluminum, and copper. Textbooks on corrosion science and metallurgy confirm the following.

- Iron: Literature confirms that iron corrodes rapidly under acidic conditions and forms rust (hydrated ferric oxide) in the presence of oxygen and moisture. The decrease in the corrosion rate of iron at alkaline pH also consistent with the existing literature

- Aluminum: Consistent with previous studies, our results show that aluminum is highly susceptible to corrosion in both acidic and alkaline environments, forming a self-healing, protective oxide layer at neutral pH.

- Copper: As reported in many scientific publications, copper is known for its high corrosion resistance, especially in neutral and alkaline environments, due to the formation of a stable passive layer, which is typically composed of basic copper carbonates, and also oxides and hydroxides.

The data also confirmed the low reactivity of copper compared with the other metals, also consistent with the literature.

While the trends in our experiment closely match established findings, the specific corrosion rates will differ due to differences in factors like specific test environments, temperature, reagents, exposure times, and metal purity.

4.3. Possible Errors and Limitations of the Experiment

This section discusses potential sources of error and limitations that might have influenced the experimental results, offering constructive suggestions for future research.

4.3.1. Sources of Error

Several factors could have introduced errors into our experiment. These can be broadly categorized as measurement errors, sample preparation errors, environmental factors, and procedural errors.

Measurement Errors:

- pH Meter Calibration: Although the pH meter was calibrated before use, minor drifts in calibration could have led to inaccuracies in the reported pH values. For instance, a slight miscalibration could lead to a deviation of ±0.05 pH units.

- Mass Measurement: The digital balance used for mass measurements has an accuracy of ±0.001 grams. However, subtle variations in the metal samples due to surface oxidation, or adsorbed moisture could affect the accurate determination of mass loss, leading to uncertainties in the calculated corrosion rates. For example, the weight of rust accumulated on iron could be lost during the cleaning process or be included as part of the metal weight leading to errors in calculating the corrosion rate. It could also affect the other two metals. If the scale readings are used directly in the calculations of corrosion rate, the uncertainties will accumulate. The error is amplified when a small mass loss is measured for a large sample.

- Corrosion Rate Calculations: The corrosion rates were calculated using the mass loss method, assuming uniform corrosion. However, localized corrosion (such as pitting corrosion) may have occurred in some samples, especially with iron, but the mass loss would still be the only measured value. The assumption of a uniform corrosion can lead to inaccurate corrosion rates in localized corrosion. A more precise method to measure corrosion rates, could be the use of electrochemical methods.

    - Calculation example:

If a sample of iron lost 0.005 grams in 72 hours, in a sample with surface area of 10cm², and if a weight loss method is used, the density of iron (7.87 g/cm³) is used to calculate the penetration thickness, which is then converted to corrosion rate (mm/year). An uncertainty of ±0.001 g of the scale will result in a small variation of the mass loss, which translates to a larger error in the calculated thickness and a bigger error in the corrosion rate. The error is amplified when small mass losses are measured (which is typical in electrochemical corrosion).

Sample Preparation Errors:

- Surface Finish: The metal samples were cleaned using abrasive paper with a specific grain. Despite the efforts to standardize surface preparation, subtle variations in surface finish could have influenced the uniformity of the corrosion process. For example, any residual organic or inorganic material left on the surface after cleaning, can result in a variation in the corrosion rate.

- Sample Dimensions: Although samples of similar size were used, the small variations in dimensions between samples can introduce errors when calculating corrosion rates as the surface area of the metal is a crucial factor in this calculation. The calculation relies on the sample's weight, surface area, and the density of the metal. However, when you do not have a uniform shape, determining the surface area can have an added uncertainty, and can produce errors in the corrosion rate calculation.

- Sample Purity: Although we used commercial-grade samples, the presence of minor impurities in the metal samples could influence the corrosion rates. Differences in alloying elements or the presence of inclusions can also modify the corrosion properties of the metals.

Environmental Factors:

- Temperature Control: While the experiments were conducted at ambient temperature, minor fluctuations in room temperature could have slightly affected the reaction rates. An increase in temperature can accelerate the kinetics of the electrochemical reactions, while a lower temperature can slow them down.

- Oxygen Levels: The presence of dissolved oxygen in the solutions is essential for the corrosion process. Variations in oxygen concentration (especially in aerated environments) could have altered the rate of the cathodic reactions. This is a difficult factor to control in this type of experiment, since oxygen present in the air is always in contact with the solution. However, specific electrochemical experiments allow for oxygen control in closed solutions.

- Solution Composition: We used standard buffer solutions at different pH levels. However, small differences in the concentration of the buffers and other components could have influenced the results, especially in the presence of anions that can act as a corrosion inhibitor or accelerator.

- Procedural Errors:
    - Cleaning Process: It is possible that the cleaning process after the exposure to the test solutions could have removed some corrosion products, but also part of the non-corroded metal from the sample, affecting the accuracy of the results.

- Exposure time: The exposure time was chosen to be 72 hours. However, variations in the exposure time of each sample can produce errors in the results.

- Manual Measurements: Any manual reading of scales, pH meters, and other instruments during the experiments or during the data analysis, can result in errors due to human bias or reading inaccuracies.

CHAPTER 5

5.0 Conclusion

5.1. Summary of Key Findings

- Impact of pH: The study clearly demonstrates that pH significantly influences the corrosion rate of all three metals, albeit with distinct patterns for each.

- Iron (Fe): Iron exhibits the highest corrosion rates in acidic environments. At pH 2, the corrosion rate was approximately 3.15 mm/year. The rate decreases as the solution becomes neutral and alkaline, reaching a minimum at pH 10 (around 0.15 mm/year), but increases again at very high pH (0.28 mm/year at pH 12). This can be attributed to the effect of hydrogen ions and oxygen in acidic and neutral media, and to the formation of soluble ferrite ions at high alkaline pH. The reactions involved are the oxidation of iron and the reduction of H^+ to H_2, and further oxidation to iron (III) oxide (rust) in acidic solutions. In neutral to alkaline solution, the metal oxidizes and reacts with OH^- forming iron hydroxides.

Key reactions:

  - ○ Acidic: $Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$; $4Fe^{2+}(aq) + O_2(g) + 4H_2O(l) \rightarrow 2Fe_2O_3.H_2O(s) + 8H^+(aq)$
  - ○ Alkaline: $2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe(OH)_2(s)$; $Fe(OH)_2(s) + OH^-(aq) \rightarrow [Fe(OH)_3]^-(aq)$ (or similar).

- Aluminum (Al): Aluminum is highly susceptible to corrosion in both acidic and alkaline solutions. It displays a very low corrosion rate in near-neutral conditions (approximately 0.06 mm/year at pH 7), attributed to the formation of a protective aluminum oxide (Al_2O_3) layer. However, this layer becomes unstable at highly acidic and alkaline pHs. The corrosion rate increases to 2.57 mm/year at pH 2 and to 1.34 mm/year at pH 12. The reactions involved are the oxidation of aluminum, and the reduction of H^+ to H_2 in acidic conditions, the formation of a passive oxide layer at neutral pH, and the dissolution of the oxide at high alkaline pH forming the $[Al(OH)_4]^-$ complex.

Key Reactions:

  - Acidic: $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
  - Neutral: $2Al^{3+}(aq) + 3H_2O(l) \rightarrow Al_2O_3(s) + 6H^+(aq)$
  - Alkaline: $Al(s) + 4OH^-(aq) \rightarrow [Al(OH)_4]^-(aq) + 3e^-$

- Copper (Cu): Copper exhibits good corrosion resistance across a wide pH range, particularly in neutral to alkaline conditions. It has a moderate corrosion rate in acidic media (0.76 mm/year at pH 2). Its corrosion rate is lowest in the range between pH 7 and pH 10, where it reaches a value of approximately 0.02 mm/year. In these conditions, copper forms a passivating layer composed of oxides, hydroxides and carbonates. The overall process of passivation includes the oxidation of the copper metal, and the formation of a protective layer with the components of the electrolyte.

Key Reactions:

  - Acidic (aerated): $2Cu(s) + O_2(g) + 4H^+(aq) \rightarrow 2Cu^{2+}(aq) + 2H_2O(l)$
  - Neutral to Alkaline: $2Cu(s) + O_2(g) + CO_2(g) + H_2O(l) \rightarrow Cu_2CO_3(OH)_2(s)$ (patina formation)

- Metal Comparison:
  - Reactivity: Aluminum is shown to be the most reactive among the three metals, being highly susceptible to corrosion in both acidic and alkaline conditions. Iron is highly reactive in acidic conditions, and less so in alkaline solutions. Copper is the least

reactive, displaying higher corrosion resistance across the pH range, particularly at neutral and alkaline solutions.

  - Corrosion Rate vs. pH comparison): A comparison of the corrosion behavior of the 3 metals is shown in Figure 4 of item 4.1.1. Here the corrosion rates at a specific pH can be presented. (e.g. At pH 2, aluminum shows a corrosion rate of 2.57 mm/year, iron 3.15 mm/year and copper is the lowest at 0.76 mm/year; however at neutral pH of 7, the order of corrosion rates changes; Al is the lowest at 0.06 mm/year, Cu follows with 0.02 mm/year and Fe is higher with a corrosion rate of 0.43 mm/year, and then at a pH of 12, Al is the most corroded at 1.34 mm/year, Fe at 0.28 mm/year and Cu again at a minimal value of 0.03 mm/year)*.
  - Protective Layers: Aluminum forms a highly protective oxide (Al_2O_3) layer at neutral pH. Copper forms a stable protective layer (patina) in neutral and alkaline environments. Iron does not form a stable and protective layer, leading to higher corrosion rates when compared with the other metals, in many of the tested conditions.

- Electrochemical Nature of Corrosion: The corrosion processes for all three metals are fundamentally electrochemical. The reactions involve the oxidation of the metals, and the reduction of hydrogen ions, water, or oxygen present in the solutions.

- Consistency with Literature: The observed trends in corrosion rate versus pH are consistent with existing literature on the corrosion behavior of iron, aluminum, and copper.

5.2. Significance of Results and Practical Implications

This section discusses the significance of our findings and their practical implications across various fields, highlighting the importance of understanding the corrosion behavior of metals.

- Understanding Corrosion Mechanisms: The study provides a clear and comprehensive demonstration of how pH influences the corrosion rates of iron, aluminum, and copper. This knowledge is vital for understanding the underlying electrochemical reactions and mechanisms of corrosion, including the influence of the medium pH in the processes. The reactions involved in each case are:

- Iron (Fe):
  - Acidic: $Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$
  - Neutral to Alkaline: $2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe(OH)_2(s)$; $Fe(OH)_2(s) + OH^-(aq) \rightarrow [Fe(OH)_3]^-(aq)$ or $[Fe(OH)_4]^{2-}(aq)$ (at high pH).
- Aluminum (Al):
  - Acidic: $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
  - Neutral: $2Al^{3+}(aq) + 3H_2O(l) \rightarrow Al_2O_3(s) + 6H^+(aq)$
  - Alkaline: $Al(s) + 4OH^-(aq) \rightarrow [Al(OH)_4]^-(aq) + 3e^-$
- Copper (Cu):
  - Acidic (aerated): $2Cu(s) + O_2(g) + 4H^+(aq) \rightarrow 2Cu^{2+}(aq) + 2H_2O(l)$
  - Neutral to alkaline: $2Cu(s) + O_2(g) + CO_2(g) + H_2O(l) \rightarrow Cu_2CO_3(OH)_2(s)$ (patina formation)

These equations show how the presence of H^+ and OH^- influence the reactions involved in the corrosion of metals.

- Material Selection and Design: The findings are critical in informing material selection for various applications, particularly in environments where the pH is likely to vary.

- Acidic Environments: In applications where acidic conditions are expected (e.g., chemical processing, industrial waste management), the study shows that iron and aluminum will have very high corrosion rates, which needs to be taken into consideration when designing equipment or structures to be used in these conditions. Copper displays lower corrosion rates, but still presents significant corrosion in acidic solutions, especially in presence of oxygen.

Therefore, when the exposure to acidic media is inevitable, these materials must be protected using a suitable corrosion protection approach, or a different material should be considered.

- Neutral Environments: The stability of aluminum at neutral pH, due to the formation of the protective Al_2O_3 layer, makes it a suitable candidate for several applications in neutral environments. However, if corrosion is a major concern, other protective means must be taken into consideration, since the results show that some corrosion is still present. Copper also shows excellent resistance at neutral pH due to the formation of a protective patina. Iron is less stable at these conditions, and is more likely to corrode.

- Alkaline Environments: In alkaline conditions, aluminum is highly susceptible to corrosion, becoming less suitable for use in these environments. Iron shows low corrosion rates under mildly alkaline solutions, but still displays corrosion at very high pH. Copper continues to be the best option at very alkaline solutions due to the higher corrosion resistance.

- Corrosion Control and Prevention: The understanding of pH's impact on corrosion is essential for designing effective corrosion prevention strategies.

- pH Adjustment: Maintaining a neutral to slightly alkaline pH environment can help mitigate corrosion in many applications by ensuring the stability of protective surface layers (as observed for aluminum and copper) and reducing the aggressiveness of hydrogen ions that can speed up electrochemical reactions that corrode metals.

- Protective Coatings: The use of protective coatings in highly acidic or alkaline environments is necessary to minimize direct contact between metals and the corrosive media. The study emphasizes that the protective layer can be inherently formed by the metal, if the correct conditions are used. However, the use of synthetic coatings, corrosion inhibitors, and alloys are also good approaches.

- Industrial and Engineering Applications: The study has implications for several key industrial and engineering sectors:

- Water Distribution: Maintaining the correct pH level in water distribution systems is important to prevent corrosion of pipelines and fittings. Iron and copper are common materials for pipelines. The results show that copper is better suited for such applications.

- Chemical Industry: The chemical industry relies on equipment that must withstand a range of pH conditions. Understanding the corrosion resistance of iron, aluminum, and copper is vital for proper selection and use of industrial equipment.

- Construction: Metals such as steel (which mainly contains iron) are commonly used as construction materials. It is necessary to evaluate the corrosion resistance in the different environmental conditions, such as rain and moisture, that the structures will be exposed to.

- Marine Applications: Marine environments can have variable pH levels. Therefore, choosing materials suitable for marine applications needs to take into consideration the effect of pH in the corrosion of metallic components of ships and other constructions.

- Environmental Considerations: Understanding how pH impacts corrosion can help prevent the release of heavy metals into the environment, reducing metal contamination of soil and water. Corrosion can be harmful if it leads to the release of heavy metals from corroded materials to the environment. This release can contaminate soil and water, impacting the wildlife, and human health.

- Advancement of Corrosion Science: This study contributes to a better understanding of the fundamental aspects of corrosion science and electrochemistry. The results obtained can be used for further studies in other metals, and under different conditions.

5.3 Concluding Remarks

This section provides concluding remarks on our study, emphasizing its importance and highlighting potential future research directions based on our findings on the corrosion behavior of iron, aluminum, and copper at different pH values.

- Summary of Findings: This study demonstrates that pH is a critical factor in determining the corrosion rates of iron, aluminum, and copper, each displaying unique behavior. The corrosion of iron is most pronounced in acidic environments, while aluminum is significantly affected by both acidic and alkaline conditions but forms a highly stable oxide layer at neutral pH, and copper exhibits good corrosion resistance across a broad pH range, especially at neutral and alkaline conditions, where a protective patina is formed. The specific corrosion rates can be observed at the tables in section 4.1.1. In general, the order of corrosion rates changes with the pH, as we have observed in our study. At pH 2 the order was Fe>Al>Cu; however, at neutral pH (7), Al is more resistant than Cu, which in turn, is more resistant than Fe.

The results corroborate the electrochemical nature of the corrosion processes, where the oxidation of the metals and the reduction of hydrogen ions (in acidic conditions) or water and oxygen (in neutral to alkaline conditions) are the main driving forces for corrosion. The chemical equations for these processes are:

- Iron (Fe):
  - Acidic: $Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$ and further oxidation in presence of oxygen: $4Fe^{2+}(aq) + O_2(g) + 4H_2O(l) \rightarrow 2Fe_2O_3.H_2O(s) + 8H^+(aq)$
  - Neutral to Alkaline: $2Fe(s) + O_2(g) + 2H_2O(l) \rightarrow 2Fe(OH)_2(s)$; $Fe(OH)_2(s) + OH^-(aq) \rightarrow [Fe(OH)_3]^-(aq)$ or $[Fe(OH)_4]^{2-}(aq)$ at higher pH.
- Aluminum (Al):

  - Acidic: $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
  - Neutral: $2Al^{3+}(aq) + 3H_2O(l) \rightarrow Al_2O_3(s) + 6H^+(aq)$
  - Alkaline: $Al(s) + 4OH^-(aq) \rightarrow [Al(OH)_4]^-(aq) + 3e^-$

- Copper (Cu):

  - Acidic (aerated): $2Cu(s) + O_2(g) + 4H^+(aq) \rightarrow 2Cu^{2+}(aq) + 2H_2O(l)$
  - Neutral to Alkaline: $2Cu(s) + O_2(g) + CO_2(g) + H_2O(l) \rightarrow Cu_2CO_3(OH)_2(s)$

These reactions highlight the crucial role of H^+ and OH^- ions in the corrosion mechanism. The metal's interaction with oxygen, water and carbon dioxide also plays an important role in the process.

- Study of the effect of temperature on the corrosion behavior at different pHs.

- Study of the effect of other anions and cations on the corrosion rate.

* Exploration of new corrosion inhibitors to be used in specific pH levels.

In conclusion, this study provides valuable insights into the pH-dependent corrosion behavior of iron, aluminum, and copper. The findings have significant practical implications for material selection, corrosion control, and the development of future corrosion mitigation strategies. This work also emphasizes the necessity of combining experimental data with a fundamental understanding of the underlying electrochemical processes to achieve a better knowledge of the corrosion mechanism and help in the design of better corrosion control systems.

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