SANGAM SKM COLLEGE - NADI
LESSON NOTES - WEEK 1
YEAR 13
CHEMISTRY

Strand: 3: Physical Chemistry
Sub-Strand: 3.1 Electrochemistry
Content Learning Outcome: Study Redox titration and limiting reagent calculation

Redox Titration

- Type of titration based on a redox reaction between the analyte and a titrant.
- A **titrant** - reagent of known concentration and volume used in the titration.
- An **analyte** - any substance undergoing analysis.
- Redox titration utilises redox reactions to determine the amount of oxidising or reducing agent in a sample.
- It may involve the use of a redox indicator.

Example:

25.30 mL of 0.025 mol L-1 potassium permanganate solution (KMnO4) reacted with 25.00 mL of an iron(II) solution acidified with dilute sulphuric acid.
i. Calculate the concentration of the iron(II) solution.
ii. How do you recognise the end-point in the titration?

Solution:

i. Derive the balanced redox equation:

$MnO_4^-{}_{(aq)} + 8H^+{}_{(aq)} + 5Fe^{2+}{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)} + 5Fe^{3+}{}_{(aq)} + 4H_2O_{(l)}$

Calculate the number of moles of $KMnO_4$:

$n = c \times V$
$= 0.025 \text{ mol } L^{-1} \times 0.0253 \text{ L}$
$= 6.325 \times 10^{-4}$ moles

Using the mole ratio, calculate the number of moles of Fe^{2+}:

$MnO_4^- : Fe^{2+}$
1 : 5
6.325 × 10-4 : x
$x = 3.1625 \times 10^{-3}$ moles

Hence, calculate the concentration of Fe^{2+}:

$c = n/ V$
$= 3.1625 \times 10^{-3} \text{ mol}/0.025 \text{ L}$
= **$0.13 \text{ mol } L^{-1}$**

ii. The end point is when the first permanent faint pink colour appeared due to excess of $KMnO_4$.

Exercise:

1. To determine the percentage of iron in razor blade, a 0.15 g sample of razor blade was dissolved in excess dilute H2SO4. When the solution was titrated with KMnO4 solution, 20 mL of 0.01 mol L-1 KMnO4 solution was used as shown in the reaction below.

$5Fe^{2+}{}_{(aq)} + MnO_4^-{}_{(aq)} + 8H^+{}_{(aq)} \rightarrow 4H_2O_{(l)} + 5Fe^{3+}{}_{(aq)} + Mn^{2+}{}_{(aq)}$

Note: [*M*: Fe = 56 g mol-1]

a. Calculate the number of moles of KMnO4 used in the titration.
b. Determine the number of moles of iron present in the razor blade solution.
c. Calculate the mass of iron present in the razor blade sample.

Limiting Reagent

- Is a reactant which is completely consumed in a reaction.
- It determines the extent to which a reaction will proceed and how much product will form.
- Is the reactant which has the least mole

Example:

Iron in the form of steel wool can be easily oxidised using copper sulphate as the oxidising agent. In an experiment a student added 1.95 g of steel wool into 50.00 mL of 1.00 mol L^{-1} solution of copper sulphate. Calculate the mass of copper produced. (M: Fe = 56 g mol^{-1}; Cu = 64 g mol^{-1})

Solution:

Steps:

1. Write a balanced equation for the reaction:

$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$

2. Determine the limiting reagent: Calculate the moles of the reactants.

Moles of Fe = m/M
$= 1.95\ g / 56\ g\ mol^{-1}$
= 0.0348 moles

Moles of $CuSO_4$ = c × V
$= 1.00\ mol\ L^{-1} \times 0.05\ L$
= 0.05 moles

Mole Ratio: Use the moles of one of the reactants to determine the expected moles of the other reactant.

Fe : $CuSO_4$
1 : 1
0.0348 : x
x = 0.0348 moles

Moles of $CuSO_4$ required = 0.0348 moles

From the balanced equation, the mole ratio Fe : $CuSO_4$ is 1:1, thus for 0.0348 moles of Fe there should be 0.0348 moles of $CuSO_4$ but 0.05 moles of $CuSO_4$ is present which is in excess. Hence, Fe is the limiting reagent.

3. Use the moles of the limiting reagent present to calculate the moles of copper (Cu):

Fe : Cu
1 : 1
0.0348 : x
x = 0.0348 moles

Moles of Cu = 0.0348 moles

4. Hence find the mass of the copper (Cu):

$m = n \times M$
$= 0.0348\ mol \times 64\ g\ mol^{-1}$
= **2.23 g**

Exercise:

The reaction between silver nitrate and ferric chloride is represented by the equation:

$3AgNO_{3(aq)} + FeCl_{3(aq)} \rightarrow 3AgCl_{(s)} + Fe(NO_3)_{3(aq)}$

A solution containing 18.00 g $AgNO_3$ was mixed with a solution containing 32.40 g $FeCl_3$ to give 10.20 g of AgCl. Determine the limiting reagent in the reaction given above.

SANGAM SKM COLLEGE - NADI
LESSON NOTES - WEEK 2 & 3
YEAR 13
CHEMISTRY

Strand: 3: Physical Chemistry
Sub-Strand: 3.1 Electrochemistry
Content Learning Outcome: Investigate the processes involved in an electrochemical cell.

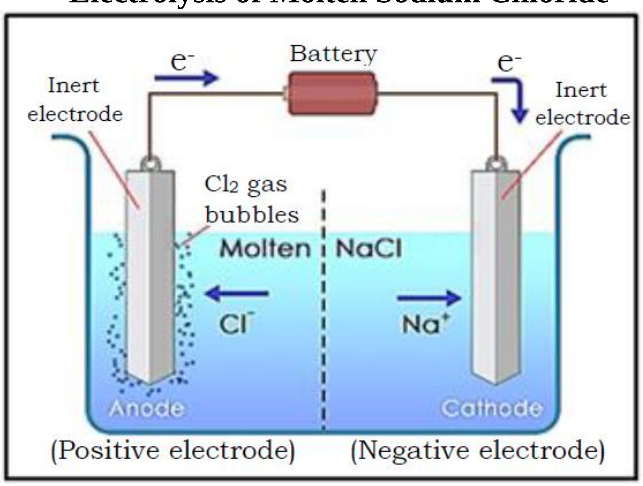
Electrochemical Cells

- system in which electron transfer occurs between the two half-reactions: oxidation and reduction
- two types of electrochemical; **electrolytic cell** and **galvanic cell.**
- **electrolytic cell**- energy from an applied voltage is used to bring about a non-spontaneous oxidation-reduction reaction.
- **galvanic cell**- produces electrical energy from spontaneous redox reactions taking place within the cell.

1. Electrolytic Cells *(recap year 11 and 12)*

- undergoes a redox reaction when electrical energy is applied from an external source such as battery.
- uses an electric current to split a compound into its respective ions by the process of electrolysis.
- made up of two electrodes connected to a DC power supply and immersed in an electrolyte which is a solution or liquid that conducts electricity.
- electrolyte contains ions that are free to move and is either a molten salt or an aqueous solution of an ionic salt.

Example: **Electrolysis of Molten Sodium Chloride**



Cathode reaction (reduction): $Na^+_{(aq)} + e^- \rightarrow Na_{(s)}$
Anode reaction (oxidation): $2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$

2. Galvanic Cell (Voltaic Cell)

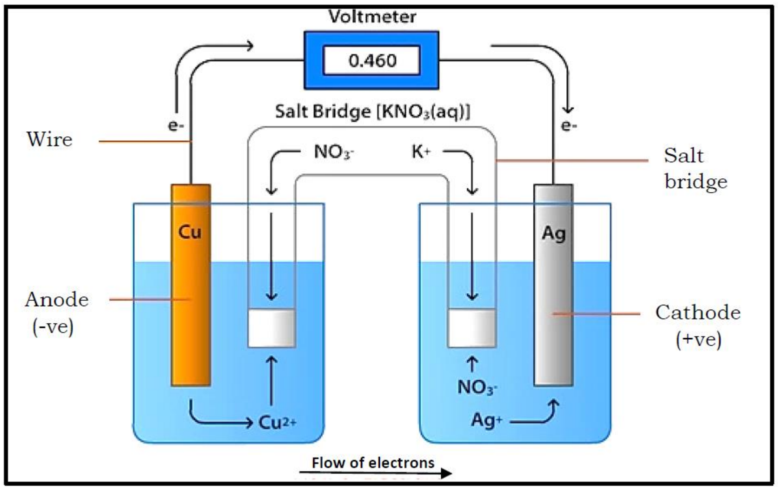
- electrochemical cell that produces electrical energy from spontaneous redox reactions taking place within the cell.

- consists of two separate compartments called **half-cells** which are connected by a salt bridge.
- anode is the negative electrode where oxidation occurs and cathode is the positive electrode where reduction occurs.

Salt Bridge

- conducting medium that allows ions to pass from one half-cell to the other to keep the half-cell electrically neutral.
- can be constructed using a glass U-tube filled with KI, KNO_3 or KCl and fitted with porous plugs at either so that it does not react with any of the chemicals in the cell.

Example:



Anode reaction (oxidation): $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$

Cathode reaction (reduction): $Ag^+_{(aq)} + e^- \rightarrow Ag_{(s)}$

Similarities between Galvanic and Electrolytic Cell

- Electron flow is always from anode to cathode.
- Both contains two electrodes dipped into electrolytic solution(s).
- In both cells, oxidation occurs at the anode and reduction occurs at the cathode.

Differences between Galvanic and Electrolytic Cell

Electrolytic Cell	**Galvanic Cell**
Cathode is negative (reduction occurs here)	Cathode is positive (reduction occurs here)
Anode is positive (oxidation occurs here)	Anode is negative (oxidation occurs here)
Electrical energy is supplied using battery or any electrical source	Electrical energy is produced through a spontaneous reaction
Converts electrical energy into chemical energy	Converts chemical energy into electrical energy
Redox reaction is not spontaneous and electrical energy has to be supplied to initiate the reaction	Redox reaction is spontaneous and is responsible for the production of electrical energy
Both the electrodes are placed in the same beaker in the electrolyte solution or a molten electrolyte	The two half cells are set-up in different beakers with electrolyte solutions and are connected through the salt bridge or porous partition
Absence of a salt bridge	Presence of a salt bridge

Cell Notation

- is a shorthand way of representing a galvanic cell.

Example:

Electrode 1 / Electrolyte 1 // Electrolyte 2 / Electrode 2

ANODE (Oxidation)   SALT BRIDGE   CATHODE (Reduction)

$Zn_{(s)}$ / $Zn^{2+}_{(aq)}$ // $Cu^{2+}_{(aq)}$ / $Cu_{(s)}$

Note:

- The components of the half-cell undergoing oxidation (anode) are written on the left hand side and the components of the half-cell undergoing reduction (cathode) are always written on the right hand side of the notation.
- A single vertical line is used to represent a phase boundary.

For example, a single line between $Zn_{(s)}$ and $Zn^{2+}_{(aq)}$ indicates that solid zinc is a different phase from aqueous Zn^{2+}.

- Species of the same phase are separated by a comma.

For example: Pt/$I^-_{(aq)}$, $I_{2(aq)}$ // $BrO_3^-{}_{(aq)}$, $Br^-_{(aq)}$/ Pt

$I_{2(aq)}$, $I^-_{(aq)}$ and $BrO_3^-{}_{(aq)}$, $Br^-_{(aq)}$ are separated by commas as they are present in the same phase and platinum (Pt) is the inert solid electrode used to conduct electricity.

- A double vertical line indicates the presence of a salt bridge.

Example 1

Write the cell notation for the following reaction.

$$Cu_{(s)} + 2Ag^+_{(aq)} \rightarrow Cu^{2+}_{(aq)} + 2Ag_{(s)}$$

Solution

Half reactions are:

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

Reduction: $2Ag^+_{(aq)} + 2e_- \rightarrow 2Ag_{(s)}$

Cell Notation: $Cu_{(s)}$ / $Cu^{2+}_{(aq)}$ // $Ag^+_{(aq)}$ / $Ag_{(s)}$

Example 2

Write the cell notation for the following reaction.

$$2Cl^-_{(aq)} + 2Fe^{3+}_{(aq)} \rightarrow Cl_{2(g)} + 2Fe^{2+}_{(aq)}$$

Solution

Half reactions are:

Oxidation: $2Cl^-_{(aq)} \rightarrow Cl_{2(g)} + 2e^-$

Reduction: $2Fe^{3+}_{(aq)} + 2e^- \rightarrow 2Fe^{2+}_{(aq)}$

Cell Notation: Pt/ $Cl^-_{(aq)}$/ $Cl_{2(g)}$ // $Fe^{3+}_{(aq)}$, $Fe^{2+}_{(aq)}$/ Pt

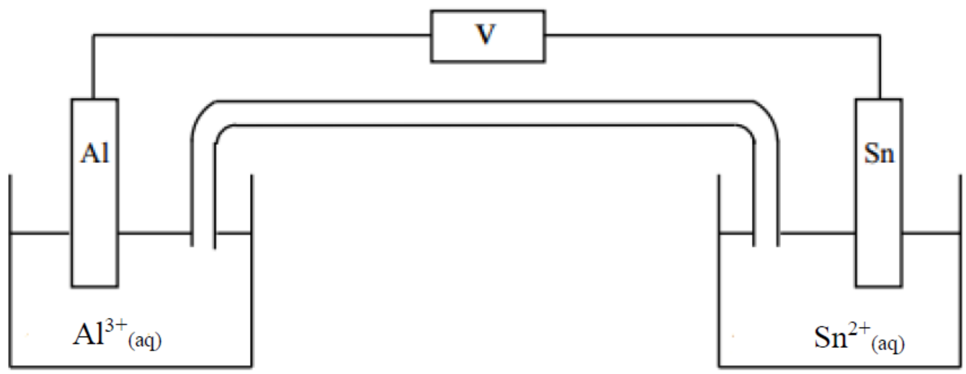
Note:

Since the reactants here are all present as gaseous and aqueous solution, an inert electrode such as platinum is used to provide electrical contact and surface for the redox reaction to take place.

Exercise:

1. Consider the following galvanic cell.

$$Al_{(s)} / Al^{3+}_{(aq)} // Sn^{2+}_{(aq)} / Sn_{(s)}$$



a. Label the anode and the cathode using the cell notation given.
b. In which direction will electrons flow?
c. In which half-cell will electrons enter the cell? At which electrode the electrons are consumed?

d. What is the purpose of the salt bridge?

e. In which direction do cations within the salt bridge move to maintain charge neutrality?

2. The following spontaneous reaction occurs when metallic zinc is dipped into a solution of copper sulphate.

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

a. What are the half-cell reactions?

b. Write the cell notation for the above reaction.

3. Write the anode and cathode half-reactions for the following galvanic cell.

$$Al_{(s)} / Al^{3+}_{(aq)} // Pb^{2+}_{(aq)} / Pb_{(s)}$$