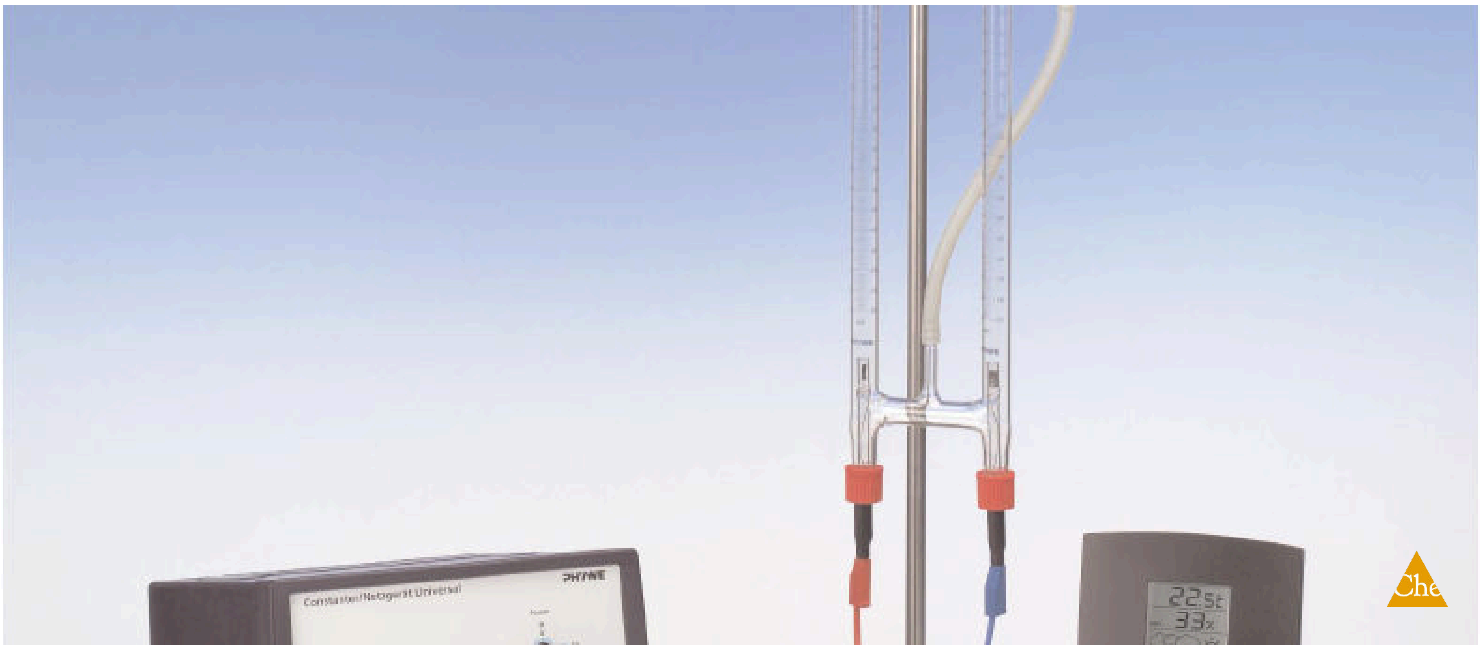
Determination of Faraday's constant



Chemistry > Physical chemistry > Electrochemistry > Electrochemical measurement set

Difficulty level	Group size	Preparation time	Execution time
hard	-	-	-

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General information

Application

Setup

Faraday's constant is a constant of nature. It is used when electrical currents lead to material flow such as in batteries. As such it sees wide use in the energy storage industry.

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Other information (1/2)

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Prior knowledge

The prior knowledge for this experiment is found in the Theory section.

Main principle

Faraday’s laws of electrolysis describe the correlation between the amounts of substances transformed in the reactions at the electrodes and the charge applied (amount of electricity). Faraday’s constant, which appears as a proportionality factor, can be determined experimentally from this dependence.

Other information (2/2)

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Learning objective

The goal of this experiment is to investigate Faraday's law.

Tasks

1. Determine Faraday’s constant from the dependence of the volumes of hydrogen and oxygen evolved on the charge applied in the hydrolysis of dilute sulphuric acid.

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Safety Instructions

Hydrogen
H220: Extremely flammable gas.
P210: Keep away from heat, hot surfaces, sparks, open flames and other ignition sources. No smoking.

Oxygen
H270: May cause or intensify fire; oxidizer
P220: Keep/Store away from clothing/…/combustible materials.

Sulphuric acid
H290: May be corrosive to metals.
H314: Causes severe skin burns and eye damage.
P280: Wear protective gloves/protective clothing/eye protection/face protection.

Theory (1/3)

If an electric current I is forced by a direct source over a time period t, changes occur in the given electrolysis system due to anodic oxidation and cathodic reduction. Oxygen is formed at the anode and hydrogen at the cathode.

Oxidation (at the anode):

$$6\,H_2O \;\rightarrow\; O_2 + 4\,H_3O^+ + 4\,e^-$$

Reduction (at the cathode):

$$4\,H_3O^+ + 4\,e^- \;\rightarrow\; 2\,H_2 + 4\,H_2O$$

The sum of these is:

$$2\,H_2O \;\rightarrow\; 2\,H_2 + O_2$$

This experiment shows that the gas volumes evolved are proportional to the amperage and to the time. Faraday's first law can be derived from this:

$$n = \frac{I \cdot t}{z_R F} = \frac{|q|}{z_R F} \quad (1)$$

where n Quantity of substance, q charge, z_R Number of transferred electrons per formula conversion and F Faraday's constant; the product of Avogadro's number N_A and the unit charge e.

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Theory (2/3)

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The Faraday constant represents the quantity of electricitity that is required to liberate or deposit a gram-equivalent (= N_A monovalent ions). The same quantity of electricity always liberates or deposits equivalent quantities of substances from different electrolytes (Faraday's second law).

Under the proper conditions (low pressure, sufficiently high temperature), the quantities of gas n evolved in the two electrode reactions can be expressed as the corresponding gas volumes V using the general equation of state for ideal gases:

$$n = \frac{pV}{RT} \quad (2)$$

where R Universal gas constant; 8.31441 J/(K mol), p pressure and T Absolute temperature.

Theory (3/3)

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With equation (1) we obtain

$$V = \frac{|q|RT}{z_R pF} \quad (3)$$

According to this relationship, linear correlations (Fig. 1) result for the function V = f(q) which has the following slope:

$$\frac{\Delta V}{\Delta |q|} = \frac{RT}{z_R pF} \quad (4)$$

From the slopes of the graphs obtained for hydrogen (z_R = 2) and oxygen (z_R = 4) Faraday's constant can be calculated, if p and T have been determined.

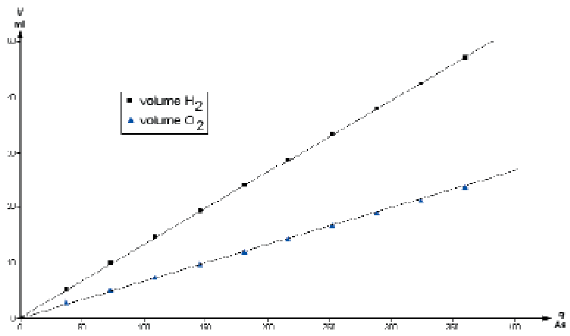

Fig. 1: Correlation between the transferred charge and the evolved volumes of hydrogen and oxygen in the electrolysis of diluted sulphuric acid (T = 296.05 K and p =100.4 kPa).

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Equipment

Position	Material	Item No.	Quantity
1	PHYWE Power supply, universal DC: 0...18 V, 0...5 A / AC: 2/4/6/8/10/12/15 V, 5 A	13504-93	1
2	Water decomposition apparatus according to Hoffmann, incl. tripod	MAU-25014002	1
3	Digital multimeter, 600V AC/DC, 10A AC/DC, 20 MΩ, 200 µF, 20 kHz, −20°C… 760°C	07122-00	1
4	On/off switch	06004-00	1
5	Connecting cord, 32 A, 750 mm, blue	07362-04	1
6	Connecting cord, 32 A, 500 mm, red	07361-01	1
7	Connecting cord, 32 A, 250 mm, red	07360-01	2
8	Right angle boss-head clamp	37697-00	2
9	Universal clamp	37715-01	2
10	Digital stopwatch, 24 h, 1/100 s and 1 s	24025-00	1
11	Weather monitor, 6 lines LCD	87997-10	1
12	Beaker, Borosilicate, low form, 600 ml	46056-00	1
13	Pasteur pipettes, 250 pcs	36590-00	1
14	Rubber caps, 10 pcs	39275-03	1
15	Funnel, glass, top dia. 80 mm	34459-00	1
16	Wash bottle, plastic, 500 ml	33931-00	1
17	Sulphuric acid, 95-97%, 500 ml	30219-50	1
18	Water, distilled 5 l	31246-81	1

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Setup and Procedure

Setup

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Set up the experiment as shown in Fig. 2.

Prepare the 20 % sulphuric acid solution required for the experiment by carefully adding 62 g of concentrated sulphuric acid to 238 g of distilled water in a 500 ml beaker.

Open the cocks of the electrolysis apparatus and fill in 200 ml of the sulphuric acid solution through the levelling vessel. Eliminate air bubbles in the apparatus by gently knocking on the tubes. Fill the measuring tubes exactly by raising the levelling vessel, then close the stopcocks.

Fig. 2

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Procedure

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Apply about 400 mA for just a few minutes to carry out electrolysis and so saturate the solution in the apparatus with the evolved gases. Turn off the power supply and open the stopcocks to again completely fill the tubes with solution, then close the stopcocks.

In the first part of the experiment, carry out electrolysis at a constant amperage of between 200 and 300 mA for 10 minutes, starting the stopwatch at the same time as current is applied. Interrupt the supply of power each minute and read off the volumes of gases that have been evolved. To do this, adjust the height of the levelling vessel so that the meniscus in the vessel is level with the meniscus in the tube being measured. Record the time and the gas volumes in a Table. When the tenth readings have been taken, select a suitable scale and plot a graph of the gas volumes measured against time.

In the second part of the experiment, subject the solution to electrolysis at three different amperages (70, 140, 210 mA), each for the same length of time (between 5 and 10 minutes). Read off the gas volumes when the chosen time has elapsed and plot them in a graph. Re-fill the tubes with acid prior to each measurement. Determine the atmospheric pressure and room temperature.

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Evaluation

Data and Results

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An electrolysis performed with I = 300 mA in steps of t = 120 s at T = 296.06 K and p = 100.4 kPa provided the linear correlations between the applied charge q = I · t and the thus-evolved volumes of hydrogen and oxygen presented in Fig. 1.

Using the slopes obtained from compensation calculation

$\Delta V(H_2)/\Delta|q| = 1.298 \cdot 10^{-7}\, m^3 As^{-1}$ and $\Delta V(O_2)/\Delta|q| = 0.640 \cdot 10^{-7}\, m^3 As^{-1}$

the Faraday constant is determined to be $F = 94477\ \frac{As}{mol}$ resp. $F = 95739\ \frac{As}{mol}$.

Literature value: $F = 96485\ \frac{As}{mol}$

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