

# CORROSION AND CORROSION PROTECTION

## Experiment 1

### Corrosion with hydrogen depolarization

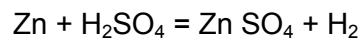
Laboratory equipment:

- analytical balance,
- beaker,
- pump rubber,
- burette,
- caliper,
- sandpaper.

Chemicals:

- zinc plate,
- 1M H<sub>2</sub>SO<sub>4</sub>,
- ethyl alcohol,

Clear the zinc sample with sandpaper, wash it with water and degrease with ethyl alcohol. After drying, weigh the sample with use of analytical balance and measure its surface using the caliper. Then place it in a measuring system – beaker filled with the 1M H<sub>2</sub>SO<sub>4</sub> sucked into burette using the rubber bulb. Since then every five minutes note the level of acid in the burette for 30 minutes. After the measurement remove the sample from the solution, wash with water and alcohol, dry and then weigh it again. The mass of dissolved zinc is proportional to the volume of evolved hydrogen. According to the reaction:



Therefore it can be determined the rate of corrosion of zinc from the cavity directly or indirectly – from the amount of hydrogen evolved.

## Experiment 2

### SEM measurement of the oxygen-depolarized cell

Laboratory equipment:

- digital Voltmeter,
- 2 beakers,
- electrical cables,
- air supply tube,
- sandpaper,

Chemicals:

- 2 steel electrodes,
- 1M NaCl,
- ethyl alcohol,

Clean steel electrodes with sandpaper paper, wash with water and then with alcohol. Build a cell according to scheme (Fig. 5.) and measure SEM of the cells. Then turn on the pump and start to supply the air to one of the electrodes. After few minutes read the SEM value of cells again.

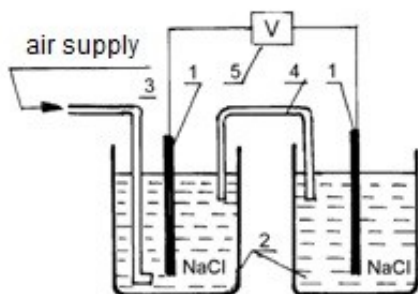


Figure 1  
Scheme of concentration corrosion cell

- 1 - steel electrodes,
- 2 - NaCl solution,
- 3 - air supply tube,
- 4 - salt bridge,
- 5 - Voltmeter.

### Experiment 3

#### Cathodic protection with sacrificial electrodes

Laboratory equipment:

- digital Voltmeter,
- 3 test tubes in a rack,
- sandpaper.

Chemicals:

- electrodes: Fe, Fe-Zn, Fe-Cu,
- 0.1M H<sub>2</sub>SO<sub>4</sub>,
- K<sub>3</sub>[Fe(CN)<sub>6</sub>] solution,
- ethyl alcohol.

The purpose of this exercise is to observe the change in the rate of corrosion of iron in contact with copper and zinc. An indicator - potassium hexacyanoferrate (III), (potassium ferricyanide) reacts with Fe<sup>+2</sup> ions origin from iron dissolving during corrosion process, forming blue compound - Fe<sub>3</sub>[Fe(CN)<sub>6</sub>]<sub>2</sub>.

Clean the electrodes with sandpaper and alcohol. Pour 2-4 cm<sup>3</sup> 0.1M H<sub>2</sub>SO<sub>4</sub> to each of three test tubes and add 2-3 drops of K<sub>3</sub>[Fe(CN)<sub>6</sub>]. Then immerse electrodes: Fe, Fe-Zn, Fe-Cu in the test tubes. Observe the intensity of solution color and rate of evolution of hydrogen.

### Experiment 4

#### Zinc electroplating

Laboratory equipment:

- DC power supply
- beaker,
- sandpaper,
- analytical balance,
- caliper,
- dryer.

Chemicals:

- electrodes: steel, Zn,
- zinc solution for galvanizing,
- ethyl alcohol.

Clean the surface of the steel electrode with sandpaper, rinse with water, degrease with alcohol, dry and weigh with use of analytical balance. Pour the solution into the beaker, immerse steel electrode, take it out and measure immersed surface. Connect the electrical circuit according to scheme on Fig.2. Turn on the DC power supply by setting the value of the current so that the current density would be approximately 2A/ dm<sup>2</sup>. The electrolysis should be carried out for 10 minutes.

After time elapsed turn off the power supply, rinse the electrode with distilled water, dry and weigh it one more time. Repeat the process of galvanizing with the current density about 10 times greater. Compare the quality of the layers produced at different current density. Pour the solution into a container for the used bath.

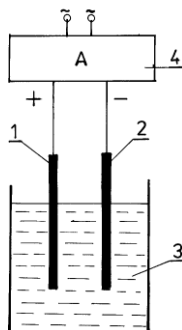


Figure 2  
Scheme of a zinc plating device  
1 - zinc plate,  
2 - steel plate,  
3 - galvanizing solution,  
4 - DC power supply.

## Experiment 5 Measurement of the corrosion potential of various metals

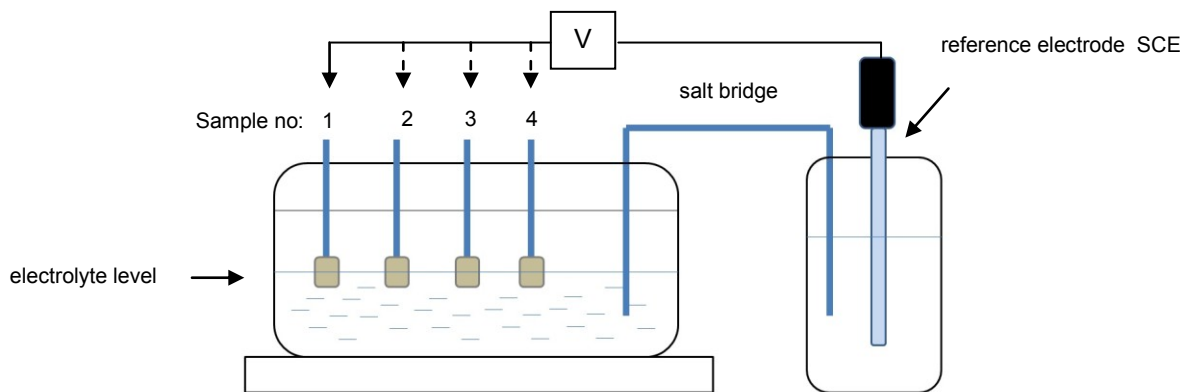
### Laboratory Equipment:

- 4 samples of metals or alloys
- measuring cell
- multimeter
- salt bridge
- reference electrode (saturated calomel electrode - SCE)
- electric cables
- sandpaper

### Chemicals:

- 1M NaCl
- ethyl alcohol

Clean the cylindrical metal samples with sandpaper, rinse with distilled water and degrease with alcohol. Pour a 1 M NaCl solution to the measurement vessel so that each sample was immersed in the solution to half of volume!!!. Connect the vessel with reference electrode (SCE) using a salt bridge. Reference electrode wire connect to COM socket of multimeter. Connect the free cable to the V/mV socket of multimeter and measure sequentially the corrosion potential for each of the four samples at 5 minute intervals for one hour. Note the results in the table in the report in  $E_{vs SCE}$  column.



## Experiment 6 Cathodic protection

### Laboratory Equipment:

- DC power supply
- beaker
- sandpaper

### Chemicals:

- two steel electrodes
- KCl solution
- $K_3[Fe(CN)_6]$  solution
- phenolphthalein
- ethyl alcohol

Clean the surface of the electrodes with sandpaper, rinse with water, degrease with alcohol and then dry them. Pour the solution into the beaker and add 5 drop of phenolphthalein and 5 drops of  $K_3[Fe(CN)_6]$  solution. Immerse the electrodes and connect them to the power supply according to the scheme on Fig.3. Turn on the DC power supply and set the value of the potential so that the it would be approximately 1V

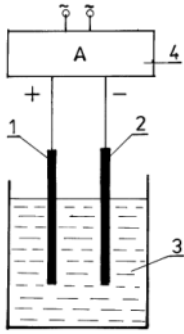


Figure 3

1. Steel electrode
2. Steel electrode
3. KCl solution + indicators
4. DC power supply

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**Experiment 1**  
**Corrosion with hydrogen depolarization**

Calculate the corrosion rate of the zinc from the measurement of the weight loss of the sample. Give the corrosion rate ratios  $V_c$  and  $V_p$  assuming density of zinc  $d = 7.14 \text{ g/cm}^3$  and using the following formulas:

$$V_c = \frac{\Delta}{s \cdot t} \text{ [g/m}^2 \text{ day]} \qquad V_p = \frac{V_c \cdot 365}{1000 \cdot d} \text{ [mm/year]}$$

	time [min]	time [day]	$m_1$ [g]	$m_2$ [g]	$\Delta m$ [g]	surface [mm <sup>2</sup> ]	surface [m <sup>2</sup> ]	$V_c$ [g/m <sup>2</sup> ·day]	$V_p$ [mm/year]
<b>Zn</b>									

time [min]	Level of the acid [cm <sup>3</sup> ]	Hydrogen volume [cm <sup>3</sup> ]	Mole number H <sub>2</sub> (Zn)	weigh Zn [g]
0				
5				
10				
15				
20				
25				
30				

Compare the corrosion rate determined gravimetrically and calculated based on the volume of hydrogen gas evolved during the reaction.

**Experiment 2**  
**SEM measurement of the oxygen-depolarized cell**

Cell	SEM [V]
Fe   NaCl   Fe	
Fe   NaCl   Fe(O <sub>2</sub> )	

- Determine the impact of the incoming air (oxygen) in the SEM cells .....
- Determine which electrode is the cathode and the anode cell that corrosion .....

- Describe the differences in the appearance of the cathode and anode corrosion cells
- .....

**Experiment 3**  
**Cathodic protection with sacrificial electrodes**

- On the basis of the color of the solution to determine where the corrosion rate of iron is the greatest?
- In the course of the study, you can also see the release of the gas (hydrogen) - on which metal hydrogen evolution reaction occurs, and how fast is it?
- Which metal is the protector in case b) and c)?

metals	Color intensity	Corrosion	Hydrogen evolution	Protector
Fe				
Fe – Zn				
Fe – Cu				

**Experiment 4**  
**Zinc electroplating**

- Calculate the theoretical weight gain of zinc for the zinc plating steel rod from the formula:

$$m_{Zn} = k \cdot I \cdot t = \dots\dots\dots$$

where: k = 1.22 [g / Ah]

I - current [A]

T - Time zinc [h]

- Calculate the current efficiency of the process of zinc in% as the ratio of the sample weight gain zinc dedicated to the theoretical amount of zinc and calculated from Faraday's law.

$$W = (\Delta m / m_{Zn}) 100\% = \dots\dots\dots$$

	Electrode surface s [dm <sup>2</sup> ]	Time t [h]	Weigh before electroplating m <sub>1</sub> [g]	Weigh after electroplating m <sub>1</sub> [g]	Δ m [g]	m <sub>Zn</sub> calc. from Faraday's law [g]	Output of the process [%]
Fe							

Determine the differences in the appearance of layers produced with different current density:

.....  
 .....

**Experiment 5**  
**Measurement of the corrosion potential of various metals**

time t [min]	CORROSION POTENTIAL E [V]							
	Sample nr 1 (.....) type of metal		Sample nr 2 (.....) type of metal		Sample nr 3 (.....) type of metal		Sample nr 4 (.....) type of metal	
	E <sub>SCE</sub>	E <sub>SHE</sub>	E <sub>SCE</sub>	E <sub>SHE</sub>	E <sub>SCE</sub>	E <sub>SHE</sub>	E <sub>SCE</sub>	E <sub>SHE</sub>
0								
5								
10								
15								
20								
25								
30								
35								
40								
45								
50								
55								
60								

Calculate the potential of metal in relation to the SHE(standard hydrogen electrode) basing on the formula:

$$E_{SHE} = E_{SCE} + E_{vs.SCE}, \quad \text{where: } E_{SCE} = +0.244 \text{ V}$$

Draw curves potential vs. time  $E = f(t)$ . Explain why there are differences in corrosive potential and explain what is the cause of increase or decrease in potential over time.

**Experiment 6**  
**Cathodic protection**

Note the observations:

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