Chapter - 5
Electrochemistry

- ✓ Electrochemistry is the branch of chemistry which deals with the studies on the processes that produce electricity through chemical reactions and use electricity to bring about chemical reactions.
- ✓ Electrochemical cells are the devices that converts chemical energy into electrical energy or electrical energy into chemical energy through redox reactions.
- ✓ Electrochemical cells can be divided into two types.
  i. Galvanic cells ii. Electrolytic cells

Galvanic cells

- ✓ Galvanic cells are devices that convert chemical energy into electrical energy.
- ✓ In a galvanic cell, the electrode at which oxidation occurs is called anode and the electrode at which reduction occurs is the cathode.
- ✓ In a galvanic cell, anode attains negative charge and cathode gets positive charge.
- ✓ In a galvanic cell, electron flow from the negative electrode (anode) to the positive electrode (cathode)
- ✓ In a galvanic cell, the direction of flow of electric current is from the positive electrode (cathode) to the negative electrode (anode)
- ✓ Daniel cell ($Zn - Cu\ cell$) is an example of a galvanic cell.

Daniel Cell

- Daniel cell is a galvani cell that converts chemical energy into electrical energy using zinc and copper electrodes.

Anode ($-ve$ electrode) : Zn rod

Cathode ($+ve$ electrode) : Cu rod

Equation of chemical reaction taking place at

Anode (Oxidation) : $Zn \rightarrow Zn^{2+}(aq) + 2e^{-}$

Cathode (Reduction) : $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu$

Cell reaction (Redox reaction) :

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

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- Electron flow from the zinc rod (anode) to the copper rod (cathode)
- The direction of flow of electric current is from the copper rod (cathode) to the zinc rod (anode).

Mg – Zn rod

Anode : Mg rod

Cathode : Zn rod

Equation of chemical reaction taking place at

Anode (Oxidation) : $Mg \rightarrow Mg^{2+}(aq) + 2e^-$

Cathode (Reduction) : $Zn^{2+}(aq) + 2e^- \rightarrow Zn$

Cell reaction (Redox reaction) :

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

- Electron flow from the magnesium rod (anode) to the zinc rod (cathode)
- The direction of flow of electric current is from the zinc rod (cathode) to the magnesium rod (anode).

Cu – Ag rod

Anode : Cu rod

Cathode : Ag rod

Equation of chemical reaction taking place at

Anode (Oxidation) : $Cu \rightarrow Cu^{2+}(aq) + 2e^-$

Cathode (Reduction) : $Ag^+(aq) + e^- \rightarrow Ag$

$$\Longrightarrow 2\,Ag^+(aq) + 2e^- \rightarrow 2Ag$$

Cell reaction (Redox reaction) :

$$Cu + 2\,Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2\,Ag$$

- Electron flow from the copper rod (anode) to the silver rod (cathode)

The direction of flow of electric current is from the silver rod (cathode) to the copper rod (anode).

Reactivity Series

✓ The series in which elements are arranged in the decreasing order of their reactivity is known as reactivity series.

Potassium	K
Sodium	Na
Calcium	Ca
Magnesium	Mg
Aluminium	Al
Zinc	Zn
Iron	Fe
Nickel	Ni
Tin	Sn
Lead	Pb
Hydrogen	H
Copper	Cu
Mercury	Hg
Silver	Ag
Gold	Au

Reactivity decreases ⟶

Applications of the reactivity series

1. To identify which metal is being displaced

✓ The more reactive metal displace the less reactive metal from its salt solution. Such chemical reactions are known as displacement reactions

Experiment 1 : Place a zinc (Zn) rod in a beaker of copper sulphate $(CuSO_4)$ solution.

Observations

i. The blue colour of the solution gradually turns colourless (Reason – **The decrease in the concentration of Cu^{2+} ions**)

ii. The size or weight of the zinc rod start increasing. (Reason : Copper atoms get deposited on Zn rod)

iii. Zinc displace copper from $CuSO_4$ solution (Reason – Zinc is more reactive than copper)

- Chemical equation of this reaction is

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

The atom underwent oxidation $= Zn$

The atom underwent reduction $= Cu$

Equation for oxidation is

$$Zn \rightarrow Zn^{2+} + 2e^-$$

Equation for reduction is

$$Cu^{2+} + 2\,e^- \rightarrow Cu$$

Experiment 2 : Take a little silver nitrate $(AgNO_3)$ solution in a beaker and dip a copper rod in it.

Observations

i. The initially colourless solution gradually turns blue (Reason – **Due to the formation of Cu^{2+}** ions)

ii. The size or weight of the copper rod start increasing. (Reason : Silver atoms get deposited on Cu rod)

iii. Copper (Cu) displace silver (Ag) from $AgNO_3$ solution (Reason – Copper is more reactive than silver)

- Chemical equation of this reaction is

$$Cu + 2\,AgNO_3 \rightarrow Cu(NO_3)_2 + 2\,Ag$$

The atom underwent oxidation $= Cu$

The atom underwent reduction $= Ag$

Equation for oxidation is

$$Cu \rightarrow Cu^{2+} + 2e^-$$

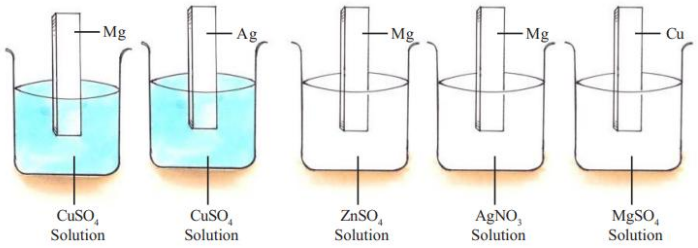
Equation for reduction is

$$Ag^+ + 1\,e^- \rightarrow Ag$$

$$OR$$

$$2\,Ag^+ + 2\,e^- \rightarrow 2\,Ag$$

Question : Observe the pictures given below. Based on the reactivity series, predict which among these undergo a displacement reaction.



Answer :

Metal	Solution	Displacement reaction	reason
Mg	$CuSO_4$	Takes place	Mg is more reactive than Cu
Ag	$CuSO_4$	Does not take place	Ag is less reactive than Cu
Mg	$ZnSO_4$	Takes place	Mg is more reactive than Zn
Mg	$AgNO_3$	Takes place	Mg is more reactive than Ag
Cu	$MgSO_4$	Does not take place	Cu is less reactive than Mg

2. To identify oxidising agent and reducing agent

- ✓ In a redox reaction, the more reactive metal gets oxidised while the ion of the less reactive metal gets reduced.
- ✓ The substance that causes oxidation is called an oxidising agent and the substance that causes reduction is called a reducing agent.
- ✓ In a redox reaction, the oxidising agent gets reduced and the reducing agent gets oxidised.
- ✓ In redox reactions, the **more reactive metal** acts as a **reducing agent** and the **less reactive** metal acts as an **oxidising agent**.

Example : 1

$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

Here Zn is more reactive than Cu, hence

The metal get oxidised $= Zn$

The ion get reduced $= Cu^{2+}$

Oxidising agent $= Cu^{2+}$

Reducing agent $= Zn$

Example : 2

$Cu + 2\ AgNO_3 \rightarrow Cu(NO_3)_2 + 2\ Ag$

Here Cu is more reactive than Ag, hence

The metal get oxidised $= Cu$

The ion get reduced $= Ag^{+}$

Oxidising agent $= Ag^{+}$

Reducing agent $= Cu$

Example : 3

$Mg + CuSO_4 \rightarrow MgSO_4 + Cu$

Here Mg is more reactive than Cu, hence

The metal get oxidised $= Mg$

The ion get reduced $= Cu^{2+}$

Oxidising agent $= Cu^{2+}$

Reducing agent $= Mg$

3. To identify the displacement of hydrogen from acid

- ✓ The metals placed above hydrogen in the reactivity series can displace hydrogen from dilute acids.

Question : Which of the following metals can displace hydrogen from hydrochloric acid?

$$Sodium, gold, silver, aluminium$$

Answer : Both the metals sodium (Na) and aluminum (Al) are places above hydrogen in reactivity series. Hence they can displace hydrogen from hydrochloric acid.

Electrolytic cells

- ✓ The cells which utilise electrical energy to bring about chemical changes are called electrolytic cells.
- ✓ An electrolytic cell convert electrical energy into chemical energy

Electrolyte

- ✓ Electrolytes are substances that conduct electricity in molten state or in an aqueous solution and undergo chemical changes.
- ✓ Acids, alkalis and salts are electrolytes in their molten state or in aqueous solution.
- ✓ Ions are free to move in the liquid state or in an aqueous solution of an electrolyte. These ions are responsible for the electrical conductivity in the electrolyte.

Electrolysis

- ✓ Electrolysis is the process by which an electrolyte undergoes chemical changes when electricity is passed through it.
- ✓ Electrodes are substances which pass electricity to the electrolytes.
- ✓ During electrolysis, the electrode connected to the positive terminal of the battery is called the anode and the electrode connected to the negative terminal is called the cathode.
- ✓ Reduction takes place at the cathode and oxidation at the anode.
- ✓ The positive ions which are attracted towards the negative electrode (cathode) are called cations.
- ✓ The negative ions which are attracted towards the positive electrode (anode) are called cations.

	Galvanic cell	Electrolytic cell
Oxidation taking place at	Anode	Anode
Reduction taking place at	Cathode	Cathode
Positive electrode	Cathode	Anode
Negative electrode	Anode	Cathode

Electrolysis of molten sodium chloride (NaCl)

- ✓ In the solid state, sodium chloride does not conduct electricity because the ions are not free to move.
- ✓ Molten NaCl is a good conductor of electricity because its ions move freely.
- ✓ $Na^+$and Cl^- are the ions in molten sodium chloride.
- ✓ The negative ion Cl^- attracted towards the anode (positive electrode).
- ✓ Chlorine gas (Cl_2) having a pale green colour, is liberated at the anode (negative electrode).
- ✓ The positive ion Na^+ attracted towards the cathode.
- ✓ Silvery molten sodium metal (Na) is formed at the cathode.

Ions in molten NaCl	$Na^+$and Cl^-
Product obtained at anode	Chlorine (Cl_2)
Product obtained at cathode	Sodium (Na)
Equation for the reaction taking place at anode	$2Cl^- \rightarrow Cl_2 + 2e^-$
Equation for the reaction taking place at cathode	$Na^+ + e^- \rightarrow Na$

Electrolysis of aqueous sodium chloride solution

- ✓ Na^+, Cl^-, H_3O^+, OH^- are the ions in sodium chloride solution.
- ✓ The negative ions Cl^- and OH^- are attracted towards the anode (positive electrode).
- ✓ When compared to Cl^-, OH^- and H_2O, the tendency to get oxidised is greater for chlorine ions (Cl^-). Hence chlorine gas is liberated at anode.
- ✓ The positive ions Na^+ and H_3O^+ are attracted towards the cathode (negative electrode).
- ✓ As H_2O has a greater tendency to get reduced, compared to Na^+ and H_3O^+ ions, hydrogen gas is liberated at the cathode.

Ions in NaCl solution	Na^+, Cl^-, H_3O^+, OH^-
Product obtained at anode	Chlorine (Cl_2)
Product obtained at anode	Hydrogen (H_2)
Equation for the reaction taking place at anode	$2Cl^- \rightarrow Cl_2 + 2e^-$
Equation for the reaction taking place at cathode	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

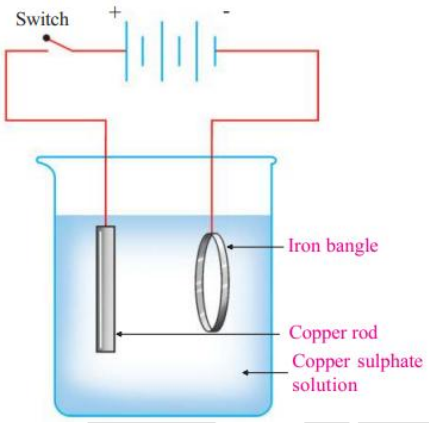
- ✓ When phenolphthalein is added to a sodium chloride solution during electrolysis, the solution around the cathode turns pink.
  (Reason – The basic compound sodium hydroxide is formed around the cathode by the combination of Na^+ ions and OH^- ions.)

Electroplating

- ✓ Electroplating is the process of coating a layer of one metal onto another metal through electrolysis.
- ✓ Electroplating is helpful for improving the appearance of the metal and also in preventing metallic corrosion.
- ✓ The object to be coated is connected to the negative terminal of the battery
- ✓ The metal which is used for plating is connected to the positive terminal of the battery.
- ✓ A salt solution of the metal to be plated is used as the electrolyte.

Metals to be covered	Electrolyte
Silver	Silver nitrate solution / sodium cyanide + silver cyanide solution
Gold	Gold Sodium cyanide + gold cyanide solution

Copper plating on iron bangle



Anode ($+ve$ electrode)	Copper rod
Cathode ($-ve$ electrode)	Iron bangle
Electrolyte	$CuSO_4$ solution
Equation of reaction taking place at anode	$Cu \rightarrow Cu^{2+} + 2e^-$
Equation of reaction taking place at cathode	$Cu^{2+} + 2e^- \rightarrow Cu$

Applications of electrolysis

- ✓ Production of metals like sodium, potassium, calcium and aluminium.
- ✓ Production of nonmetals like hydrogen, oxygen and chlorine.
- ✓ Manufacture of compounds like sodium hydroxide and potassium hydroxide.
- ✓ Gold plated jewellery, silver plated utensils, chromium plated iron objects etc are produced through electroplating.
- ✓ Electrolysis is useful in the purification of metals such as copper, gold etc.

Electrochemical cells as energy sources

- ✓ A cell is a device that helps to convert the energy released in a chemical reaction directly into electricity.
- ✓ Cells perform two main functions.

1. As portable sources of electrical energy:

   Examples range from the button cells, used in electronic watches, to the lead acid cells used in vehicles.

2. As storage devices of electrical energy provided by an external source:

   Such cells can be used for powering electric vehicles, emergency power distribution and storing solar energy.

Different types of cells

- ✓ There are two main types of cells.
  i. Primary cells
  ii. Secondary cells

Primary cells

- ✓ In primary cells, the cell become dead, when the electrical energy produced by chemical reaction is used up.
- ✓ Primary cells cannot be recharged and reused.
- ✓ Examples of primary cells are
  i. Dry cells (commonly used in clocks)
  ii. Button cells (used in watches)

Secondary cells

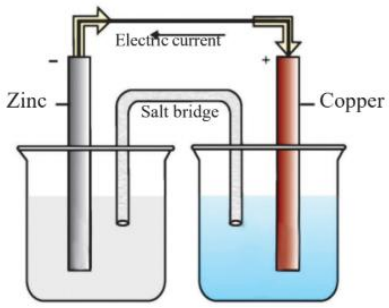
- ✓ A secondary cell can be recharged and used again.
- ✓ Examples of primary cells are
  i. Lead acid cell (used in vehicles)
  ii. Nickel – cadmium cell (used in flash lights)
  iii. Lithium ion battery (Used in smartphones, laptops and electric vehicles)
  Lithium ion battery
- ✓ Lithium ion batteries are used in portable electronic devices due to their
  i. high energy density
  ii. long lifespan
  iii. low self discharge rate.
- ✓ Lithium ion batterries are also used as satellite batteries

Fuel cells

- ✓ Fuel cells are galvanic cells designed to convert the chemical energy produced by the combustion of fuels (hydrogen, methane and methanol) into electrical energy.
- ✓ Fuel cells operate continuously as long as fuel is supplied.
- ✓ One of the **advantages** of fuel cells is that they produce **electricity** with **high efficiency**.

Let us assess

Examine the diagram of a galvanic cell and find out whether the given statements are true or false.



a. As the cell operates, the mass of the zinc rod, Zn(s) decreases.

b. The copper electrode is the anode.

c. Electrons flow from the zinc electrode to the copper electrode through the external circuit.

d. During cell reaction, reduction takes place at the copper electrode.

e. During cell reaction, the concentration of Cu^{2+} decreases.

2. The chemical equation of the reaction taking place in a galvanic cell is given.

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

a. Represent diagrammatically the above galvanic cell.

b. Which electrode has a negative charge?

c. What is the reaction taking place at each electrode?

3. A galvanic cell is made up of the following half cells:

- A magnesium electrode immersed in magnesium sulphate solution.
- A nickel electrode immersed in nickel sulphate solution.

The reactions taking place in the half cells are given.

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

Draw the cell diagram and label the anode, cathode and direction of flow of electrons.

4. When an aqueous sodium chloride solution is electrolysed, the reaction that takes place at the anode is:

(i) $Na^{+}(aq) + e^{-} \rightarrow Na(s)$

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

(iii) $H^{+}(\text{aq}) + e^{-} \rightarrow \frac{1}{2}H_2(g)$

(iv) $Cl^{-}(\text{aq}) \rightarrow \frac{1}{2}Cl_2(\text{g}) + e^{-}$

5. Write the products obtained when the following solutions undergo electrolysis.

(i) Aqueous solution of NaCl

(ii) Aqueous solution of CuSO4 (using copper electrodes)

6. Brass is an alloy of zinc and copper. When brass comes into contact with saline water, corrosion of metal takes place and zinc gradually dissolves in the solution, leaving copper behind. Explain why zinc dissolves in comparison to copper.

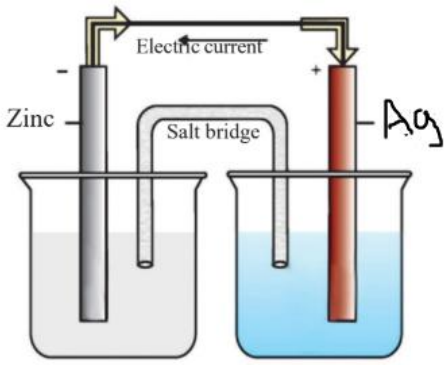
7. List the cells we use in our daily life and classify them into primary and secondary cells.

8. Define fuel cells and write its advantage.

Answers

1. a. True b. False c. True d. True e. True

2. a.



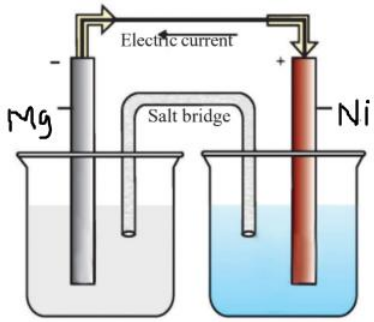
b. Zinc (Anode)

c. Chemical reaction taking place at

Anode **Cathode**

$Zn \rightarrow Zn^{2+} + 2e^-$ $2Ag^+ + 2\,e^- \rightarrow 2Ag$

3.



Anode (negative electrode) – Mg rod

Cathode (positive electrode) – Ni rod

Direction of flow of electrones is from Mg rod (anode) to Ni rod (cathode)

4. (iv) $Cl^-(aq) \rightarrow \frac{1}{2}Cl_2(g) + e^-$

5. i. Chlorine (Cl_2), Hydrogen (H_2), Sodium hydroxide (NaOH)

ii. Pure copper (Cu)

6. Since zinc is more reactive than copper, zinc gets oxidized and enters the solution as zinc ions (Zn^{2+}).

7.

Primary cell	Secondary cell
i. Dry cell	i. Lead acid cell
ii. Button cell	ii. Nickel – cadmium cell
	iii. Lithium ion battery

8. Fuel cells are galvanic cells designed to convert the chemical energy produced by the combustion of fuels into electrical energy.

Advantages

- Operate continuously as long as fuel is supplied.
- Produce electricity with high efficiency.
- Does not cause pollution

Question Bank with Answers

1. **Assertion (A)** : When electricity is passed through an aqueous solution of sodium chloride, sodium is obtained at the cathode.

Reason (R) : The reduction tendency of Na^+ is highier than that of H_2O.

Which among the following is **correct**?

a. Both (A) and (R) are correct, and R is the correct explanation of (A).

b. Both (A) and (R) are correct, and R is not the correct explanation of (A).

c. (A) is correct but (R) is not correct.

d. Both (A) and (R) are not correct.

2. Two activities are given.

(i) Ag rod is dipped in $ZnSO_4$ solution

(ii) Fe rod is dipped in $CuSO_4$ solution

a. In which of these solutions does the displacement reaction take place?

b. Which metal undergo oxidation in the above reaction?

3. Some metals and their salt solutions are given.

$Zn, Fe, Ag, Cu, Mg, ZnSO_4$ solution, $FeSO_4$ solution, $CuSO_4$ solution, $MgSO_4$ solution, $AgNO_3$ solution

[Hint: Reactivity : $Mg > Zn > Fe > Cu > Ag$]

a. Which of the given metals are used to make a galvanic cell with highest cell voltage?

b. Write the chemical equation of the reaction taking place at the cathode in this cell.

c. Which of the given metals acts only as the anode when galvanic cells are made?

4. Molten sodium chloride is a good conductor because its ions can move freely.

a. Which is the gas liberated at the anode during electrolysis of molten sodium chloride?

b. Write the chemical equation of the reaction taking place at the cathode.

5. Electroplating is one of the application of electrolysis. Copper can be coated on an iron bangle by this process.

a. Which is the metal connected to the negative terminal of the battery here?

b. Write the chemical equation of the reaction taking place at the anode.

6. Some galvanic cells and their cell voltages are given in a table. Analyze them and answer the questions.

Galvanic cell	Cell voltage
$Zn - Cu$	1.10 V
$Al - Cu$	2.00 V
$Mg - Cu$	2.72 V

a. Among the electrodes used, which metal has the highest tendency to get oxidised in the redox reaction ?

b. If a galvanic cell is constructed using magnesium and zinc electrodes, from which metal to which metal will be the direction of electron flow ?

c. Write the redox reaction equation that takes place in this cell.

7. The products of electrolysis of aqueous NaCl are

A. Na, H_2 and Cl_2

B. H_2 Cl_2 and NaCl

C. H_2, Cl_2 and NaOH

D. H_2, Cl_2 and HCl

8. a. What is a fuel cell?

b. Write one advantage of fuel cell.

9. A Cu rod is dipped in silver nitrate ($AgNO_3$) solution.

a. What is the change in colour that takes place in the solution?

b. Which metal is oxidised here?

10. Two galvanic cells are given

i) Mg – Cu ii) Zn – Cu

a. Which of them will give higher voltage? Why?

b. Write the direction of electron flow in the first cell.

c. Write the equation of reaction taking place at the anode of the cell (ii).

11. The rcdox reaction in a cell is given below.

$$Mg(s) + 2Ag^+(aq) \rightarrow Mg^{2+}(aq) + 2Ag(s)$$

a. Draw the diagram of this cell.

b. Mark the direction of electric current in this cell.

Answers

1. d. Both (A) and (R) are not correct.

2. a. (ii) Fe rod is dipped in $CuSO_4$ solution

$$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$$

b. **Fe** undergoes oxidation (Reason : Fe is more reactive than Cu)

3. a. Mg – Ag cell

b. $2Ag^+ + 2e^- \rightarrow 2Ag$

c. Mg (Reason : Most reactive metal)

4. a. Chlorine

b. $Na^+ + e^- \rightarrow Na$

5. a. Iron (Fe)

b. $Cu \rightarrow Cu^{2+} + 2e^-$

6. a. Mg

b. Mg rod to Zn rod

c. $Mg + Zn^{2+} \rightarrow Mg^{2+} + Zn$

7. C. H_2, Cl_2 and NaOH

8. a. Fuel cells are galvanic cells designed to convert the chemical energy produced by the combustion of fuels into electrical energy.

b. Fuel cells produce electricity with high efficiency.

9. a. The initially colourless solution gradually turns blue.

b. Copper (Cu)

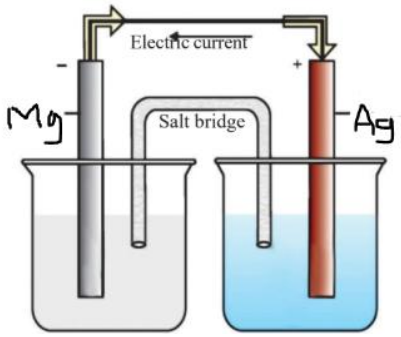
10. a. Mg – Cu cell

Reason : Magnesium is more reactive than zinc.

b. Mg rod to Cu rod.

c. $Zn \rightarrow Zn^{2+} + 2e^{-}$

11. a.



$MgSO_4$ solution $AgNO_3$ solution

b. The direction of flow of electric current is from the silver rod (cathode) to the magnesium rod (anode).