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ARNIKO AWASIYA H.S SCHOOL

A project work for the partial fulfillment of chemistry department

SUBMITTED BY

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CERTIFICATE OF APPROVAL

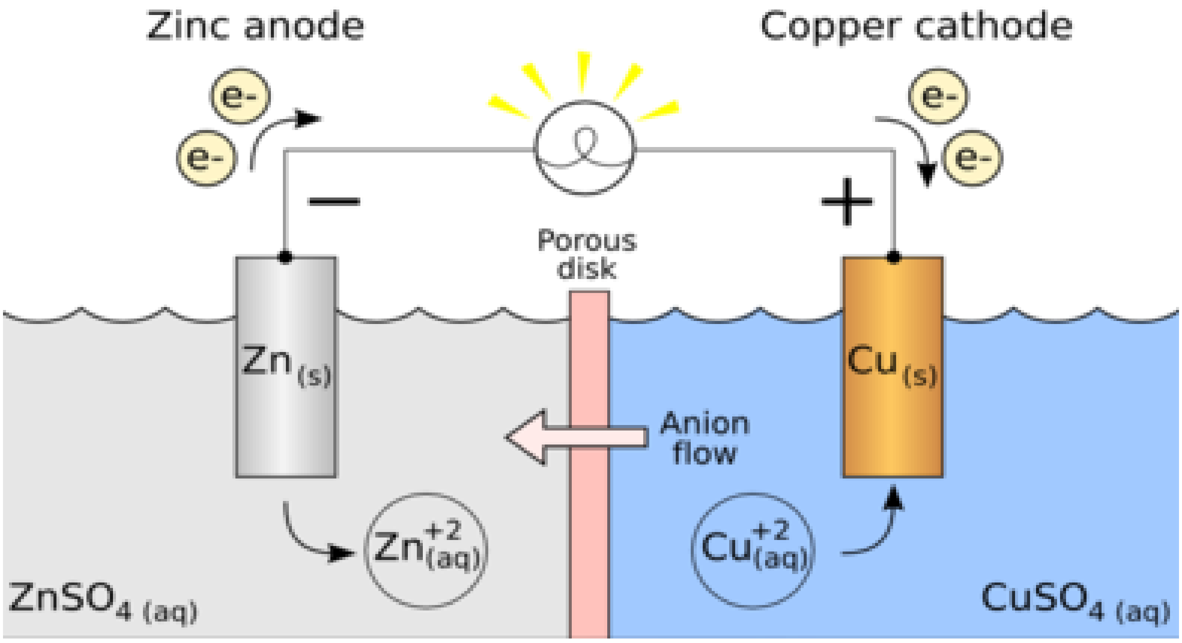
DEPARTMENT OF CHEMISTRY:-

This is to certify that we have examined the chemistry project work titled "To study the working principle of galvanic cell and calculate standard emf with the help of reference electrode" submitted by Barsha Pariyar is a bonafide student of class XII bearing Roll no. 801015. We hereby accord our approval of the project work carried out and presented in a manner required for its acceptance for the partial fulfillment of the Arniko Awasiya college. It is understood that this approval does not necessarily accept the every statement made, opinion expressed or conclusion drawn as recorded in this project. It only signifies the acceptance of the project work for the purpose for which it has been submitted.

Signature of external examiner ***Signature of Chemistry teacher***

INTRODUCTION

It involves a chemical reaction that makes the electrical energy. During a redox reaction, a galvanic cell utilizes the energy transfer between the electrons to convert chemical energy into electerical energy.



Galvanic cell utilizes the ability to separate the flow of electrons in the process of oxidization and reduction, causing a half reaction and connecting each with a wire so that a path can be formed for the flow of electrons through such wire. This flow of electrons is essentially called a current. Such current can be made to flow through a wire to complete a circuit and obtain its output in any device such as a television or a watch.

A galvanic cell can be made out of any two metals. These two metals can form the anode and the cathode if left in contact with each other. This combination allows the galvanic corrosion of that metal which is more anodic. A connecting circuit shall be required to allow this corrosion to take place.

Setup of a Galvanic Cell

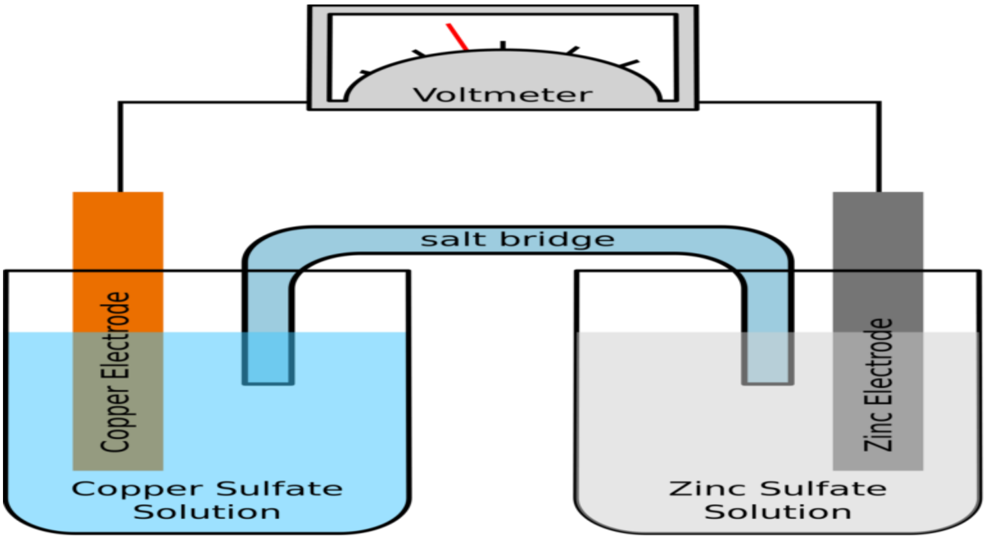
In order to create a galvanic cell, one would have to go through the following setup. The cell would ideally include two electrodes. One of these electrodes, the cathode, shall be a positively charged electrode while the other, shall be the anode, the negatively charged electrode.

These two electrodes shall form the two essential components of the galvanic cell. The chemical reaction related to reduction shall take place at the cathode while the oxidation half-reaction shall take place at the anode. As has already been said, any two metals can be used to create the chemical reaction.

Understanding the Galvanic Cell with an Example

Let us take an example where the two metals involved in the chemical reaction are zinc and copper. As the chemical reaction takes place, Zinc would end up losing two electrons. This will be taken up by copper to become elemental copper. Since these two metals will be placed in two separate containers and would be connected by a conducting wire, an electric current would be formed, which would transfer all electrons from one metal to another.

At the same time, the two metals shall be immersed in a salt solution, say, Zinc sulphate and Copper sulphate in this case. In this case, the two solutions are not mixed together directly but can be joined using a bridge or a medium. This medium shall be responsible for the transfer of ions but also make sure that the two solutions do not come to mix with each other



Such bridge helps in completing the circuit for carrying the electric charge and also makes sure that the solutions in the containers with the metals remain neutral and do not mix with each other. As long as the salt bridge does not interfere with the redox reaction, under which oxidization and reduction are taking place, it does not matter which salt bridge is being used in the chemical reaction.

Some Important Terms

Some of the important terms brought into use in a galvanic cells are listed below:

- *Phase boundaries: It refers to the two metals which act as cathode and anode.*
- *Salt bridge: The connecting bridge or medium that allows a redox reaction to take place.*
- *Oxidation and reduction: The chemical processes that allow the electric current to form and flow through a galvanic cell.*

WORKING OF THE GALVANIC CELL

A piece of zinc going into a solution as zinc ions, with each Zn atom giving up 2 electrons, is an example of an oxidation half-reaction.

$$Zn(s) \rightarrow Zn2+(aq) + 2e-$$

The oxidation number of $Zn(s)$ is 0 and the oxidation number of the Zn^{2+} is +2. Therefore, in this half-reaction, the oxidation number increases, which is another way of defining an oxidation. In contrast, the reverse reaction, in which Zn^{2+} ions gain 2 electrons to become Zn atoms, is an example of reduction.

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

In a reduction there is a decrease (or reduction) in oxidation number. Chemical equation representing half-reactions must be both mass and charge balanced. In the half-reactions above, there is one zinc on both sides of the equation. The charge is balanced because the 2+ charge on the zinc ion is balanced by two electrons, $2e^-$, giving zero net charge on both sides.

Another example of reduction is the formation of solid copper from copper ions in solution.

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

In this half-reaction the oxidation number of the aqueous copper is +2, which decreases to 0 for the solid copper, and again charge and mass are balanced. However, no half-reaction can occur by itself. A redox reaction results when an oxidation and a reduction half-reaction are combined to complete a transfer of electrons as in the following example:

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

The electrons are not shown because they are neither reactants nor products but have simply been transferred from one species to another (from Zn to Cu^{2+} in this case). In this redox reaction, the $Zn(s)$ is referred to as the reducing agent because it causes the Cu^{2+} to be reduced to Cu. The Cu^{2+} is called the oxidizing agent because it causes the $Zn(s)$ to be oxidized to Zn^{2+}.

EXPERIMENTAL DATA

Any half-reaction can be expressed as a reduction as illustrated in the case where equation (1) can be reversed to equation (2). A measure of the tendency for a reduction to occur is its reduction potential, E, measured in units of volts. At standard conditions, 25 °C and concentrations of 1.0 M for the aqueous ions, the measured voltage of the reduction halfreaction is defined as the standard reduction potential, E°.

Standard reduction potentials have been measured for many half-reactions and they are listed in tables. A short list is also provided at the end of the In-Lab section. For the reduction half-reactions in equations (2) and (3), the standard reduction potentials are –0.76 V for zinc and +0.34 V for copper. The more positive (or less negative) the reduction potential, the greater is the tendency for the reduction to occur. So Cu2+ has a greater tendency to be reduced than Zn2+. Furthermore, Zn has a greater tendency to be oxidized than Cu. The values of E° for the oxidation halfreactions are opposite in sign to the reduction potentials: +0.76 V for Zn and –0.34 V for Cu.

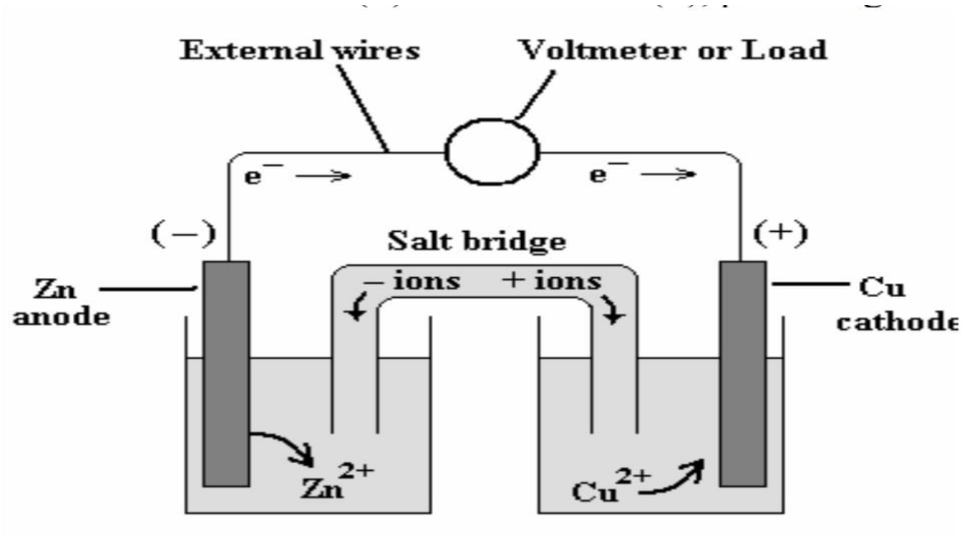


Figure 1. Galvanic cell (or battery) based on the redox reaction in equation (4).

The cell potential, Ecell, which is a measure of the voltage that the battery can provide, is calculated from the half-cell reduction potentials:

Ecell = Ecathode – Eanode

At standard conditions, indicated by the superscript o , the standard cell potential, E°cell, is based upon the standard reduction potentials, as shown in equation (5).

$E°cell = E°cathode – E°anode$

(5) Based on the values for the standard reduction potentials for the two half-cells in equation (4) [–0.76 V for zinc anode and +0.34 V for copper cathode], the standard cell potential, E°cell, for the galvanic cell in Figure 1 would be:

$E°cell = +0.34\ V – (–0.76\ V) = +1.10\ V$

The positive voltage for Eo cell indicates that at standard conditions the reaction is spontaneous. Recall that ΔGo = − nFEo cell, so that a positive Eo cell results in a negative ΔGo . Thus the redox reaction in equation (4) would produce an electric current when set up as a galvanic cell.

When conditions are not standard, the Nernst equation, equation (6), is used to calculate the potential of a cell. In the Nernst equation, R is the universal gas constant with a value of 8.314 J/(K⋅mol), T is the temperature in K, and n is the number of electrons transferred in the redox reaction, for example, 2 electrons in equation (4). Q is the reaction quotient for the ion products/ion reactants of the cell. The solid electrodes have constant "concentrations" and so do not appear in Q. F is the Faraday constant with a known value of 96,500 J/(V⋅mol).

$$E_{cell} = E^0_{cell} - \left(\frac{RT}{nF}\right)(\ln Q)$$

Standard Reduction Potentials:

Electrode	E^0
$Ag^+ + e^- \rightarrow Ag$	+0.80 V
$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34 V
$Pb^{2+} + 2e^- \rightarrow Pb$	-0.13 V

$Fe^{2+} + 2e^- \rightarrow Fe$	-0.44 V
$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76 V
$Al^{3+} + 3e^- \rightarrow Al$	-1.66 V
$Mg^{2+} + 2e^- \rightarrow Mg$	-2.34 V

Galvanic cell	E^0cell measured	Anode	Equation of anode half reaction	Cathode	Equation for cathode half reaction
Cu-Zn	0.783 V	Cu	$Cu \rightarrow Cu^{2+} + 2e$	Zn	$Zn^{2+} + 2e^- \rightarrow Zn$
Cu-Fe	0.986 V	Cu	$Cu \rightarrow Cu^{2+} + 2e$	Fe	
Zn-Fe	-0.109 V	Zn	$Zn \rightarrow Zn^{2+} + 2e$	Fe	

Reference

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