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2024





ADDIS ABABA

2016E.C (2023/2024) ACADAMIC YEAR

HIGH SCHOOL DIVISION

2nd QUARTER CHEMISTRY HANDOUT (ELECTROCHEMISTRY) FOR GRADE 12

Name _____ Sec. _____

UNIT TWO

ELECTROCHEMISTRY

- Electrochemistry deals with the inter conversion of electrical energy and chemical energy.
- It is primarily concerned with two processes that involve oxidation- reduction reactions.
 - The generation of electric current from chemical reaction.

i.e. Chemical reaction Electrical energy

- The use of current to produce chemical changes.

i.e. Electrical Energy chemical changes

Oxidation- Reduction Reactions

- Are also called Redox reactions
- Redox reaction is a reaction in which one or more electrons are transferred. The species that lost electrons is said to be oxidized, and the species that gained the electron is said to be reduced
- Reactions which do not involve transfer of electrons from one species into another are called non-redox reactions.

Example:

Oxidation

NOTE: double displacement reaction is **non- redox reactions**.

- Oxidation of a substance can be defined in several ways:

- Oxidation is the combination of a substance with oxygen.

Examples: →

→

- Oxidation is the removal of hydrogen from a compound

Examples: →

→

- Oxidation is defined as the loss of electrons.

Examples. →

..

→

- Oxidation is defined as an algebraic increase in oxidation number (oxidation state)

Examples. →

→

Reduction

- Since electrons cannot be created or destroyed, oxidation and reduction always occur simultaneously in ordinary chemical reactions, and to the same extent.
- Reduction is defined as

i) Addition of Hydrogen

Example: $\text{H}_2 + \text{H}_2$

$\text{H}_2 + \text{O}$

ii) Removal of oxygen

Example: $\text{CaO} + \text{H}$

iii) Gain of electrons

Example: $+$

iv) An algebraic decrease in oxidation state

Example: $-$

$2 + 2\text{e}^-$

Worked Examples

1. Which of the following reactions are redox rxns?

a)

b)

c)

d)

Solution: Electron transfer reaction is a redox rxn.

a)

$\rightarrow$

▪ Elements in their free existence and molecules of an element have a zero oxidation state

$\rightarrow$

▪ Thus the rxn is redox because;

$\rightarrow$

$\rightarrow$

b)

$\rightarrow$

$\rightarrow$

$\rightarrow$

... Reduction

Self Test

Classify the following reactions as redox and non- redox reactions

A)

$\rightarrow$

B)

$\rightarrow$

C)

$\rightarrow$

D)

$\rightarrow$

E)

$\rightarrow$

Balancing Oxidation- Reduction [Redox] Reactions

⇒ Balancing Redox Reaction involves;

- Balancing the number of reactants and products
- Balancing the net electrical charges in the reactant side to the net electrical charge product side.

i. Balancing Redox Reactions using change in Oxidation- Number Method

Step 4:

Step 5: Balance the remaining substances by Inspection

Self Test

Balance the following chemical equations using oxidation number method.

A)

B)

C)

The Ion- Electron Method for Balancing Redox Reactions

= This method is useful if the net ionic equation is given or no spectator ions are present.

Steps used for balancing redox reaction using ion- electron method are:

Step 1: Identify the element undergoing oxidation and reduction; write separate equations for oxidation and reduction half reactions.

Step 2: Balance the electron oxidized or reduced.

Step 3: Determine the change in oxidation number of element oxidized or reduced

Step 4: Balance any element other than hydrogen and oxygen that may be present.

Step 5: Balance charges by adding H⁺ if the solution is acidic and OH⁻ if the solution is basic.

Step 6: Balance Hydrogen by adding water molecules

Step 7: Balance the number of electrons used for oxidation to those electron used for reduction

Step 8: Add the two half reactions together and eliminate any thing that appears in identical form sides of the equation.

Examples: Balance the following chemical equation using Ion- electron method.

a)

b)

Solution:

Write the oxidation and reduction half rxns

+7

Balance the oxidation half rxn?

Step 1: is balanced

Balance the reduction half rxns.

Step 1:

Step 2: Balance the charge difference by adding

(-6)

(+2) to the reactant side

Step 3: Add water to balance the hydrogen

Step 4: Balance the electrons in oxidation half rxn to the electrons used in reduction half rxn.

Step 5: Combine the two half- reactions

Solution: b. $\rightarrow$ [] (in basic medium)
Write the oxidation and Reduction half reactions.

Step 1: Balance the oxidation half- reaction

Charge: 0... Balanced
Balance the Reduction half- rxn

Step 1 $\rightarrow$ []

Step 2: add to balance the charge difference

(+2) $\rightarrow$ []
-2 + (-2)

add OH to the reactant side

Step 3: The hydrogen is balanced, so balanced for reduction to balanced for oxidation.

$\times$ []
 $\times$ []

Step 4: Combine the half reactions:

$\rightarrow$ []
 $\rightarrow$ [] Balanced

Self Test

Balance the following redox reaction using ion-electron method

A. $\rightarrow$

B. $\rightarrow$

C. $\rightarrow$

Electrolysis of Aqueous solutions

Electrical conductivity is the ability of a substance to transmit electricity.

Generally there are two types of electrical conduction.

To conduct electricity, a substance must contain freely mobile, charged species.

Ionic compounds cannot conduct electricity in the solid state. This is because the ions of the compound are held together by strong electrostatic forces and are in a fixed position in the crystal lattice. Therefore, they are unable to move and carry an electric charge, preventing the conduction of electricity.

i. Metallic conduction

Is also called electronic conduction because the charge carriers are free electrons.

Transmits electricity without undertaking chemical change.

- There is no net movement of metal atoms.
 - Conductivity of metal decreases with increase in temperature due to the enhanced thermal vibration of metal atoms disrupting the movement of electrons passing through them
- Example: Metals and Graphite*

ii. Electrolytic conduction

- Refers to the conduction of an electric current by the motion of mobile ions through a substance called Electrolytes.
- The conductivity of electrolytes increases with increase in temperature. This is due to increase in ionic mobility.
- Causes chemical changes to occur on the surface of the electrodes.
- The strength of an electrolyte depends upon its tendency to dissociate and upon the magnitude of the charge on the ions.
- Electrolytes are molten ionic compounds (pure liquid) or aqueous solution that conduct electricity due to moving ions.
- The positively charged ions (cations) migrate spontaneously toward the negative electrode, while negatively charged ions (anions) move toward the positive electrode.
- Electrodes are usually metal surfaces upon which oxidation or reduction half-reactions occur. Generally, Oxidation occurs at the Anode and reduction occurs at the cathode.

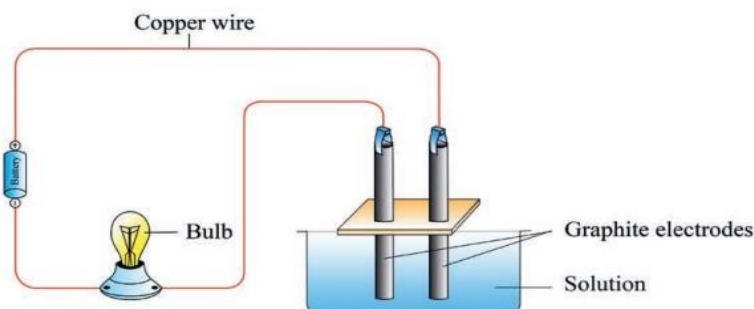
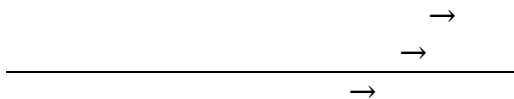


Figure 2. 1: Electrical Conductivity Apparatus

Example:- Electrolysis of molten or fused CaCl_2

-
- Migrates towards the negative electrode, **cathode**.
- Migrates towards the positive electrode, **Anode**.



Worked Examples

Define the following basic terminologies of an electrolytic cell.

- | | |
|----------------------|--------------------------------------|
| a) Electrolytic cell | c) Electrodes |
| b) Electrolysis | d) Overall reaction or cell reaction |

Solution: Electrolytic cell – is a device where electrolysis is taken place

- A typical electrolytic cell contains
- Direct electric source

- An electrolyte
- connecting wires
- electrodes

b) **Electrolysis** is a non-spontaneous redox reaction proceeding under the effect of electric current on electrodes placed in an electrolyte.

▪ is a process by which electrical energy is used to produce chemical changes

c) **Electrodes** are poles (or rods) of carbon or metal at which the current as a flow of electrons enters or leaves an electrolyte

d) **Overall rxn or cell reactions** is a reaction that takes place in the electrolytic cell and is always a redox rxn.

Preferential Discharge

- When two or more ions of similar charge are present under similar conditions in an electrolyte solution, one is preferentially selected by the electrode and discharged.
 - This selection of the electrode for one kind of ions is called preferential discharge.
 - There are different factors influencing the preferential discharge of ions, but three factors are most important.

1. Nature of the Electrode.

- Inert electrodes like Graphite or platinum do not affect the product of electrolysis. These electrodes do not react with ions of the solution.
- Active Electrodes like copper directly take part in the reactions and react with the electrolyte. These electrodes affect the product of electrolysis.

2. Position of the Ions in the Electrochemical series

- Just as certain elements are more reactive than others, ions of these elements vary in their ability to gain or lose electrons. This in turn depends on the atomic or ionic structure of the given atom or ion concerned.
- A series of ions arranged in the order of their ability to gain or lose electrons (tendency to discharge) is known as an electrochemical series.
- The reactivity series is sometimes quoted in the reverse order of standard electrode potentials.

Table : the ease of discharge of some common cations and Anions.

Li^+
 K^+
 Na^+
 Mg^{2+}
 Al^{3+}
 Mn^{2+}
 Zn^{2+}
 Cr^{3+}
 Fe^{2+}
 Cd^{2+}
 Pb^{2+}
 Fe^{+3}
 H^+
 Cu^{+2}
 Ag^+
 Au^{3+}

The ease of discharge increase down ward

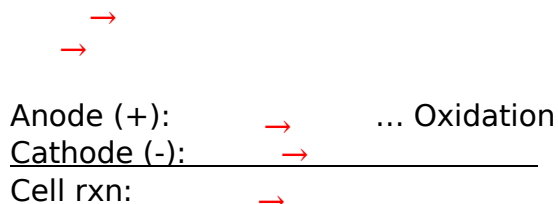
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The ease of discharge increase down ward

- If two or more positive ions migrate to the cathode, the ion lower in the series is discharged preferentially.
Similarly, anions of the lower series are discharged at the anode preferentially.

3. Concentration of the Ions in the Electrolyte

- If an electrolyte contains a higher concentration of ions that are higher in the electrochemical series than those that are lower, then the higher ions get discharged in preference to the lower ones.
Example:- The electrolysis of concentrated KCl in water
- In the solution, there are more K^+ ions than H^+ ions; therefore, K^+ at the cathode, and Cl^- ions at the anode are preferentially discharged.



Write the electrolysis of the following aqueous solutions.

- Concentrated NaCl solution (Brine solution)
- Dilute NaCl solution
- Dilute H_2SO_4 solution
- CuSO_4 solution, using Inert Electrodes.
- CuSO_4 solution, using copper Electrodes.

Self Test

- Write the overall reaction for the electrolysis of dilute NaOH solution.
- What are the substances produced at each electrode in the electrolysis of
 - Concentrated AlBr_3 (aq)
 - Diluted PbCl_2 (aq)
- The electrolysis of concentrated CuF_2 produce;
 - Copper and Oxygen
 - Hydrogen and oxygen
 - Copper and fluorine
 - Hydrogen and Fluorine

Quantitative Aspects of Electrolysis

Faraday's First Law of Electrolysis

States that "the amount of substance consumed or produced at one of the electrodes in an electrolytic cell is directly proportional to the amount of electricity that passes through the cell". If m is the mass of substance liberated or deposited at an electrode due to the passage of charge Q , then according to Faraday's 1st law

Where Z is a proportionality constant and is called electrochemical equivalent but $Q = It$

- ∴ The quantity of electrical charge equivalent to one mole of electrons is called the Faraday, F .
- ∴ One faraday equals one mole of electrons times the charge on an electron.

$$\frac{m}{M} = \frac{Q}{nF}$$

Thus, charge (Q) =
Number of moles of electrons ×

$$m = \frac{Q \times M}{n \times F}$$

where m is given mass
M is molar mass
Q is current
t is time in second.

Example 1: How many grams of copper are produced when 30A of current are passed through copper (II) for an hour? (

Solution: Copper is produced by the reaction;

Faraday's Second Law of Electrolysis

State that “when the same quantity of charge, Q, is passed through different electrolytes, then mass of different substances deposited (, at the respective electrodes will be directly proportional to their equivalent masses (E, ...]

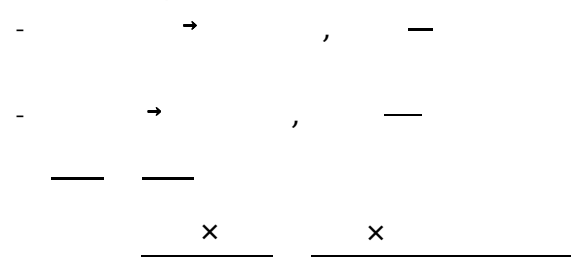
$$\frac{m_1}{E_1} = \frac{m_2}{E_2} = \dots$$

Or , , etc...

Where, m is mass of one substance
E is equivalent mass of substance one
m is mass of second substance
E is equivalent mass of second substance.

Example: Two cells are connected in series. One contains AlCl₃ and the other contains AgNO₃. The electrolytes. What mass of Ag is deposited when 18.0gm of Al is deposited at cathode?

Solution: The equivalents of Aluminum and silver can be obtained from their reduction half rxn



Self Test

- Two cells are connected in series. One contains FeSO₄ and the other contains CuSO₄. Calculate the mass of Cu deposited when 200.0gm of Fe is deposited at the Cathode?
- 0.4 moles of electrons are passed through three electrolytic cells in series, that contain Zn²⁺ ions. How many grams of each metal will be deposited at cathode?

Industrial Application Of Electrolysis Electroplating and Electro refining of Metals

- **Electroplating**- Is the process that involves the coating of a solid surface with a metal application of electrolysis.
 - || Is the coating of less active metal on the surface of more active metal by electrolysis.
 - || This technique is used to protect metals that corrode by the application of a thin coating of that resists corrosion.

Example: House hold articles such as spoons, trays, etc are electroplated to prevent the article from corrosion & to make them look more attractive.
 - || The pure plating metal is used as Anode, and the article to be plated is made to be the cathode.
- Electro refining:** is an industrial electrolytic process in which a sample of impure metal acts as the anode and sample of the pure metal acts as the cathode.
 - || Is purification of metals by using electrolysis.

Example: Metals such as copper, Silver, Aluminum and Gold are refined by electrolysis
 - || Salts of the impure metal (Anode) are used as electrolyte.

Self Test

1. If a silver strip were immersed in aqueous solution containing Cu^{2+} ions, what would you expect to happen?
 - A. Ag would be oxidized
 - B. Cu^{2+} ions would be reduced
 - C. Copper would be deposited on the silver strip.
 - D. No reaction would occur.
2. If an iron ring is plated with zinc metal then which of the following is true?
 - A. It is a voltaic cell and the reaction is spontaneous.
 - B. It is a voltaic cell and the reaction is not spontaneous.
 - C. It is an electrolytic cell and the reaction is spontaneous.
 - D. It is an electrolytic cell and the reaction is not spontaneous.

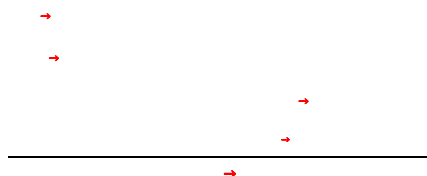
Extraction of Metals

- || Electrolytic reduction is suitable for highly electropositive metals, such as sodium, magnesium and Aluminum.
- || Aluminum is extracted by the electrolysis of molten aluminum oxide to enhance electrical conductivity and to lower the melting point.
- || At the cathode, metal ions are reduced and at the anode, anions are oxidized.

Manufacture of non metals and some compounds

- || Non metals such as Hydrogen, oxygen and Chlorine can be produced by electrolysis at the industrial level.

Example: Write the cathode and anode rxns during the electrolysis of molten mixture of KF and HF .



Worked Examples

Explain the electrolytic reduction of sodium metal from sodium chloride.

Solution: Active metals like Na are strong reducing agents and cannot be reduced by carbon.

- Sodium is obtained by the electrolysis of molten mixtures of about 40% NaCl and 60% CaCl₂ in down cell.

→



CaCl₂ is added to lower the melting point of NaCl from 800°C to 600°C.

Voltaic (Galvanic) Cells

A galvanic cell is a device in which chemical energy is transformed into electrical energy.

The three common types of galvanic cells are

- A. Primary cell
- B. Secondary Cells
- C. Fuel cells

A. Primary Cells:- Are cells which cannot be re-charged. i.e. the electrode reactions cannot be reversed.

Spontaneous chemical reactions generate electricity in galvanic cell, and Anode is negative and cathode is positive in galvanic cell in contrast to electrolytic cell.

Construction of Zn-Cu Galvanic cell

Zn-Cu galvanic cell is also called Daniel cell.

It has two compartments separated by salt bridge.

- || Anode:- is Zn in ZnSO₄ solution
- || Cathode:- is Cu in CuSO₄ solution

Anode rxn:

Cathode rxn:

Cell rxn:-

A voltmeter can be inserted in the circuit to measure the potential difference between the electrodes.

An ammeter can be inserted to measure the current.

In Galvanic cell, electrons are released from the anode and consumed at the cathode, thus electrons flow through the wire from Anode to the cathode. Therefore in contrast to the electrolytic cell, Anode is Negative and the cathode is positive.

Salt Bridge:- is an inverted "U" tube that connects the compartments of the voltaic cell.

- The salt bridge contains a solution of non reactive electrolyte, NaCl or NH₄NO₃ that facilitates the flow of ions between the compartments.

A. Salt Bridge serves three functions

- i. It allows electrical contact between the two solutions.
- ii. It prevents mixing of the electrode solution.

iii. It maintains the electrical neutrality in each half-cell as ions flow in to and out of the salt bridge.

Cell Notation

- Is a short hand representation of a galvanic cell. For example, a Daniel cell in which the electrode concentrations are each 1.0 molar is represented as: $\text{Zn} | \text{Zn}^{2+} (1.0 \text{ M}) || \text{Cu}^{2+} (1.0 \text{ M}) | \text{Cu}$ the anode is written on the left, the cathode on the right and concentrations and other data are given in parentheses.
- The vertical line (|) indicates phase boundaries, and the double vertical line (||) indicates a salt bridge or porous portions.

Worked Examples

- Write the cell notation for a galvanic cell consisting of an Al electrode placed in 1M Al (NO_3^-) solution and Ag electrode placed in 1.0M AgNO₃.

Solution:- $\rightarrow$
 $\rightarrow$

- For the cell notation $\text{Zn} | \text{Zn}^{2+} || \text{Cu}^{2+} | \text{Cu}$ write;
 - The half cell reaction at the anode
 - The half cell reaction at the cathode
 - The overall cell reaction.

Solution: A. Anode (-):- $\rightarrow$
 B. cathode (+):- $\rightarrow$
 C. Overall rxn:- $\rightarrow$
 Self Test

- Write the cell notation of an electrochemical cell constructed by the reaction;
 $\rightarrow$
- Write the oxidation and reduction half reactions of the following cell notations.

- Write the balanced chemical equation for the cell notations of the following galvanic cells.

Standard Reduction potentials

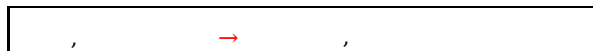
- The cell potential is the difference between two electrode potentials, one associated with the cathode and the other associated with the anode.
- By convention, the potential associated with each electrode is chosen to be the potential for reduction to occur at the electrode.
- Reduction potential is a measure of the tendency of a species to gain electrons in a reduction reaction.
- Oxidation potential is a measure of the tendency of a species to lose electrons in an oxidation rxn.

Standard electrode potentials (electrode potentials at a concentration of 1.0M) are tabulated for reduction rxns and are denoted as;

The cell potential (E_{cell}) is the difference of the cathode minus the standard reduction potential of the anode,

$$E_{cell} = E_{cathode} - E_{anode}$$

Every galvanic cell involves two half-cells and therefore it is not possible to directly measure the standard reduction potential of a half reaction. However taking the reduction of $2H^+ + 2e^- \rightarrow H_2$ as standard reduction potential of exactly 0.0v, we can determine the standard reduction potential of other half-reactions.



Notice the following key points as you use reduction potentials tables:

i) All values are relative to the standard hydrogen electrode (SHE)

ii) The values apply to the half-cell reactions as read in the forward (left to right) direction.

iii) The more positive E° is the greater tendency for the substance to be reduced.

Example: the reaction;

$2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$, is the highest E° value for all the half cell reactions. Therefore Fe^{3+} is the strongest oxidizing agent. At the other extreme, Li^+ is the weakest oxidizing agent. But Li metal is the strongest reducing agent since it is oxidized most readily.

$Li^+ + e^- \rightarrow Li$, which is the lowest E° value implying that Li^+ is the weakest oxidizing agent. But Li metal is the strongest reducing agent since it is oxidized most readily.

iv) Electrode potentials are intensive properties, which means they are not affected by the size of the electrodes or by the amount of solutions present, but depend on the concentration of the solution and the pressure of a gas.

Example:-

And

v) When we reverse a reaction the sign of E° changes, but the magnitude remains the same.

Example:-

vi) The more positive the reduction potential the greater the tendency to accept electrons. As a result, when two half cells are coupled, the reduction with higher (more positive) reduction potential proceeds as reduction, while the other proceeds as oxidation.

vii) The strength of the oxidizing agent (reductant) increases, from top to bottom and the strength of reducing agent increases upwards.

Worked Example:

Calculate standard E_{cell}° of a galvanic cell which is made of Cr electrode in a 1.0M $Cr(NO_3)_3$ solution and an Fe-electrode in 1.0M $Fe(NO_3)_2$.

Solution:- Higher negative values of E° shows that the substance gets oxidized. i.e.



52. What is the standard emf of a Galvanic cell which is made of a Cd electrode in a 1.0 M $\text{Cd}(\text{NO}_3)_2$ solution and Cr electrode in a 1.0 M Cr solution at 25°C?

Solution:

Anode (-): $\text{Cr} \rightarrow \text{Cr}^{3+} + 3\text{e}^-$,

Cathode (+): $\text{Cd}^{2+} + 2\text{e}^- \rightarrow \text{Cd}$

Self Test

1. $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ and $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$
What voltage is produced at standard conditions by combining the half rxns with these standard electrodes potential?
2. Calculate the standard cell potential for the reaction $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$
If $E^\circ_{\text{Zn}^{2+}/\text{Zn}} = -0.76\text{V}$ and $E^\circ_{\text{Cu}^{2+}/\text{Cu}} = +0.34\text{V}$
3. Evaluate the cell potential of the cell that is made for the rxn: $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$
given that $E^\circ_{\text{Zn}^{2+}/\text{Zn}} = -0.76\text{V}$ and $E^\circ_{\text{Cu}^{2+}/\text{Cu}} = +0.34\text{V}$. Is the cell rxn spontaneous?

1.5.4 Electromotive Force (emf)

- Is the driving force that allows electrons to overcome the resistance in the wire and move around the circuit.
- The Emf in a galvanic cell comes from the redox reaction that pushes electrons from the Anode to the cathode through the external circuit.
- Electrons on the negative electrode repel each other and have more potential energy than electrons on the positive electrode.
- The potential difference between the electrodes causes electrons to flow through the circuit from the Anode(-) to the cathode (+) .
- Electrical work (w):- is the product of the emf of the cell and the total charge (in coulombs) that passes through the cell

$$w = nFE$$

- By definition, $w = nFE$ where Faraday (F) is the total charge contained in one mole of electrons.

- The maximum amount of electrical energy that can be obtained from the chemical reaction is calculated as:-

- The maximum amount of useful work that can be obtained from a reaction is denoted as the change in free energy

- For a spontaneous process, ΔG is Negative, so E should be positive.

- For standard state conditions;

↳ We use the relation called the Nernst equation to determine the effect of concentration upon cell potential.

↳ Consider a redox rxn of the type;

$$a\text{Ox} + b\text{Red} \rightleftharpoons c\text{Red} + d\text{Ox}$$
 at equilibrium, thus

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln Q$$

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{RT}{nF} \ln K$$

	K		Reaction under standard state conditions.
Negative		Positive	Favors product formation
Zero		Zero	Reactants and products are equally favored
Positive		Negative	Favors reactants formation

Table 4.3 Relationship among

Worked Examples

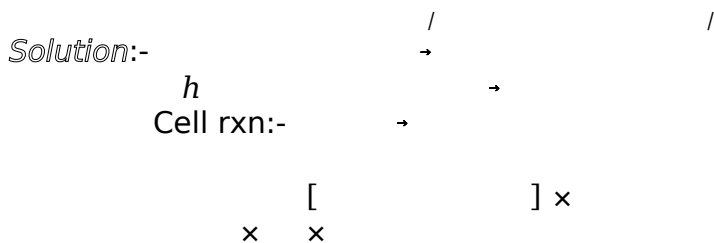
63. Calculate the equilibrium constant for the following reaction at 298k ?

$$\text{Fe}^{2+} + \text{Ce}^{4+} \rightarrow \text{Fe}^{3+} + \text{Ce}^{3+}$$
 given that,

Solution:-

$$\frac{[\text{Fe}^{3+}][\text{Ce}^{3+}]}{[\text{Fe}^{2+}][\text{Ce}^{4+}]}$$

64. Calculate ΔG° for the following reaction at 25°C given that,



Self Test

1. Consider the following standard reduction potentials;



Which pair of substance will react spontaneously?

- A. Sn with Cd^{2+} C. Cd with Sn^{2+}
- B. Sn^{2+} with Cd D. Cd with Sn

2. The cell potential, for an oxidation-reduction reaction was found to be 1.10v. What can be said about this rxn?

- A. At equilibrium C. Non spontaneous
- B. Endothermic D. Spontaneous

3. The equilibrium constant for the reaction with Cd is 1.5×10^{23} . Calculate ΔG° for the reaction at 25°C

- A. -12.6KJ B. -28.9KJ C. 0KJ D. 28.9KJ

4. Calculate the equilibrium constant for the rxn;



- A. 1.5×10^{23} C. 1.5×10^{23}
- B. 1.5×10^{23} D. 1.5×10^{23}

1.5.6 Batteries

A battery is a galvanic cell, or a series of combined galvanic cell, that can be used to provide direct current at constant voltage.

The Dry cell Battery

Are galvanic cells that cannot be recharged.

Zinc wall of the cell serves as Anode.

The Graphite rod through the center of the cell serves as cathode.

The space between the electrodes is filled with an electrolyte consisting of NH_4Cl and ZnCl_2 .

When the cell operates, its half rxns are:

- A build up of ammonia gas cannot be permitted to occur around the cathode since it would discontinue the current. This is prevented by a reaction between NH_3 and Zn^{2+} to form the complex ion $[\text{Zn}(\text{NH}_3)_2]^{2+}$, which crystallizes as chloride salt.

The overall Rxn of a dry cell is

Alkaline Battery

- They can generate current up to 50% longer than the same size of dry cell batteries.
- In Leclanche dry cell, when too large a current is drawn, the ammonia forms a gaseous layer around the carbon cathode. When this happens the voltage drops sharply, and then returns slowly to its normal value of 1.5V.
- "Alkaline" dry cells avoid the limitations of a Leclanche dry cell by using NaOH or KOH instead of acidic salt NH_4Cl .
- The cell potential of an alkaline battery doesn't decline under high current loads, because no gas is formed.

Nickel -Cadmium Battery

- Is rechargeable and is used in battery-operated tools and portable computers.
- The cadmium metal serves as Anode
- The NiO_2 serves as cathode.

Over all rxn:-

N.B:- The solid products of this rxn, Ni(OH)_2 and Cd(OH)_2 , adhere to the electrodes. This permits the reactions to be readily reversed during charging.

The lead storage Battery

- Is mainly used in automobiles, which consist of six identical cells joined together in series.
- Is a secondary rechargeable cell.
- Each cell has a lead (Pb) Anode.
- The cathode is made of PbO_2 packed on a metal plate.
- Sulphuric acid is used as an Electrolyte.

Over all rxn:-

Fuel Cells

- Are Galvanic cells in which the reactants are continuously supplied?
- Fuel cells do not require recharging, but operate continually as far as there is a continuous supply of fuel and oxygen from the air.
- A hydrogen-oxygen fuel cell uses:
 - Inner electrodes as anode and cathode

- KOH and NaOH as electrolytes

Cathode rxn:

Over all rxn:-

N.B:- The cell is maintained at high temperature, and the water produced by the cell reaction evaporates as it is formed.

Self Test

- Which of the following statements is true?
 - A battery is rechargeable
 - A battery produces electricity spontaneously
 - Generally a battery has no liquid components
 - A battery produces the same amount of electricity regardless of composition.
- The cathode material of a lead storage battery is
 - Pb
 - PbO₂
 - Graphite
 - PbSO₄
- Which of the following is false about lead storage battery?
 - It provides a total of 12 voltage
 - Its electrolyte is dilute H₂SO₄
 - Pb is used as cathode
 - PbSO₄ is produced at both electrodes

Corrosion of Metals

- Corrosion is the deterioration of metals by an electrochemical process.
- Corrosion causes enormous damage to buildings bridges, ships and cars.
- The most familiar example of corrosion is rusting of iron.
- Some of the major steps in rusting of iron are

Anode:

Cathode:-

The overall rxn:

Worked Examples

- Aluminum does not rust as iron does? Why

Solution:- Aluminum has much greater tendency to oxidize than does iron. However, the layer of insoluble aluminum oxide (Al₂O₃) that forms on the surface when the metal is exposed to air serves to protect aluminum from further corrosion.

- Metals such as Magnesium and zinc also prevent further oxidation by forming a coat of their oxides. Such metals are called self-protecting metals.

- What are the methods used to prevent rusting of iron?

Solution:- Some of these methods are

- Painting
- Oiling or greasing
- Using adherent oxides
- Alloying
- Coating with another metal
- Using preventive anode.