

Chapter #2

Topic:

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Electrochemistry

Electrochemistry: Branch of chemistry that deals with study of interconversion of Electrical energy to chemical energy and vice versa.

Electrical energy $\rightleftharpoons$ Chemical energy

- Electrochemical processes include redox (oxidation-reduction) reactions.

Oxidation - Reduction (Redox): The Reaction in which oxidation and reduction occurs simultaneously is called redox reactions.

Difference between oxidation and reduction.

Oxidation	Reduction
Addition of oxygen is called oxidation.	Addition of hydrogen is called reduction.
Removal of hydrogen is also known as oxidation.	Removal of oxygen is also known as reduction.
In oxidation electrons are lost.	In reduction electrons are gained.
Increase in oxidation state.	Decrease in oxidation state.
Increase in no. of bond of C with oxygen.	Decrease in no. of bond of C with hydrogen.

Decrease in no. of bond of C with hydrogen.

Increase in no. of bond of C with hydrogen.

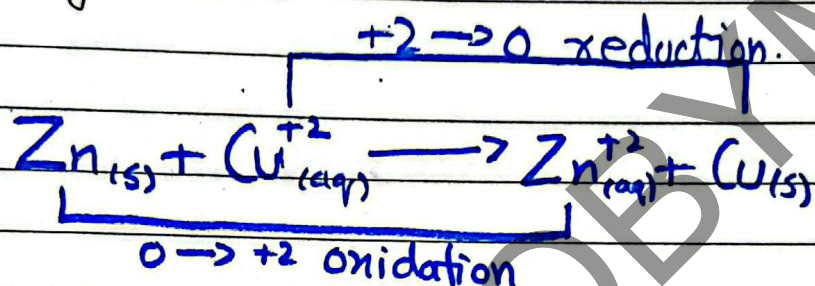
Example #2.1

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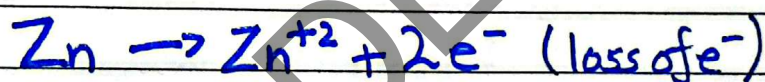
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When a piece of zinc metal is dipped in an aqueous solution of CuSO_4 , it is observed that a dark brown layer of copper begins to form at zinc surface...



In terms of loss and gain of Electron:



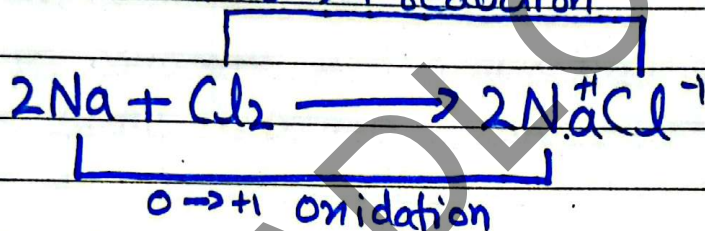
In this reaction, zinc metal loses two electrons and changes into Zn^{+2} ions while Cu^{+2} ions gain two electrons and give copper metal. The two processes taking place simultaneously are called oxidation-reduction reactions.

Oxidation Number:

"The oxidation number (oxidation state) is defined as the apparent charge, positive or negative, which an element would have in a compound."

Oxidation - Reduction in Terms of Change in Oxidation Number.

- Oxidation and reduction can be explained through changes in oxidation number.
- Oxidation means increase in oxidation number. (e^- loss).
- Reduction means decrease in oxidation number. (e^- gain).
- In the reaction: $0 \rightarrow -1$ reduction



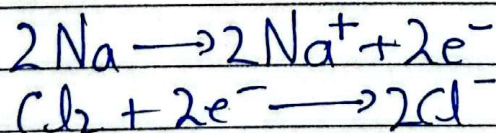
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Explanation:

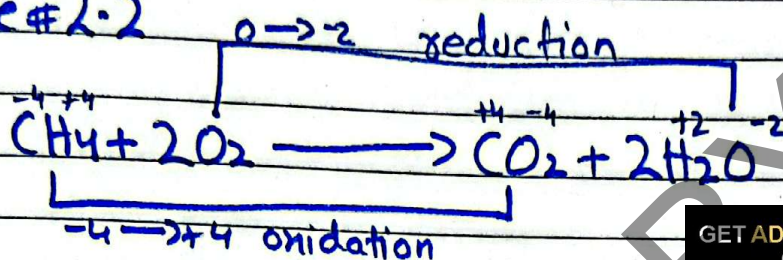
- Sodium (Na) has oxidation number 0 in element form.
- In NaCl, Na becomes +1, which means it **lost one electron**, so it is **oxidized**.
- Cl_2 has oxidation number 0 in element form.
- In NaCl, Cl_2 becomes -1, which means it **gained one electron**, so it is **reduced**.



Oxidation States of some Elements.

SO_4^{-2} sulphate	NO_3^{-1} nitrate	PO_4^{-3} phosphate	CO_3^{-2} carbonate	Cl^{-1} Chloride
	S^{-2} Sulphide	OH^{-1} hydroxide.		

Example #2.2



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Oxidation:

Carbon is oxidized in this reaction because its oxidation state increases from -4 to +4 hence it is oxidized.

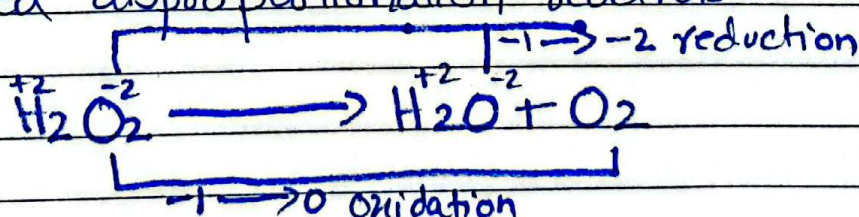
Reduction:

Oxygen is reduced in this reaction because its oxidation state decreases from 0 to -2 hence it is reduced.

Disproportionation Reaction:

"In some redox reactions, the same element is oxidised and reduced simultaneously, Such reaction are called disproportionation reactions."

Example:



In this reaction same oxygen atom is oxidized and reduced.

$-1 \rightarrow 0$ oxidation

$-1 \rightarrow -2$ reduction

so, this is a disproportionation reaction.

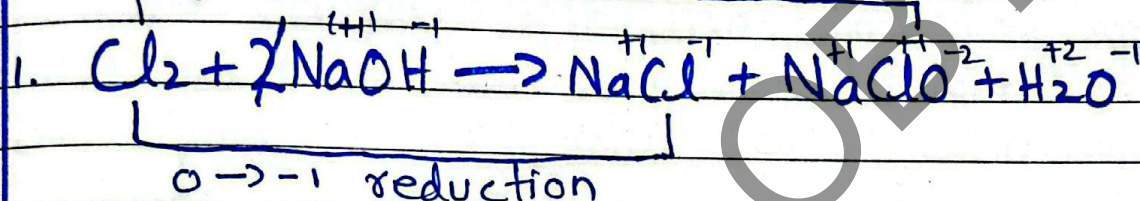
Concept assessment #2.1

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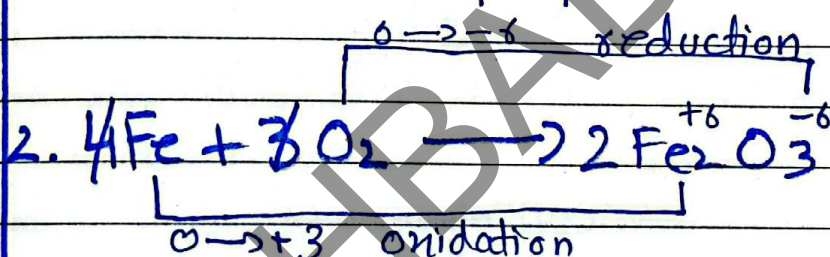
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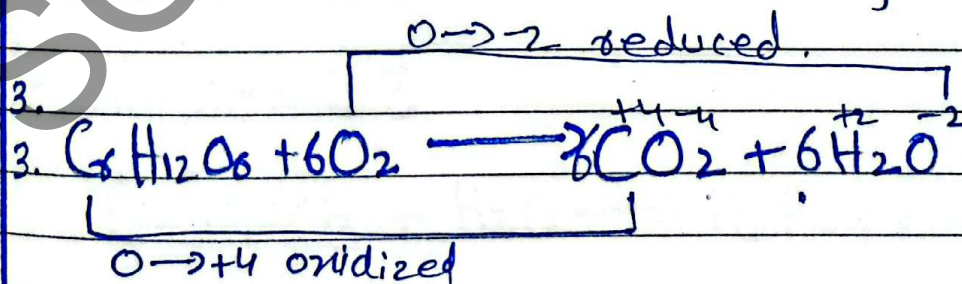
Use the oxidation number change method to identify the atoms undergoing oxidation or reduction in the following redox reactions.



In this reaction Cl_2 is oxidized and reduced so this is a disproportionation reaction.



In this reaction Fe is oxidized because its O.S increases by $0 \rightarrow +3$, and O_2 is reduced because its O.S decreases by $0 \rightarrow -2$.



In this reaction C is oxidized because its O.S

increases by 0 to +4, and O is reduced because its o.s is decreased by 0 to -2.

Balancing of Equation.

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- 1) Redox Method (Oxidation Number Method).
- 2) Ion Electron Method.

① Redox Method (Oxidation Number Method).

Based on principle that in redox reaction, total no. of e^- lost by any element must be equal to the total no. of e^- gained by another.

$$\text{no. of } e^- \text{ gain} = \text{no. of } e^- \text{ loss.}$$

Steps:

- ① Assign oxidation state to each element.
- ② Identify the element changing oxidation state.
- ③ Draw bridge between elements and show e^- loss and gain.
- ④ Multiply e^- loss and gain by a number to get common.
- ⑤ Inspection Method H & oxygen.

Example #2.3



Step 1: Assign O.S to each element.

$$\text{P} + \text{HNO}_3 + \text{H}_2\text{O} \longrightarrow \text{H}_3\text{PO}_4 + \text{NO}$$
$$P + H\overset{+5}{N}O_3 + H_2O \longrightarrow H_3\overset{+5}{P}O_4 + \overset{+2}{N}O$$
$$p^0 \rightarrow p^{+5}$$

5e⁻ loss (oxidation)

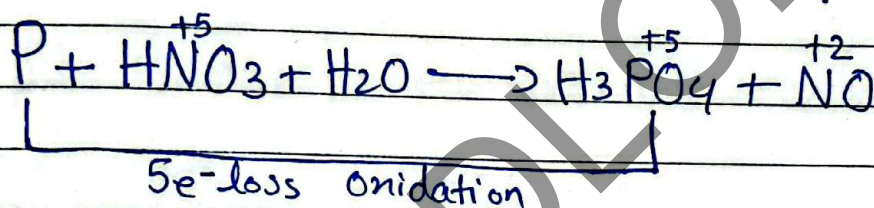
$$N^{+5} \rightarrow N^{+2}$$

$3e^-$ gain (reduction)

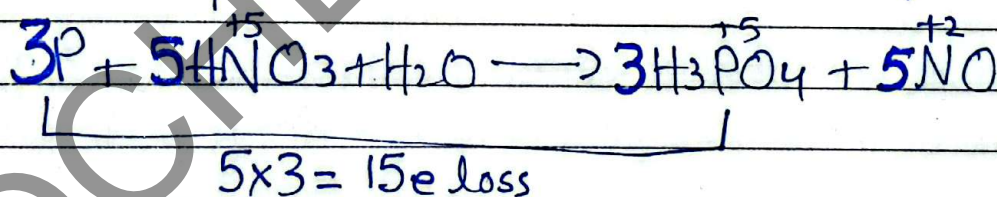
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$3e^-$ gain reduction.



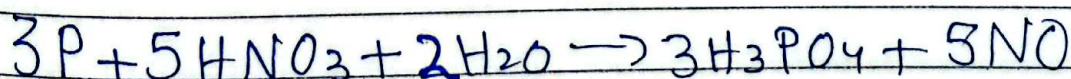
Step 4: Equalize no. of e^- loss and gain by multiply.



→ coefficient 3 with P and H_3PO_4 .

→ add coefficient 5 with HNO_3 and NO .

Step 5: Balance rest of equation by inspection method.



by multiplying coefficient of H_2O by 2, we will get the equation balanced.

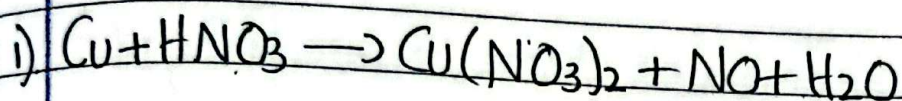
Concept Assessment #2.2

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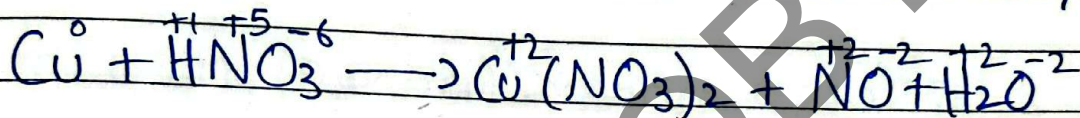
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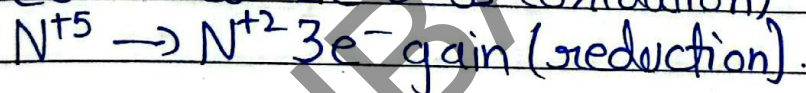
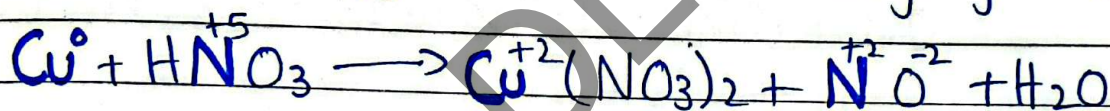
Using the redox method balance the following equation.



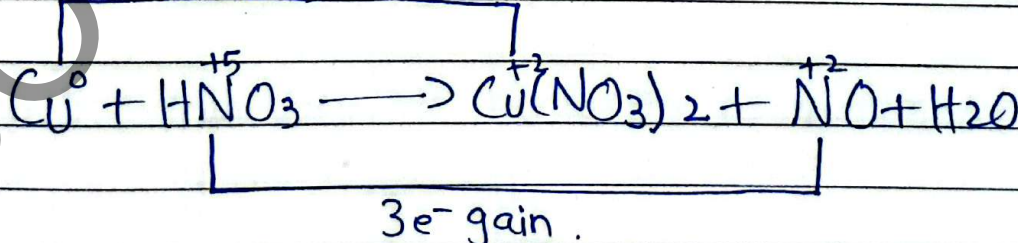
Step 1: Assign oxidation number to all elements/atom.



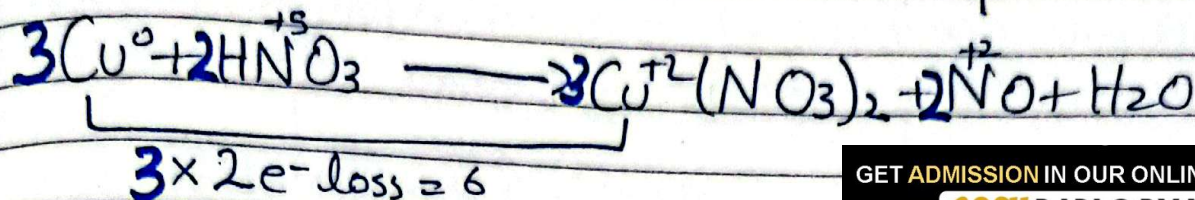
Step 2: Identify element with changing O.S.



Step 3: Draw bridge between atom whose O.S. changed by the number of e^- gained or lost.



Step 4: Equalize no. of e^- loss and gain by multiply.



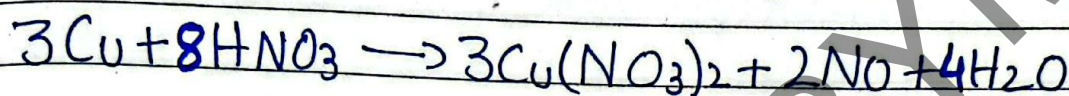
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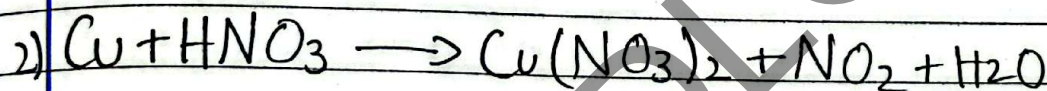
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→ coefficient 3 with Cu and Cu^{+2}
 → add coefficient 2 with HNO_3 and NO.

Step 5: Balance rest equation by inspection method.



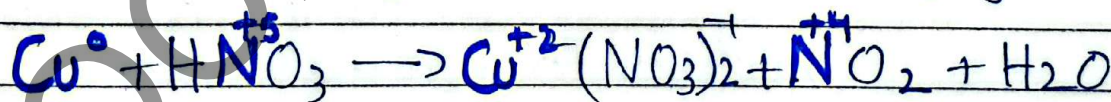
by multiplying (equation)^x coefficient of H_2O by 4 we will get the balanced equation.



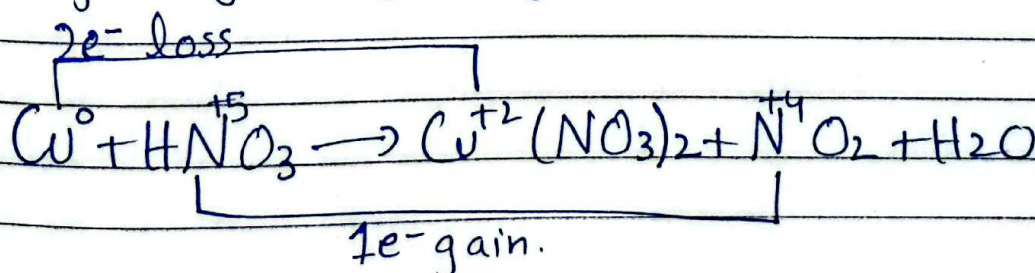
Step 1: Assign oxidation number to each element.



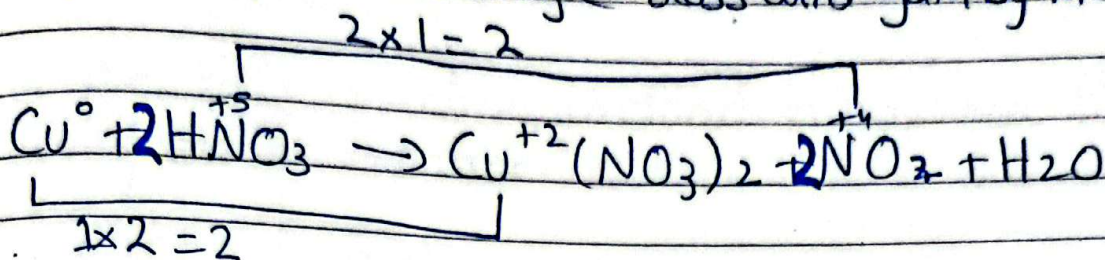
Step 2: Identify element with changing O.S



Step 3: Draw bridge between atom whose O.S changes by no. of e^- gain or lose.



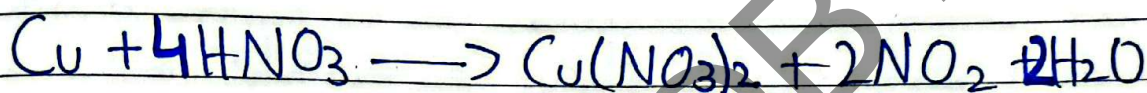
Step 4: Equalize no. of e^- loss and gain by multiply.



→ coefficient 2 with HNO_3 and NO_2 .

→ coefficient 1 with Cu and Cu^{+2}

Step 5: Balance the rest equation by inspection method.



by multiply 2 with HNO_3 and H_2O we get the balanced equation

② Balancing by ion-electron method.

Steps:

- ① balance each half reaction other than oxygen & hydrogen.
- ② Balance oxygen by adding H_2O .
- ③ Balance hydrogen by adding H^+ ions.
- ④ Balance charge of equation by adding e^- .
- ⑤ loss of e^- = gain of e^- by multiplying with suitable digit.
- ⑥ Add two equations
- ⑦ Check equation and balance total no. of charges on both sides.

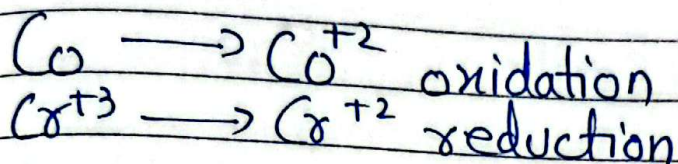
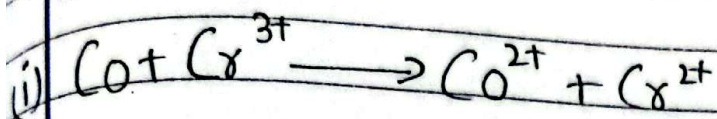
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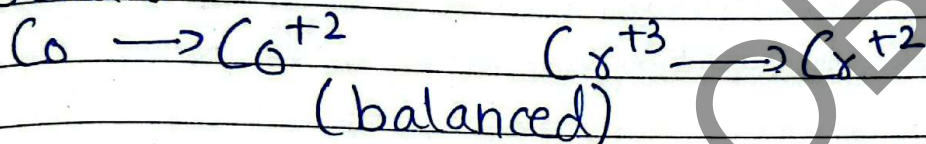
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Concept Assessment #2.3

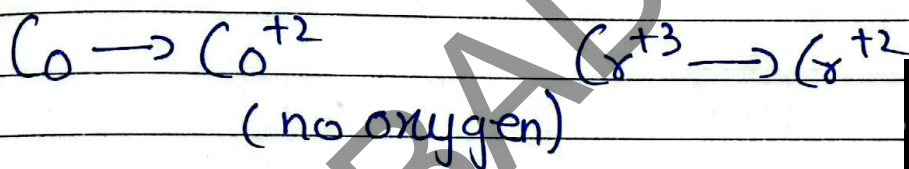
Use the half reaction method to balance the following redox reactions.



Step 1: Balance each half reaction other than O and H.



Step 2: Balance oxygen by adding H_2O

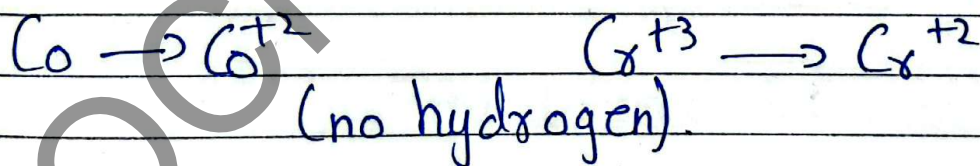


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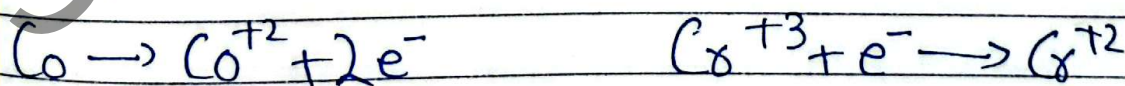
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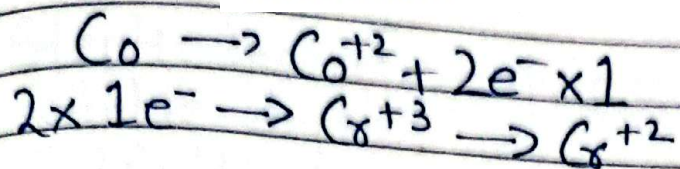
Step 3: Balance hydrogen by adding H^+ ions



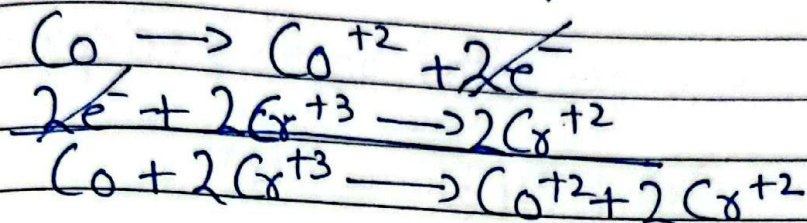
Step 4: Balance charge of equation by adding e^- .



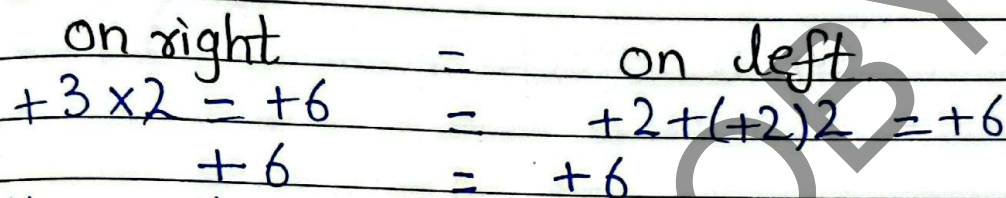
Step 5: Loss of e^- = gain of e^- by multiplying with suitable digit.



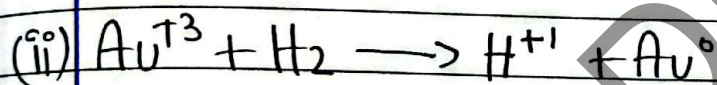
Step 6: Add two equation.



Step 7: Check equation balance on both sides.



Hence the equation is balanced.



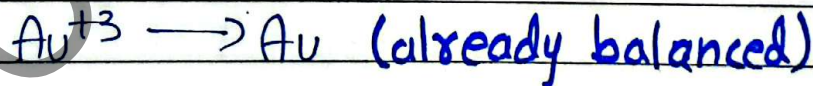
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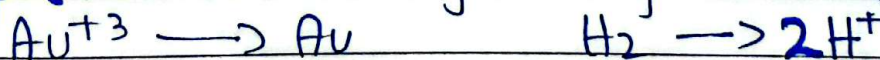
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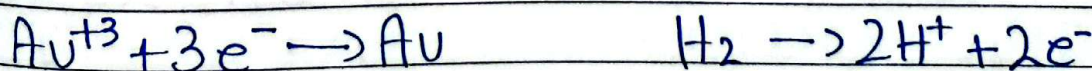
Step 1: Balance each half reaction other than O and H



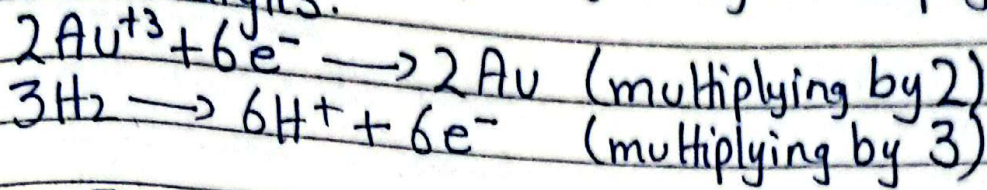
Step 2: Balance H by adding H^+ ions.



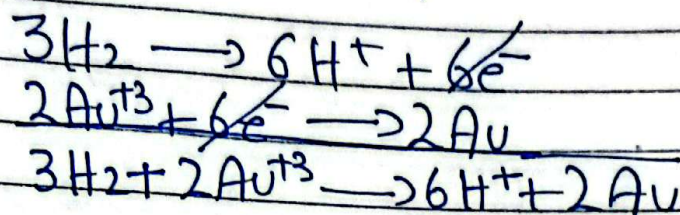
Step 3: Balance the charge by adding electrons.



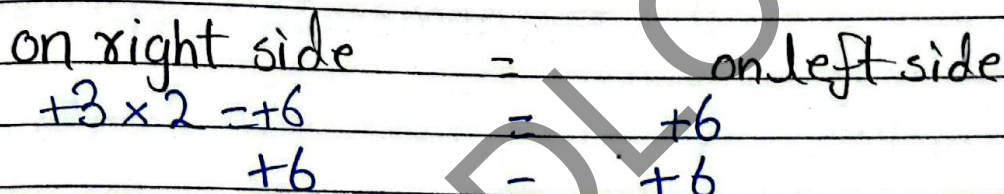
Step 4: lose of e^- = gain of e^- by multiplying with suitable digits.



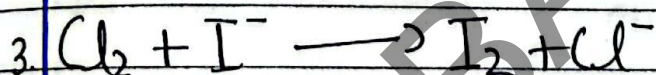
Step 5: Adding both equations.



Step 6: Check if equation is balanced on both sides.



So, the equation is balanced.

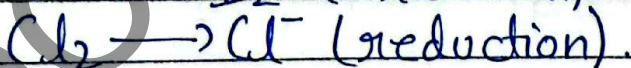


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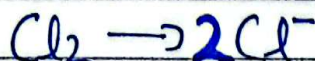
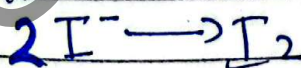
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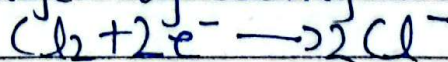
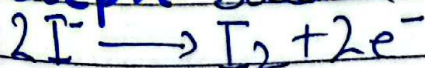
Half reactions:



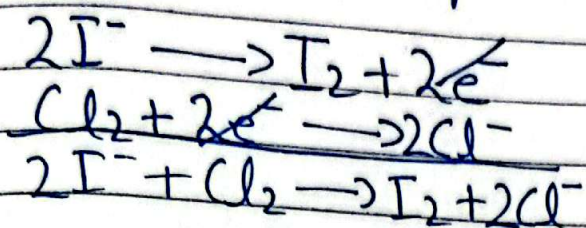
Step 1: Balance the half reactions except O and H.



Step 2: Balance the charges by adding e^-



Step 3: Add both equations.



Step 4: Check if both equations are balanced on both sides.

left side = right side
2 = 2

So, the equation is balanced.

Concept Assessment #2.4

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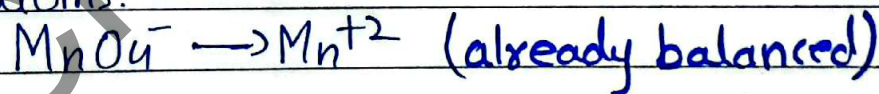
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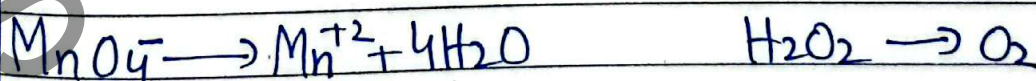
Construct a redox equation from following half equations.



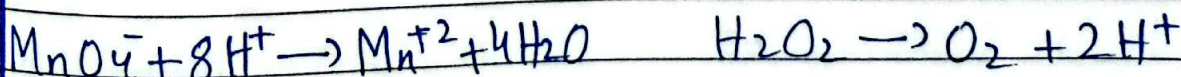
Step 1: balance the half reactions except O and H atoms.



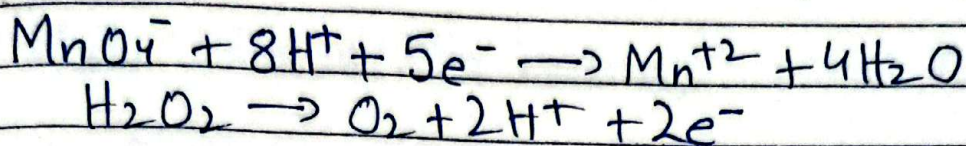
Step 2: balance oxygen atoms by adding water.



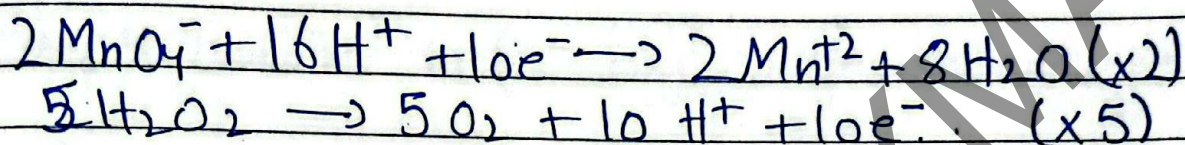
Step 3: balance H atoms by adding H^+ ions.



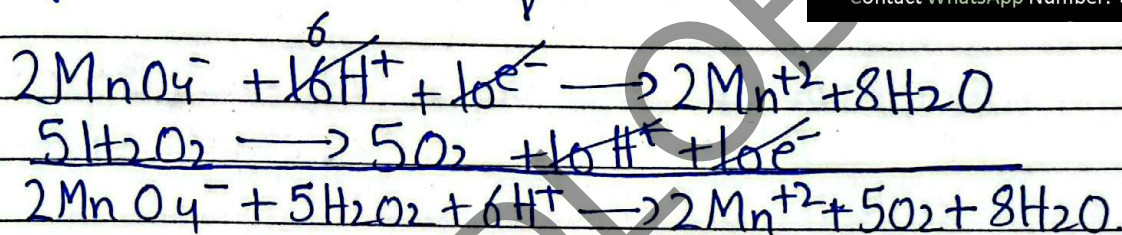
Step 4: Balance the charges by adding e^- .



Step 5: loss of e^- = gain of e^- by multiplying with suitable digits.



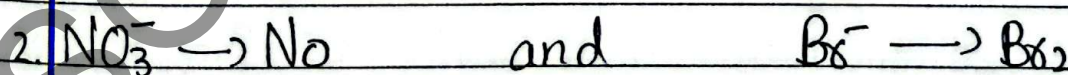
Step 6: Add both equations.



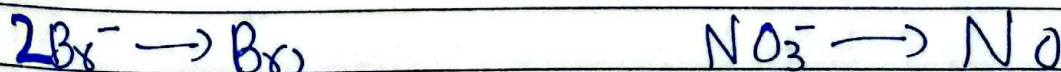
Step 7: Check if both equations are balanced on both sides.

$$\begin{array}{rcl} 5(6) + 6 + (-2) & = & 5(6) + 4 + 8(0) \\ +4 & = & +4 \end{array}$$

So, the equation is balanced.



Step 1: balance the half reactions except O and H atoms.

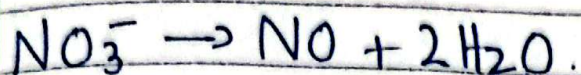


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Step 2: Balance the O-atoms by adding H_2O

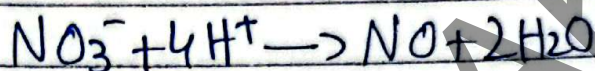
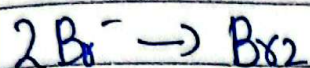


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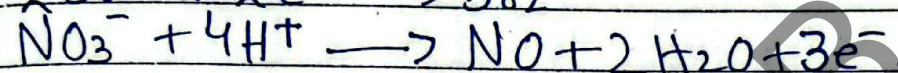
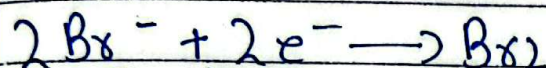
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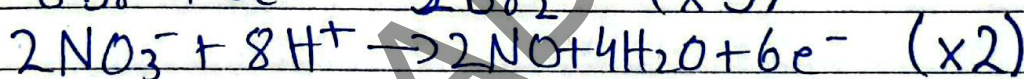
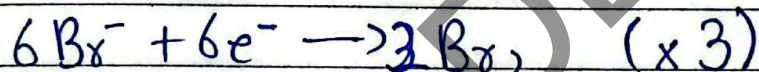
Step 3: Balance the H-atoms by adding H^+ ions.



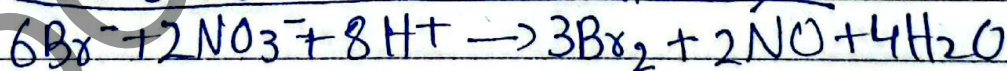
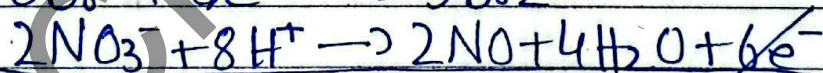
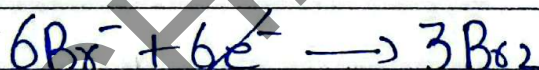
Step 4: Balance the $(e^- \text{ by})^x$ charges by adding e^- .



Step 5: loss of e^- - gain of e^- by multiplying with suitable digits.



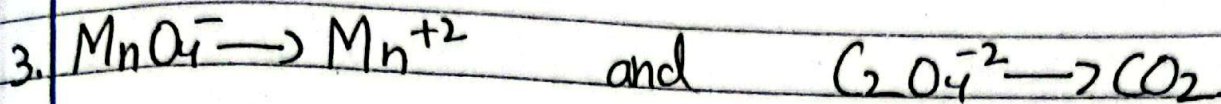
Step 6: Adding both equations.



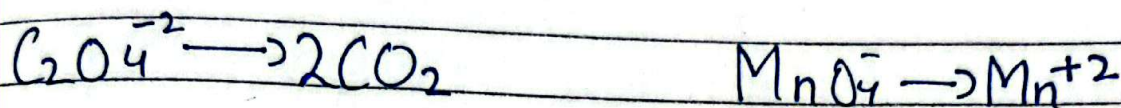
Step 7: Check if both equations are equal on both sides.

right side	=	left side.
$-6 + (-2) + 8$	=	0
0	=	0

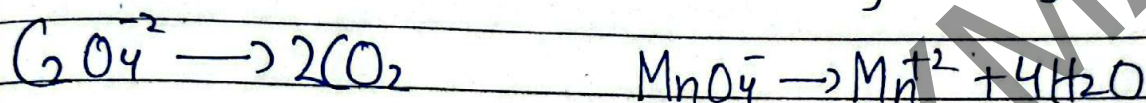
So, the equation is correct and balanced.



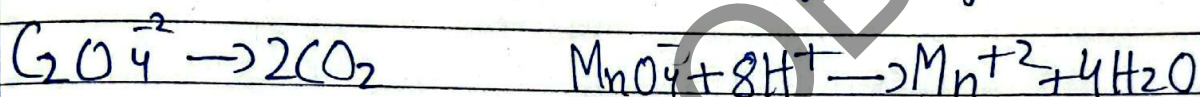
Step 1: Balance the half reactions except O and H atoms.



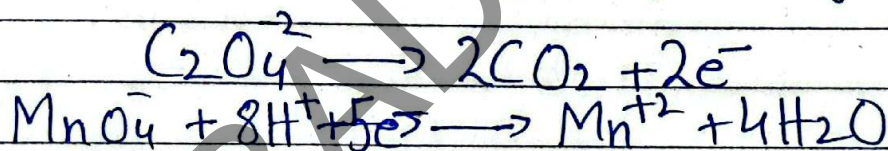
Step 2: Balance the O-atom by adding H_2O



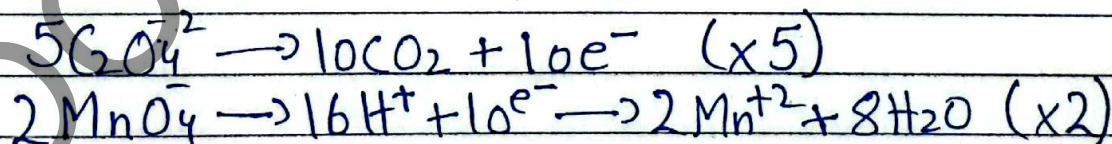
Step 3: Balance the H-atom by adding H^+ ions.



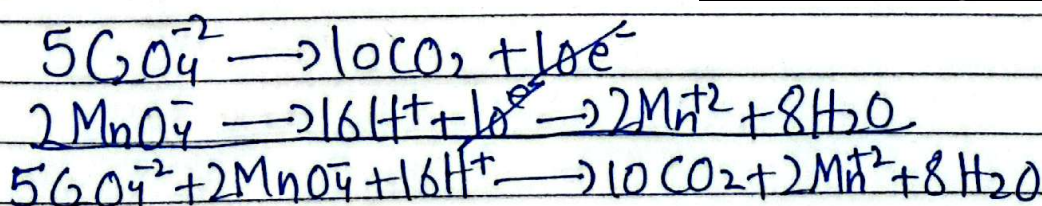
Step 4: Balance the charges by adding e^- .



Step 5: Bss of e^- - gain of e^- multiply with suitable digit.



Step 6: Adding the equations



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Step 7: Check if both sides of equation are balanced.

$$\begin{array}{ccc} \text{right side} & = & \text{left side} \\ -10 + (-2) + 16 & = & +4 \\ +4 & = & +4 \end{array}$$

So, the equation is balanced.

Oxidizing and Reducing Agent

Oxidizing agent: An oxidizing agent is a substance that is reduced in a chemical reaction.

OR.

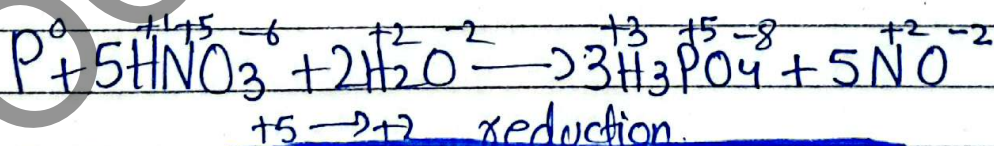
An oxidizing agent is a substance which oxidize other substance but itself is reduced.

Reducing agent: A reducing agent is a substance that is oxidized in a chemical reaction.

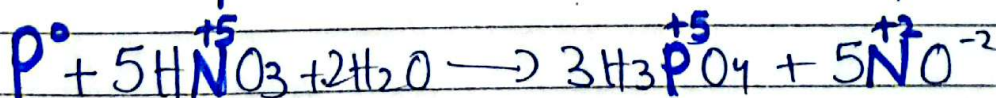
OR.

An Reducing agent is a substance which reduce other substances but itself is oxidized.

For Example:



$+5 \rightarrow +2$ reduction.



$0 \rightarrow +5$ oxidation

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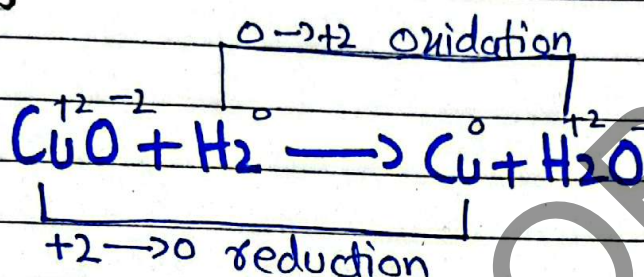
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- Here, Phosphorus is doing oxidation so it is a reducing agent.
- And, nitric acid is doing reduction so it is an oxidizing agent.

Concept Assessment # 2.5

Identify the oxidizing and reducing agent in the following reaction.



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oxidizing agent: Copper oxide is oxidizing agent because its O.S is decreased from +2 to 0 so, it is reduced, hence the compound that is reduced is oxidizing agent.

reducing agent: H_2 is reducing agent because its O.S is increased from 0 to +2 so, it is oxidized, hence the compound that is oxidized is reducing agent.

GALVANIC CELL (Daniel Cell):

The cell which involves spontaneous redox reaction to generate electricity is called galvanic or daniel cell.

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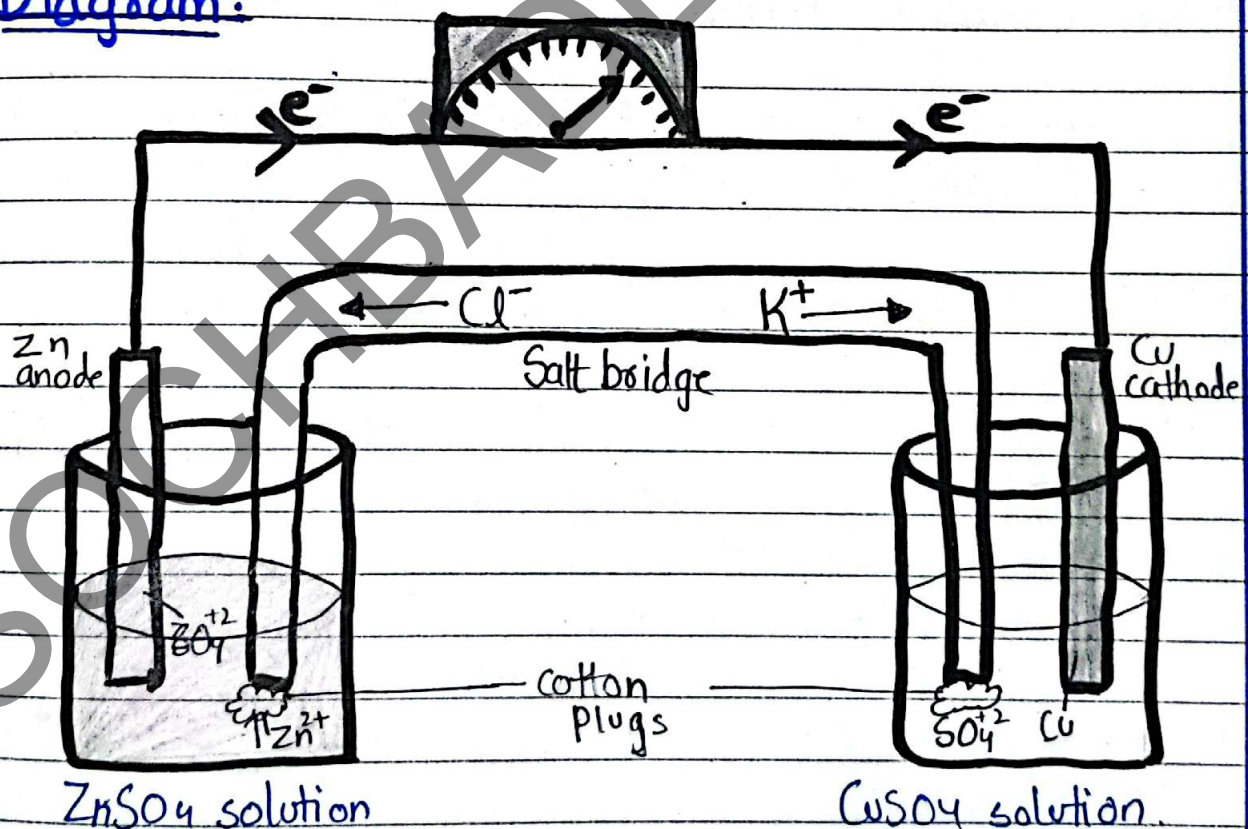
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Construction:

- Zn rod act as anode (oxidation)
- Cu rod act as cathode (reduction)
- Zn rod dipped 1M ZnSO_4 solution.
- Cu rod dipped 1M CuSO_4 solution
- A voltmeter to measure current.
- A salt bridge (inverted U tube) containing inert electrolyte such as KCl.

Diagram:



Function of Salt Bridge:

- Salt bridge is U-shaped tube filled with electrolyte/salt in gel form.
- It allows the movement of ions from one solution to another.
- Prevents the mixing of two solutions.
- Maintains 'electrical neutrality' within internal circuit by allowing ions between two half cell.
- Prevents the build-up of charge in either half cell, which would stop the reaction.

Electrodes and Half Cell:

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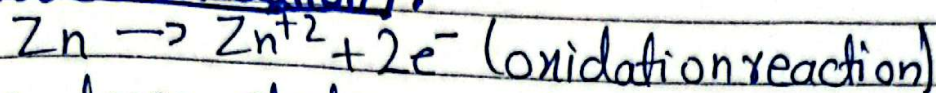
- Oxidation half cell (Anode): where electrons are lost. It is also called oxidation half reaction.
- Reduction half cell (Cathode): where electrons are gained. It is also called reduction half reaction.
- Zn is the anode and Cu is the cathode.

Electron flow:

- Zn has more tendency to lose electrons than Cu, so Zn undergoes oxidation and Cu reduction.
- The Zn electrode acquires negative charge relative to Cu electrode.
- The electrons flow from Zn electrode to Cu electrode.

Anodic and Cathodic Reactions:

• Anode (Oxidation):



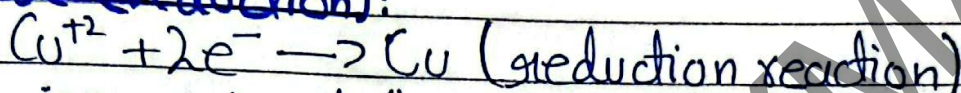
- Zinc loses electrons, releasing Zn^{+2} ions into the solution.

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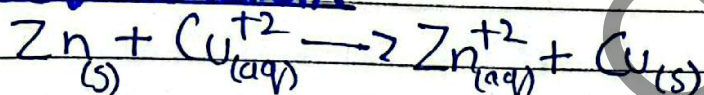
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• Cathode (Reduction):



- Copper ions gain electrons, depositing Cu on the electrode.

• Net Reaction:



CELL POTENTIAL:

Electromotive force: The force with which e^- are pushed to flow through the external wire from anode to cathode is called cell potential or electromotive force.

Unit: volts (V)

Cell Potential: The emf produced by a galvanic voltaic cell.

Denoted by: E°_{cell}

Dependency: E°_{cell} depends upon the difference in electrode potential of two half cells joined in series.

Anode Cathode in terms of potential:

Anode: Electrode with negative reduction potential (oxidation).

Cathode: Electrode with positive reduction potential (reduction).

Standard Conditions: E°_{cell} = Standard Cell Potential.

RTP conditions: Pressure: 1 atm

Temperature: 25°C

Conc: $1\text{M}/\text{mol dm}^{-3}$

Formula for Standard Cell Potential:

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

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Spontaneous / Non-spontaneous reaction

if E°_{cell} = positive $\vee$ then reaction is spontaneous

if E°_{cell} = negative then reaction is non-spontaneous.

- Single half cell potential cannot be measured directly. relative half cell potential can be determined by coupling with (S.H.E) (0.00V)

- **Standard electrode potential E°** is tendency of a half cell to undergo reduction relative to Standard Hydrogen Electrode (SHE)

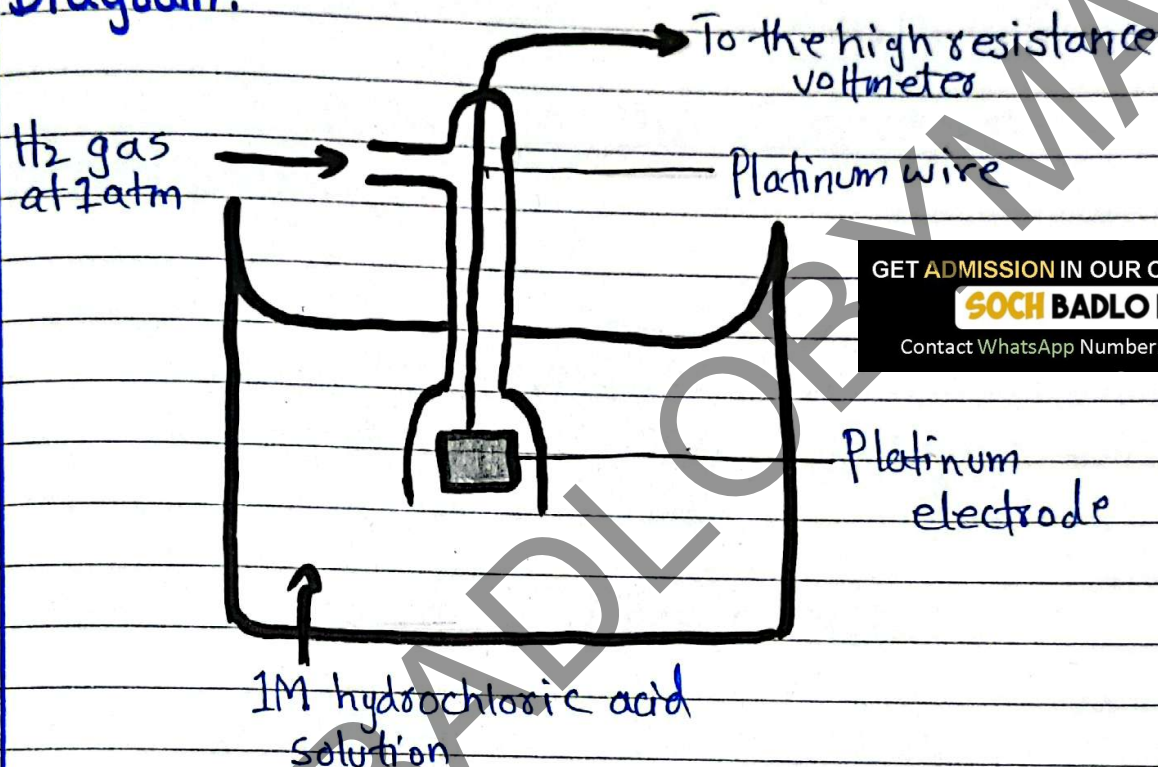
Standard Hydrogen Electrode (SHE):

- Standard Hydrogen Electrode is the primary reference electrode used to measure the standard electrode potential of other half cells.
- Standard Electrode Potential of S.H.E is zero volts at all temperatures.
- potential of single electrode cannot be measured therefore SHE is used.

Construction:

- SHE consists of a platinum foil coated with finely divided platinum.
- It is surrounded by hydrogen gas at 1 atm pressure in contact with 1M HCl solution at 298K.

Diagram:



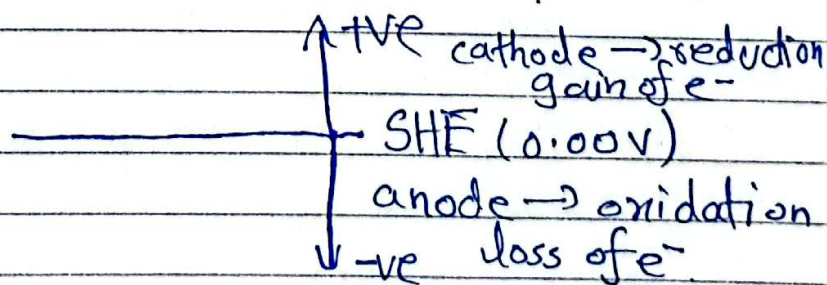
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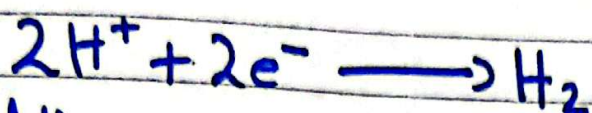
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* When SHE is combined with element whose reduction potential is greater than SHE then element be reduced.

* Element whose reduction potential is less than SHE then Element be oxidized.



Reduction: (act as cathode)



$$E_{\text{H}^+/\text{H}_2}^\circ = 0.00\text{V}$$

Oxidation
(Reduction)*: (act as anode)



$$E_{\text{H}_2/\text{H}^+}^\circ = 0.00\text{V}$$

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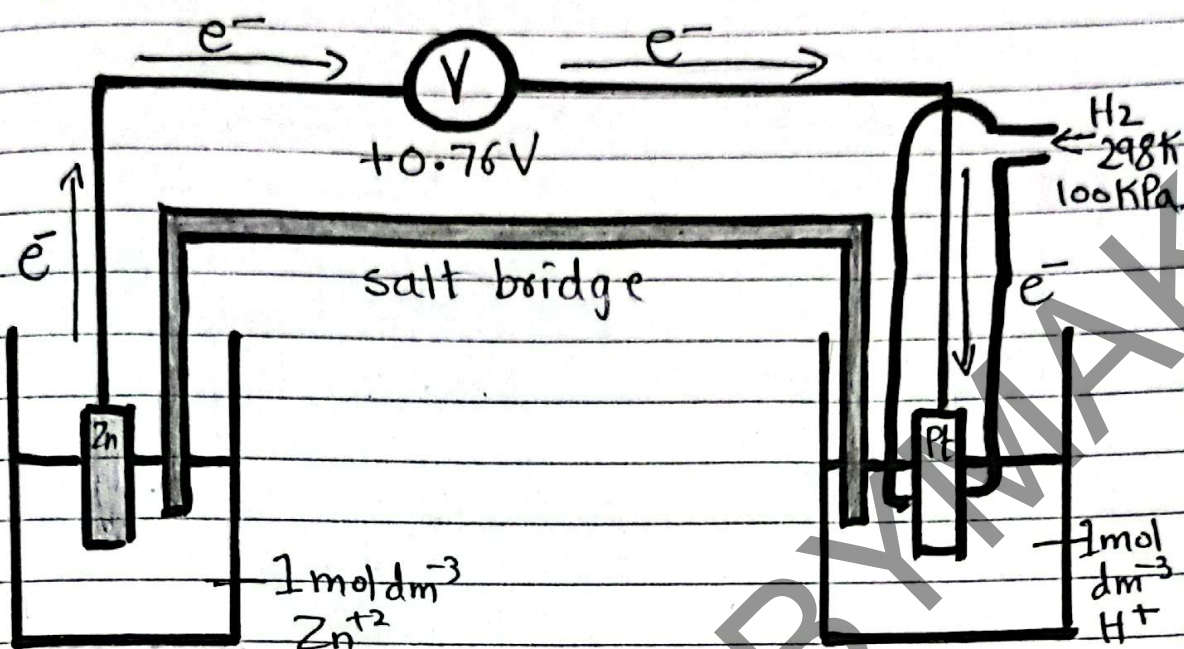
METHOD to MEASURE the ELECTRODE POTENTIAL:

- SHE is primary reference electrode used to measure Standard electrode potential of other half cells such as copper and zinc.

① Determination of Zn^{2+}/Zn Electrode:

- The electrochemical cell is set up by dipping Zn electrode in 1M ZnSO_4 solution and S.H.E in 1M HCl Solution.
- The two solutions are connected by using a salt bridge.
- Two electrodes are connected to a voltmeter through external wire, the deflection of voltmeter indicates the current flowing.
- Zn electrode act as anode and SHE act as cathode because when circuit is complete, electron begin to flow from Zn electrode to hydrogen electrode.
- The reading given by voltmeter is:
 $E_{\text{cell}}^\circ = 0.76\text{V}$.

Diagram:



Reactions:

- At anode: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$
- At cathode: $2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$
- Overall Reaction: $\text{Zn} + 2\text{H}^+ \rightarrow \text{Zn}^{2+} + \text{H}_2$

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Calculations:

- From the process it is clear that Zn loses electrons and undergoes oxidation, so we find (E°_{anode} or $E^\circ_{\text{Zn}^{2+}/\text{Zn}}$)
- H^+ 's oxidation potential will be (negative)^x positive and reduction potential will be negative.

So,

$$E^\circ_{\text{cell}} = +0.76\text{V}$$

As we know,

$$E^\circ_{\text{cell}} = E^\circ_{\text{anode}} + E^\circ_{\text{cathode}}$$

$$0.76 = E^\circ_{\text{anode}} + 0$$

$$0.76 = E^\circ_{\text{anode}}$$

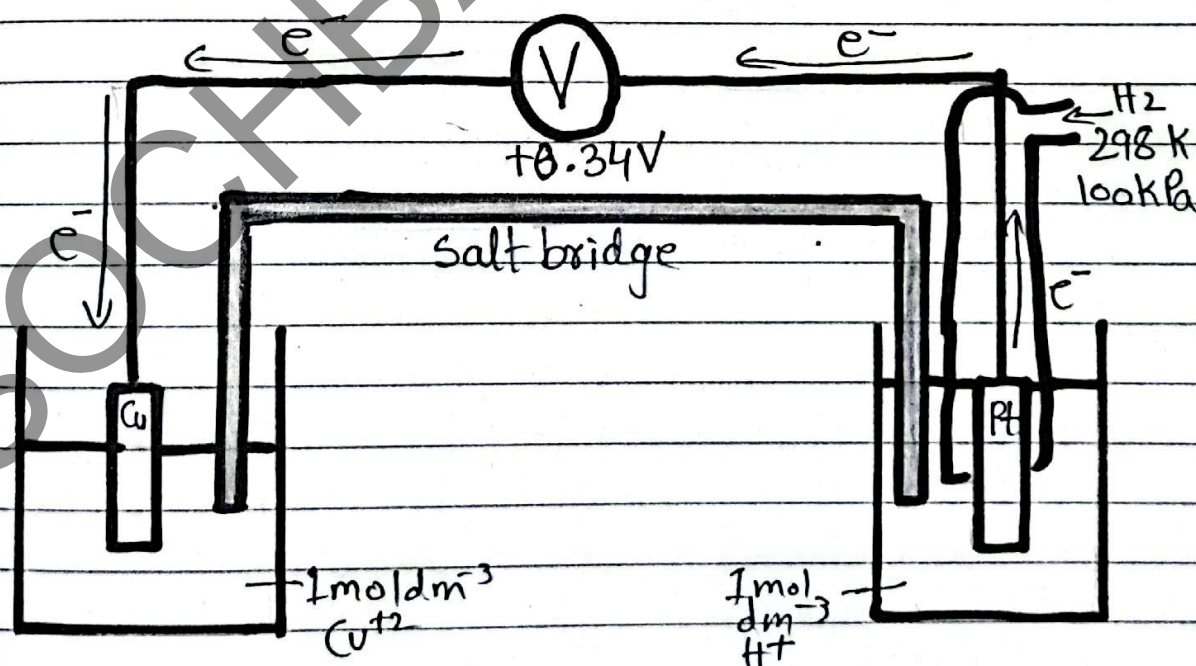
$$E^\circ_{\text{anode}} = 0.76\text{V}$$

So, at oxidation potential = $+0.76V$
at reduction potential = $-0.76V$.

② Determination of Cu^{+2}/Cu Electrode.

- An electrochemical cell is constructed by dipping Cu rod in $1M CuSO_4$ solution, and S.H.E dipped in $1M HCl$ solution.
- These solutions are connected by a salt bridge.
- Two electrodes are connected to a voltmeter which measures the current.
- When the circuit is complete, the electrons begin to flow from the hydrogen rod to Cu rod, So hydrogen electrode acts as anode and Cu electrode as cathode.
- The reading at E°_{cell} by the voltmeter is $E^\circ_{cell} = +0.34V$.

Diagram:



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Reactions:

- At anode: $\text{H}_2 \rightarrow 2\text{H}^+ + 2\text{e}^-$ (Oxidation)
- At cathode: $\text{Cu}^{+2} + 2\text{e}^- \rightarrow \text{Cu}$ (Reduction)
- Overall reaction: $\text{Cu}^{+2} + \text{H}_2 \rightarrow 2\text{H}^+ + \text{Cu}$

Calculation:

- from the process it is clear that, Cu gain electrons and is undergoing reduction so we find (E°_{cathode} or $E^\circ_{\text{Cu}^{+2}/\text{Cu}}$).
- so, its reduction potential will be positive and oxidation potential negative.

As,

$$E^\circ_{\text{cell}} = +0.34\text{V}$$

as we know,

$$E^\circ_{\text{cell}} = E^\circ_{\text{anode}} + E^\circ_{\text{cathode}}$$

$$0.34 = 0.00 + E^\circ_{\text{cathode}}$$

$$E^\circ_{\text{cathode}} = +0.34\text{V}$$

$$E^\circ_{\text{Cu}^{+2}/\text{Cu}} = +0.34\text{V}$$

Reduction potential = +0.34V

Oxidation potential = -0.34V

DETERMINATION OF CELL POTENTIAL:

- A cell reaction consist of two half cells.
- Reduction take place in the half cell having greater value of redox potential (cathode).
- Oxidation take place in the half cell having smaller value of redox potential (anode).
- equation having smaller value of Redox potential

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is reversed and added to equation of half cell having greater value of E° .

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

Steps for finding cell potential:

1. Analyze the value of reduction potential of equations and identify oxidation and reduction reactions.
2. Reverse the equation having small reduction potential.
3. add the reversed reaction and the second reaction
4. calculate E° of cell by using formula.
5. specify the direction of electrons.

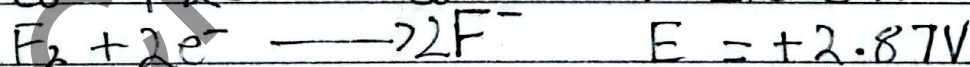
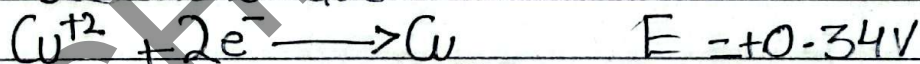
Concept Assessment #2.6

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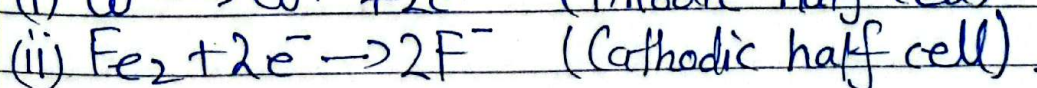
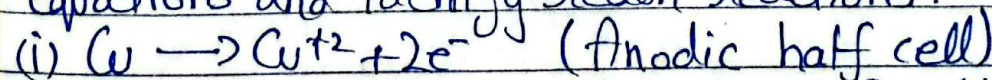
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The Standard reduction potential for the following half reactions are:

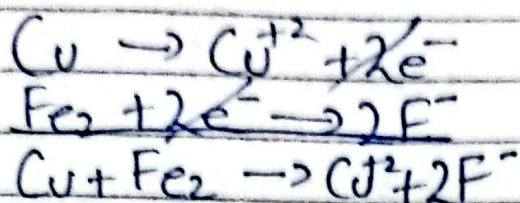


Estimate E°_{cell} for Cu-F₂ cell, write cell reactions, choose cathode and show the direction of electron flow.

Step 1: Analyze the value of red potential of equations and identify redox reactions.



Step 2: Add both equations.



Step 3: Find cell potential.

$$\begin{aligned}\mathcal{E}^{\circ}_{\text{cell}} &= \mathcal{E}^{\circ}_{\text{Cathode}} - \mathcal{E}^{\circ}_{\text{anode}} \\ &= +2.87\text{V} - (+0.34\text{V}) \\ \mathcal{E}^{\circ}_{\text{cell}} &= +2.53\text{V}\end{aligned}$$

So, the reaction is feasible and spontaneous.

Step 4: Specify direction of e^{-} .

• Direction is from anode to cathode.
Cu to Fe.

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FEASIBILITY OF CHEMICAL REACTIONS:

- A reaction is feasible (spontaneous) if it can occur on its own without needing outside energy.
- In electrochemistry we check feasibility using cell potential ($\mathcal{E}^{\circ}_{\text{cell}}$).

Rule:

$\mathcal{E}^{\circ}_{\text{cell}} > 0 \rightarrow +\text{ve}$ reaction is feasible.

$\mathcal{E}^{\circ}_{\text{cell}} < 0 \rightarrow -\text{ve}$ reaction is not feasible

Concept Assessment #2.7

i) Using emf data, argue on the following:

i) Can Mg displace Cu from a solution of copper (II) sulphate?

Ans) Yes, Mg is more reactive than Cu because it has a lower standard reduction potential ($E^\circ \text{Mg} = -2.37\text{V}$) than Cu ($E^\circ \text{Cu}^{2+}/\text{Cu} = +0.34\text{V}$). So, Mg can displace Cu from its salt solution.

ii) Can iodine displace chlorine from potassium chloride solution?

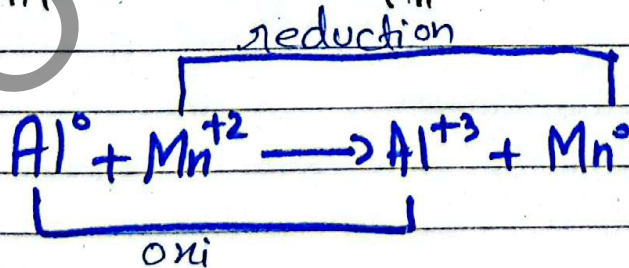
Ans) No, Iodine is less reactive than chlorine. Cl has a higher reduction potential ($E^\circ \text{Cl}_2 = +1.36\text{V}$) than iodine ($E^\circ \text{I}_2 = +0.53\text{V}$), So iodine cannot displace Cl from its salt solution.

2) Is the following reaction feasible?



The standard reduction potential values are $E^\circ_{\text{Al}} = -1.66\text{V}$, $E^\circ_{\text{Mn}} = -1.18\text{V}$.

Sol)



$$E^\circ_{\text{Al}} = -1.66\text{V} \text{ (anode)}$$

$$E^\circ_{\text{Mn}} = -1.18\text{V} \text{ (cathode)}$$

According to formula:

$$E^{\circ} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$

$$E^{\circ} = -1.18 - (-1.66V)$$

$$E^{\circ} = -1.18 + 1.66$$

$$E^{\circ} = +0.48$$

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So the reaction is feasible.

NERNST EQUATION: (formative)

"Nernst equation is used to calculate the electrode potential of an electrochemical cell under non-standard condition (when concentrations or pressure are not 1M or 1atm)"

Uses:

- It predicts cell voltage in real condition.
- Understand whether redox reaction will proceed spontaneously under given condition.
- relate concentration to voltage.
- calculate equilibrium condition.

Mathematically:

$$E = E^{\circ} + \frac{RT}{nF} \ln \frac{[\text{oxidized specie}]}{[\text{reduced specie}]}$$

- E is electrode potential
- E° is standard electrode
- no. of e^- transferred = n
- R = universal gas constant = 8.314 J mol^{-1}

- $T = K$ (Temperature)
- $F = 1 \text{ mol of } e^- \text{ charge} = 9.648 \times 10^4 \text{ C mol}^{-1}$
- [oxidized specie] = concentration of specie with higher oxidation state.
- [reduced specie] = conc of specie with lower oxidation state.
- $\ln = \text{natural log} = 2.303 \log$.

$$E = E^\circ + \frac{2.303RT}{nF}$$

(1) F

$$\begin{aligned} &= \frac{2.303 \times 8.314 \times 298K}{9.648 \times 10^4 \text{ C mol}^{-1}} \\ &= 0.059 \text{ J C}^{-1} \\ &= 0.059 \text{ V} \end{aligned}$$

As,

$$1 \text{ J/C} = 1 \text{ V}$$

$$E = E^\circ + \frac{0.059}{n} \log \frac{[\text{oxidized specie}]}{[\text{reduced specie}]}$$

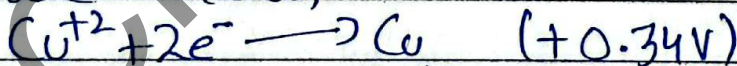
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Electrode potentials and Concentration:

As, we know,



- For this value Cu^{+2} ions should have a concentration of 1 mol/dm^3

Q. What will happen if conc of Cu^{+2} ion decreased by adding water?

Ans). If you decrease the conc of any species in an equilibrium reaction, then according to 'Le Chatelier's Principle' the position of equilibrium shifts in such a way as to encounter a counter

or nullify the change.

- The position of equilibrium will shift to left increasing the conc of Cu^{2+} ions again.
- Since E° measures the position of equilibrium, the electrode potential would become less positive. So, it will no longer be a standard potential.

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Condition for conc of ion :

- 1) An increase in conc of ion (Cu^{2+} ion) make reduction easier because more ions are available to gain electrons (increases electrode potential).
- 2) A decrease in conc of ion (Cu^{2+} ion) make reduction harder because less ions are available to accept electrons (decreases - electrode potential).

Ions conc $\propto$ Electrode potential.

Concept Assessment #2.8

The standard electrode potential of $\text{Fe}^{3+}/\text{Fe}^{2+}$ is 0.77V. What is the electrode potential of the system containing 1 mol/dm^3 of Fe^{3+} and 0.2 mol/dm^3 of Fe^{2+} ions?

Sol) According to formula:

$$E = E^\circ - \frac{0.059}{n} \log \frac{[\text{oxidized specie}]}{[\text{reduced specie}]}$$

$$E = 0.77 - 0.059 \log \frac{0.2}{1}$$

$$E = 0.77 - 0.059(-0.69)$$

$$E = 0.81071 \text{ V}$$

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ELECTROCHEMICAL SERIES

"A list of arrangement of elements in the order of their standard electrode potential with reference to standard hydrogen electrode is called an electrochemical series".

Activity series: "series in which elements are arranged in increasing order of reduction potential is called activity series".

Increase in Reactivity	Li	Very Active Metals	Co	Moderately active metals
	K		Ni	
	Ba		Sn	
	Sr		Pb	
	Ca			
	Na		H ₂	
	Mg	Intermediate active	Cu	Moderately noble metal
	Al		Ag	
	Mn		Hg	
	Zn			
	Cs		Pt	very noble metal
	Fe		Au	
	Cd			

Applications:

① feasibility of reaction.

if $E^{\circ} = +ve$ reaction is feasible

if $E^{\circ} = -ve$ reaction is non-feasible

② Calculation of voltage of cell.

For example, Cu and Zn.

$$\begin{aligned} E^{\circ}_{\text{cell}} &= E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} \\ &= 0.34 - (-0.76) \\ &= 1.1V. \end{aligned}$$

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③ Comparision of metal act as anode and cathode.

- if metal is more active then it have small reduction potential and is present at anode.
- if metal is less active then it have greater reduction potential and is present at cathode.

④ reactivity:

- Element having low/less E° reduction is more active / reactive, For Example Li and F, Li is more active than F as it has low E° reduction
- Element having more/greater E° reduction is less active / reactive, For Example, Zn and Cu, Cu have more reduction potential so it is less reactivier then Zn.

⑤ Displacement Reaction: Element which is more reactive will easily loose e^{-} and displace other atom in a reaction.

For Example:



As, zinc have less E° reduction it is more active so, it can displace Cu in CuSO_4 solution

Concept Assessment #2.9

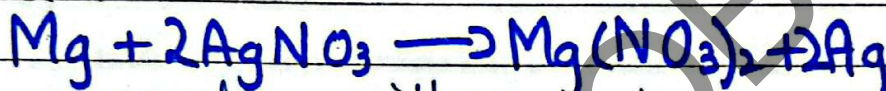
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Predict whether a replacement reaction will occur in the following instances.

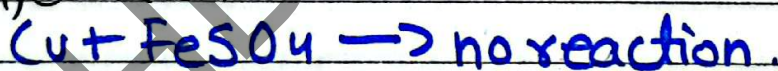
- a) Magnesium ribbon is in a solution of silver nitrate.



Yes, a reaction will occur, because Mg is more reactive than Ag, so it displaces Ag from AgNO_3 .

$\text{Mg} > \text{Ag}$ in reactivity.

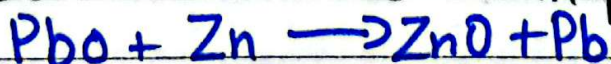
- b) A copper plate is dipped in iron(II) sulphate (FeSO_4) solution.



No, reaction will occur, because Cu is less reactive than Fe so it cannot displace Fe from FeSO_4 .

$\text{Cu} < \text{Fe}$ in reactivity.

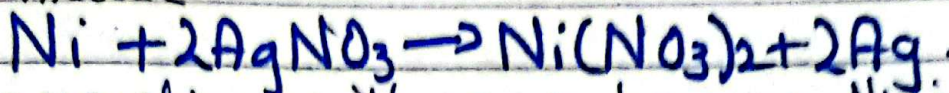
- c) lead(II) oxide is heated with powdered zinc.



Yes, a reaction will occur, because Zn is more reactive than Pb, so it reduce PbO to Pb metal.

$\text{Zn} > \text{Pb}$ in reactivity.

d) Nickel wire is placed into a solution of silver nitrate.



Yes, a reaction will occur because Ni is more reactive than Ag, so it displaces Ag from AgNO_3 solution.

$\text{Ni} > \text{Ag}$ in reactivity.

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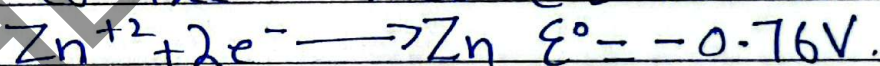
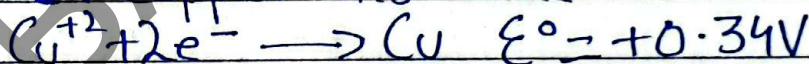
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THE RELATIVE REACTIVITY OF SPECIES AS OXIDIZING AGENTS OR REDUCING AGENTS

Oxidizing Agents: Those elements with high reduction potential are strong oxidizing agents.

Reducing Agents: Those elements with less reduction potential are strong reducing agents.

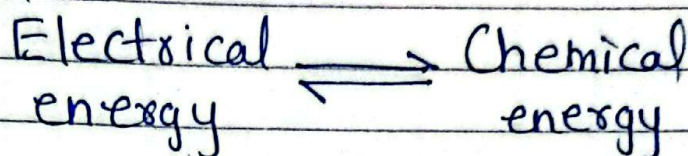
For Example: Copper and Zinc



- As, Cu have more reduction potential, so it is a strong oxidizing agent.
- As, Zn have less reduction potential, so it is a strong reducing agent.

TYPES OF ELECTROCHEMICAL CELLS.

Electrochemical Cell: Devices that convert electrical energy into chemical energy and vice versa



★ It has two types of cells:

- 1) Electrolytic cell (Electric $\rightarrow$ Chemical)
- 2) Galvanic cell. (Chemical $\rightarrow$ Electric)

ELECTROLYTIC CELL: "Device that convert electrical energy into chemical energy by electrolysis."

→ At anode (+) oxidation occurs and anions (-) passes through it. (loss of e^-)

→ At cathode (-) reduction occurs and cations (+) passes through it. (gain of e^-)

Component of electrolytic cell:

- ① Battery
- ② container
- ③ 2 Electrodes metal
- ④ Electrolyte. (that ionizes in water to high extent).

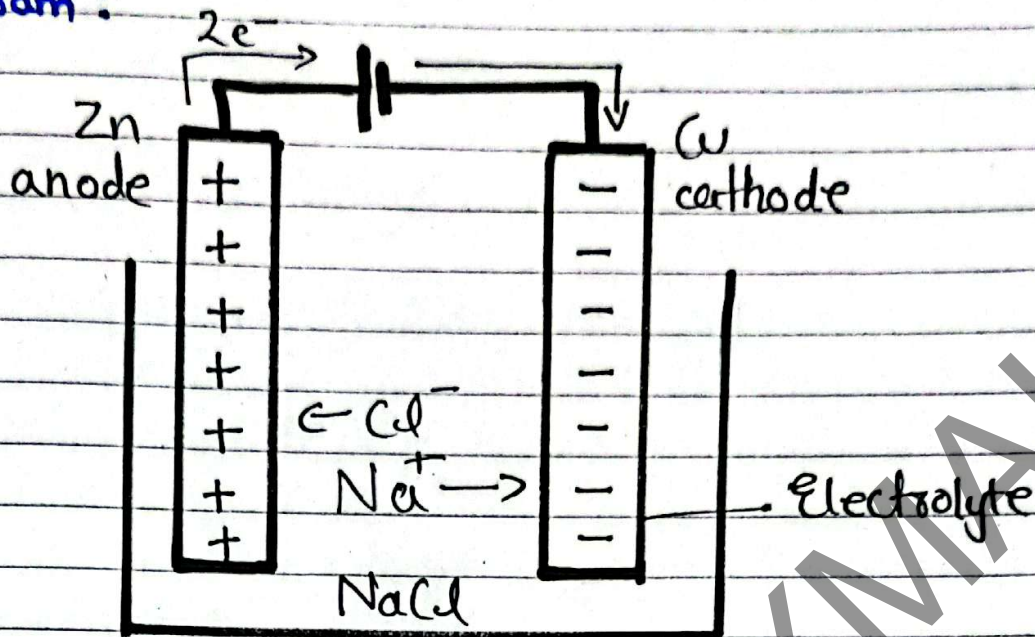
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- This redox process convert electrical energy into chemical energy by stimulation non-spontaneous reaction.

Diagram:



Reactions:

- Anode: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$
- Cathode: $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$

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Electricity battery: NaCl will be turned into Na^+ and Cl^- ions and will take position at anode and cathode.

Direction of flow of e^- :

Anode $\rightarrow$ cathode
externally wire.

ELECTROLYSIS: "Process in which chemical reaction occurs by means of electrical current in molten state or in aqueous state".

Process:

- ① Electrodes are connected with battery (+) anode - cathode.

② Electric current battery flows electrodes

Setup of Electrolysis:

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- 1) Use a container to hold the setup.
- 2) Insert two rods and connect them to a battery.
- 3) The rod connected to the + terminal is called the anode.
- 4) The rod connected to the - terminal is called the cathode.
- 5) Oxidation happens at anode (loss of e^-)
- 6) Reduction happens at cathode (gain of e^-).
- 7) An electrolyte is needed to let ions move.
Use an ionic salt like NaCl.
- 8) The reaction won't start on its own it needs the battery, so it's non-spontaneous.
- 9) A better electrolyte means more efficient conversion of electrical energy.

Unit:

- SI unit of charge is the coulomb (C)
- $1\text{ C} = 6.25 \times 10^{18} e^-$

Faraday (F): one mole of e^- and amounts is equal to 96487 C.

- SI unit of 'current' is 'ampere'.

• an **ampere** is equal to 'Cs'

• "**Ampere** is the amount of current flowing when one coulomb passes a specific point in one second."

FACTOR INFLUENCING PRODUCT OF ELECTROLYSIS.

- "Electrolysis involves the use of an electric current to drive a non-spontaneous redox reaction."
- The substance liberated during electrolysis depend on the following factors:
 - The state of electrolyte.
 - Position in the electrochemical or redox series
 - Concentration of electrolyte.

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1) The state of electrolyte:

Case I: Molten/fused electrolyte.

- Cations and Anions are discharged at the electrodes.

Case II: Aqueous Electrolytic Solution.

Brine: Conc Aq solution of NaCl.

- In this case electrolyte contain both ions from dissolve electrolyte and water.
- Water will produce H_2 at cathode.

2) Position in the electrochemical or redox series.

- Substances which are high in activity series are oxidized because they have less E°_{red} .
- Substances which are low in activity series are reduced because they have high E°_{red} .
- In this case cations are more reactive than hydrogen and do not discharge instead remain in solution, and water molecules are discharged at the cathode to liberate hydrogen.

3. Concentration of Electrolyte:

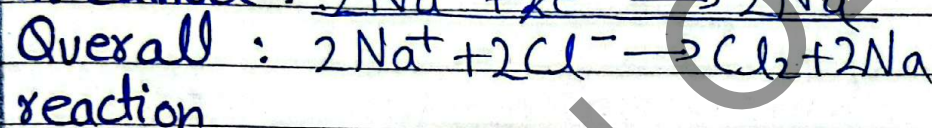
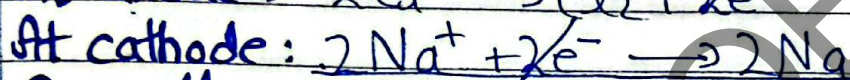
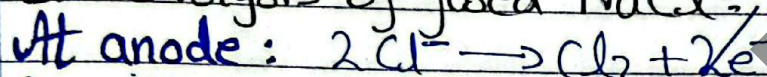
- The ions present and their concentration in the electrolyte affect which ions are discharged during electrolysis.
- Higher concentration ions are more likely to get discharged.
- Less reactive ions or those in higher conc get liberated first.

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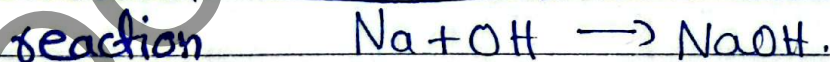
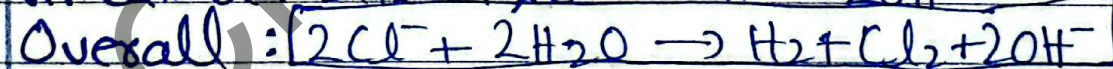
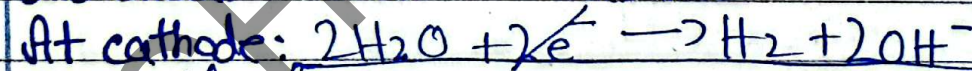
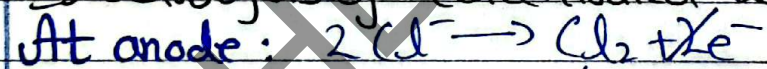
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1. Electrolysis of fused NaCl:



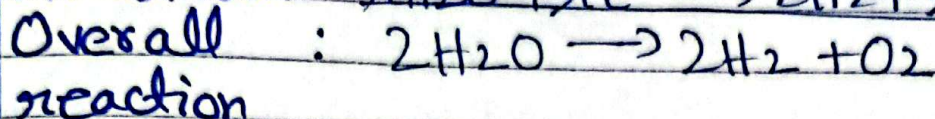
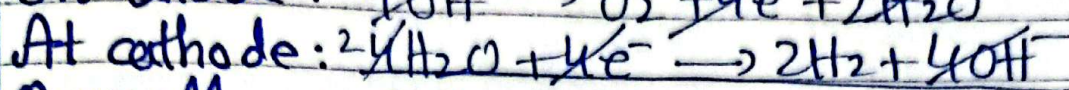
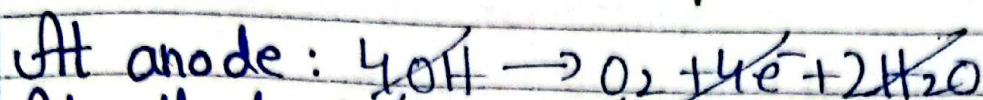
- Only Na^+ and Cl^- ions are present.
- No water, so Na is metal and Cl_2 gas is formed.

2. Electrolysis of concentrated aqueous NaCl (brine):



- Due to high Cl^- concentration, Cl^- ions are discharged instead of OH^- .
- Na^+ stays in solution (doesn't get discharged).
- H_2 gas at cathode, Cl_2 gas at anode, alkaline solution remains.

3. Electrolysis of dilute aqueous NaCl.



- Cl⁻ concentration is low, so OH⁻ gets discharged instead.
- At cathode, water is reduced to H₂, Na⁺ stays in solution.
- Final result - H₂ at cathode, O₂ at anode, and solution becomes basic.

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Relation Between the Faraday Constant, Avogadro Constant, and the Charge on the Electron.

① Relationship between current and time:

- Amount of charge produced at Electrode during electrolysis depends on quantity of current in ampere (I) and time (t) for which a constant current (I)^x is passed through electrolyte.

Mathematically: $Q = I \cdot t$

Q = quantity of charge in C.

I = current in ampere

t = time of electrolysis in sec.

② Relation between Faraday Constant, Avogadro's and charge on electron.

1 Faraday = amount of electrical charge carried by 1 mol of e^- .

$$\text{charge on } e^- = 1.602 \times 10^{-19} \text{ C}$$

$$\text{charge on one mole of } e^- = 1.602 \times 10^{-19} \text{ C} \times 6.022 \times 10^{23} \text{ per mol.}$$

$$\text{charge on one mole of } e^- = 96485.332 \text{ C/mol}$$

Faraday is unit of electrical charge named by Michael Faraday.

• It represents the charge carried by one mole of electron.

$$96500 \text{ C/mol} = 1 \text{ Faraday.}$$

↳ Thus quantity of charge in electrolysis can be determined by no. of Faraday's charge which passes.

• For calculation its value is taken 96500 C/mol

Relationship between the Faraday constant (F), Avogadro constant (N_A) and the charge on e^- can be expressed by the equation:

$$F = N_A e.$$

F = the Faraday constant (charge per mole of e^-)

N_A = the Avogadro constant (number of entities per mole)
 $= 6.02 \times 10^{23}$

e^- = the elementary charge (charge on a single e^-)
 $= 1.602 \times 10^{-19} \text{ C}$

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Experimental Determination of Avogadro Constant By Electrolytic Method.

- Electrolysis can be used to measure Avogadro's Constant experimentally.
- A known electrolyte is electrolyzed, and the metal deposited at the cathode is measured.
- Current (in amperes) is passed for a known time to determine the deposited metal mass.
- In the experiment, silver nitrate (AgNO_3) is electrolyzed using silver (Ag) electrodes.

Experimental Data =

Current = $I = 0.1$ ampere

Time = $30 \text{ min} \times 60 = 1800 \text{ s}$.

Mass of Ag = 0.201 g

$Q = I \times t$

$Q = 0.1 \times 1800 \text{ s}$

$Q = 180 \text{ coulomb charges}$.

$0.201 \text{ g of Ag} = 180 \text{ coulomb}$

$1 \text{ mol Ag} = 107.86 \text{ g Ag} = x$

by cross multiplication

$$x = \frac{180 \times 107.86}{0.201} = \boxed{96598 \text{ Coulomb}}$$

Cathodic Reaction: $\text{Ag}^+ + 1\text{e}^- \rightarrow \text{Ag}$

$1 \text{ mol of } \text{e}^- = 1 \text{ mol Ag}$.

Charge on $1 \text{ mol } \text{e}^- = 96598 \text{ C}$

Charge of $\text{e}^- = 1.609 \times 10^{-19} \text{ C}$

no. of e^- in $1 \text{ mol} = \frac{96598}{1.609 \times 10^{-19}}$

$$= \boxed{6.0298 \times 10^{23}}$$

very close to accepted value $N_A = \boxed{6.022 \times 10^{23}}$

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Concept Assessment #2.10

- i) An electrolytic cell is connected to power source for two hours. If the current flowing through the cell is 0.5A during this time. Find the mass of the substance liberated during this time interval. (the molar mass of the substance is 107.9g/mol).

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Data:

$$\text{Time} = 2 \text{ hours} = 7200 \text{ seconds}$$

$$\text{Current} = I = 0.5\text{A}$$

$$\text{Molar Mass} = M = 107.9\text{g/mol}$$

$$n = 1 \text{ (Assuming } 1e^- \text{ involved for } \text{Ag}^+)$$

$$\text{Faraday's constant (F)} = 96500 \text{ C/mol}$$

Calculation:

first we have to calculate charge.

$$Q = It$$

$$Q = (0.5)(7200)$$

$$Q = 3600\text{C}$$

Now, by calculating moles of substance liberated.

$$\text{moles} = \frac{Q}{nF}$$

$$= \frac{3600}{(1)(96500)}$$

$$= 0.0373$$

$$\text{moles} = 0.0373$$

Now to calculate mass of substance

$$\text{mass} = \text{moles} \times M$$

$$= 0.0373 \times 107.9$$

$$\text{mass} = 4.023\text{g}$$

- 2) Calculate the charge in coulombs when 3 moles of electrons flow through a circuit.

Number of moles of electrons (n) = 3 mol

Faraday's constant (F) = 96500 C/mol

Charge calculation:

As we know

$$Q = nf$$
$$= (3)(96500)$$

$$Q = 289500 \text{ C}$$

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- 3) Calculate the mass of silver deposited at the cathode during electrolysis of AgNO_3 solution, if you use a current of 0.1 A for 20 minutes.

Data:

Current (I) = 0.1 A

time (t) = 20 min = $20 \times 60 = 1200$ seconds.

Molar mass of Ag = 107.87 g/mol.

Number of e^- transferred (n) = 1

Faraday's constant (F) = 96500 C

Calculation of charge:

$$Q = It$$

$$= (0.1)(1200)$$

$$Q = 120 \text{ C}$$

Calculation of moles:

$$\text{moles} = \frac{Q}{nF}$$

$$= \frac{120}{(1)(96500)}$$

$$= 0.001244 \text{ mol}$$

$$\boxed{\text{mole} = 0.001244 \text{ mol}}$$

Calculate mass of Ag

mass = moles $\times$ molar mass

$$= 0.001244 \times 107.87$$

$$\text{mass} = 0.134\text{g}$$

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EXERCISE

1. Multiple Choice Questions (MCQs)

- i. Which of the following elements is reduced in the reaction? $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$
a) Na ☒ b) Cl c) Both Na and Cl d) Neither Na nor Cl
- ii. In the reaction $\text{Fe} + \text{CuSO}_4 \rightarrow \text{FeSO}_4 + \text{Cu}$, what is the reducing agent?
☒ a) Fe b) Cu c) SO_4^{2-} d) CuSO_4
- iii. What is the oxidation number of sulphur in H_2SO_4 ?
a) -2 b) 0 c) +2 ☒ d) +6
- iv. Which of the following best describes the process of oxidation?
a) Gain of electrons ☒ b) Loss of electrons
c) Gain of protons d) Loss of protons
- v. In the electrolysis of water, which substance is liberated at the cathode?
☒ a) H_2 b) O_2 c) H_2O d) OH^-
- vi. Which of the following metals would not react with HCl to produce H_2 gas?
a) Zn b) Fe ☒ c) Cu d) Mg
- vii. The Winkler method measures the concentration of which substance to determine BOD?
☒ a) O_2 b) CO_2 c) N_2 d) H_2
- viii. In a redox reaction, the oxidizing agent:
a) Loses electrons ☒ b) Gains electrons
c) Loses protons d) Gains protons
- ix. The standard electrode potential of the hydrogen electrode is defined as:
☒ a) 0 V b) 1 V c) -1 V d) 0.5 V
- x. What is the relationship between the Faraday constant (F), Avogadro's number (N_A), and the charge of an electron (e)?
☒ a) $F = N_A \times e$ b) $F = e / N_A$ c) $F = N_A / e$ d) $F = e^2 \times N_A$

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SLO based Questions

① Find oxidizing state by algebraic sum.

1. CO_2

$$x + 2(0) = 0$$

$$x + 2 = 0$$

$$x = -2$$

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2. $\text{K}_2\text{Cr}_2\text{O}_7$

$$2(+1) + 2x + (-2)7 = 0$$

$$+2 + 2x - 14 = 0$$

$$-12 + 2x = 0$$

$$-12 = 2x$$

$$\frac{-12}{2} = x$$

$$x = +6$$

3. $(\text{SO}_4)^{2-}$

$$x + (-2)4 = -2$$

$$x - 8 = -2$$

$$x = -2 + 8$$

$$x = +6$$

② What are the rules for assigning O.S to an element?

① Oxidation State of any uncombined or any free atoms is equal to zero. For Example H_2 , N_2 , O_2 , etc.

② Oxidation State of simple ions is same as their charge on it. For example Na^+ , PO_4^{3-} , SO_4^{2-} etc.

③ In complex ion sum of oxidation state of atom is equal to the charge on their ion. For example in $CO_3^{2-} = -2$

$$x + (-2)(3) = -2$$

$$x = +4$$

④ Oxidation state of each atom in molecule is counted separately and the algebraic sum is zero. For example HCl

Elements

O.S.

IA

+1

IIA

+2

IIIA

+3

hydrogen

+1 (for metal hydride -1)

Oxygen

-2

per oxide

-1

super oxide

$(-1/2)$

VIA

-1

VIIA

0

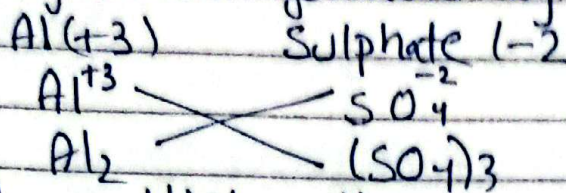
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⑤ For complex ions such as $K_2Cr_2O_7$, we represent the one element with variable (x) and multiply it with the given number of atoms.

② For an ionic compound such as aluminium sulphate we simply find its formula by following trick.



we simply multiply the O.S of Aluminium (ignoring sign) with sulphate to form $(\text{SO}_4)_3$ and O.S of sulphate with Aluminium to form Al_2 hence it becomes $\text{Al}_2(\text{SO}_4)_3$.

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③ What happen when salt bridge is removed in half cells?

Ans) If salt bridge is removed in half cells then the circuit will break, electrons will stop flowing, and the reaction will stop because the salt bridge keeps ions balanced in both half cells.

④ Why H_2 gas is added in S.H.E?

Ans) H_2 gas is added in SHE to provide hydrogen ions (H^+) and maintain a reversible redox reaction at the platinum electrode. Basically it's there to act as the standard for measuring electrode potentials, like the zero point for comparin all other electrodes.

5) Difference between electrolyte and non-electrolyte

Electrolyte	Non-electrolyte
Electrolyte conducts electricity because it forms ions in sol.	Non-electrolyte doesn't conduct electricity, because it doesn't have ions to carry current.
It has ionic bonds	It has covalent bonds
In presence of water, electrolyte dissociates completely/partially.	In presence of water, non-electrolyte is dissolved.
It has strong and weak electrolytes	They are mostly organic compounds.
They have ions in solution	They only have molecules.
Example: NaCl, HCl, KOH etc.	Example: Glucose, ethanol, urea, etc.

6) Difference between Galvanic and Electrolytic cell.

Galvanic cell	Electrolytic cell
Galvanic or Daniel cell converts chemical energy into electrical energy	Electrolytic cell converts electrical energy into chemical energy.
It is a spontaneous reaction.	It is a non-spontaneous reaction.

Galvanic cell is self-powered as it generate electricity.

Electrolytic cell needs external battery to get powered.

At anode it is negative, and oxidation occurs.

At anode it is positive, and oxidation occurs.

At cathode it is positive, and reduction occurs.

At cathode it is negative, and reduction occurs.

In galvanic cell electron flow from anode to cathode using external circuit.

In electrolytic cell electron flow from power source to cathode.

A salt bridge is required to maintain charge balance.

Salt bridge is not required.

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It is used in batteries, and voltaic cells.

It is used in electrolysis of water and electroplating.

Example: Daniell cell (Zn-Cu)

Example: electrolysis of NaCl or water.

SHORT ANSWER QUESTIONS

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(i)

The oxidation potential of Zn is $+0.76\text{ V}$ and its reduction potential is -0.76 V . explain with reasons.

Ans) • The oxidation potential of Zn is $+0.76\text{ V}$, meaning it loses electrons easily and undergoes oxidation.
• Its reduction potential is -0.76 V , meaning it does not easily gain electrons.

(ii)

A salt bridge maintains the electrical neutrality in the cell.

Ans) A salt bridge maintains the electrical neutrality in the electrochemical cell by allowing the flow of ions. It prevents charge build-up and completes the circuit.

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(iii)

Na and K can displace hydrogen from acids but Cu and Pt cannot.

Ans) Na and K are highly reactive metals that lie above (in) the hydrogen in the activity series, so they can displace hydrogen from acids. Cu and Pt lie below the hydrogen and cannot displace it.

(iv)

Define oxidation in terms of electron transfer.

Ans) Oxidation is defined as the loss of electron by an atom or ion, which results in an increase in its oxidation number.

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(v)

What is the oxidation number of oxygen in H_2O_2 ?

Ans) In hydrogen peroxide (H_2O_2), each oxygen has an oxidation number of -1 due to presence of an oxygen-oxygen single bond.

(vi)

Identify the reducing agent in the reaction.



Ans) In this reaction, zinc donates electrons and gets oxidized. Therefore, zinc is the reducing agent.

(vii)

State the purpose of Winkler method.

Ans) Winkler method is used to determine the amount of dissolved oxygen (DO) in water. It helps check water quality and how suitable it is for aquatic life.

(viii)

Explain what happens at the anode during the electrolysis of aqueous sodium chloride.

Ans). During electrolysis of aqueous NaCl, Chloride ions (Cl^-) lose electrons at the anode to form chlorine gas (Cl_2). This is an oxidation process. $2\text{Cl}^- \rightarrow \text{Cl}_2(\text{g})$

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(ix)

What does a positive standard electrode potential indicate about a substance's tendency to gain electrons?

Ans). A positive standard electrode potential indicates a strong tendency of the substance to gain electrons and undergo reduction.

(x)

Calculate the oxidation number of chromium in $\text{K}_2\text{Cr}_2\text{O}_7$?

Ans). Let the oxidation number of Cr be x .

$$2(+1) + 2x + 7(-2) = 0$$

$$2 + 2x - 14 = 0$$

$$2x = 12$$

$$x = \frac{12}{2}$$

$$\boxed{x = +6}$$

So, the O.S of Cr is +6

(xi)

Describe the role of standard hydrogen electrode in electrochemistry.

Ans) The standard hydrogen electrode (SHE) is used as a reference electrode with zero potential. It helps in comparing the electrode potentials of other half-cells.

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(xii)

What is the significance of the activity series of metals?

Ans) The activity series ranks metal by their reactivity. It helps to predict whether a metal can displace another from its salt solution.

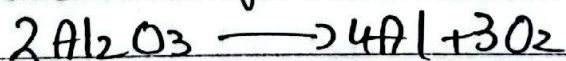
(xiii)

How can you deduce the feasibility of a redox reaction using electrode potentials?

Ans) Feasibility of a redox reaction is judged by calculating cell potential. If the E_{cell} is positive then reaction is spontaneous and feasible.

(xiv)

Bauxite ore is used for the commercial ...



sol) Use Faraday's law.

$$\bullet \text{ Total charge} = I \times t$$

$$= 15 \times 36000$$

$$\text{Total charge} = 540000 \text{ C}$$

3 moles of electrons produce 2 moles of Al.
according to Faraday's law:

$$\text{Moles} = \frac{Q}{(3)(96500)}$$

$$\text{Moles} = \frac{540000}{289500}$$

$$\text{Moles} = 1.865 \text{ mol}$$

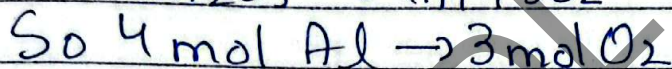
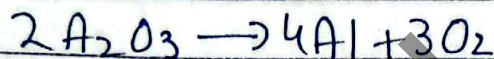
by using formula:

$$\text{Mass} = \text{moles} \times \text{molar mass} \quad \because \text{Al molar mass}$$

$$\text{Mass} = 1.865 \times 27 \quad \text{is } 27$$

$$\text{Mass} = 50.36 \text{ g}$$

Now we will find moles of O_2 gas to get volume
From the reaction:



then 1.865 mol Al will give:

$$\frac{3}{4} \times 1.865 = 1.39875 \text{ moles of O}_2$$

According to volume's formula:

$$\text{Volume} = \text{moles} \times \text{molar volume (24.5)}$$

$$= 1.39875 \times 24.5$$

$$= 34.27 \text{ dm}^3$$

$$\text{Volume of O}_2 \text{ at STP} = 34.27 \text{ liters}$$

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(xv)

Which compound will give more mass of metal, when 15 ampere current is passed through a molten mass of these salts for 1 hr.

- a) NaCl b) CaCl_2 .

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Firstly, find time and charge

$$t = 1 \text{ hour} = 3600 \text{ s.}$$

$$Q = It$$

$$= (15)(3600)$$

$$= 54000 \text{ C}$$

Now, compare both salts using equivalent mass.

For Na^+ :

$$n = 1 (1e^- \text{ per } \text{Na}^+)$$

$$\text{Molar mass} = 23 \text{ g mol}^{-1}$$

$$1 \text{ mol Na needs } 96500 \text{ C}$$

$$\text{moles of Na} = \frac{54000}{96500}$$

$$\text{moles of Na} = 0.559 \text{ mol}$$

$$\text{Mass} = \text{moles} \times \text{Molar mass}$$

$$= 0.559 \times 23$$

$$\text{mass} = 12.86 \text{ g}$$

For Ca^{2+} :

$$n = 2 (2e^- \text{ per } \text{Ca}^{2+})$$

$$\text{Molar mass} = 40 \text{ g mol}^{-1}$$

$$1 \text{ mol Ca needs } 2 \times 96500$$

$$= 193000 \text{ C.}$$

$$\text{Moles of Ca} = \frac{54000}{193000}$$

$$= 0.2798 \text{ mol}$$

$$\text{moles of Ca} = 0.2798 \text{ mol}$$

$$\text{Mass} = 0.2798 \times 40$$

$$\text{mass} = 11.19 \text{ g}$$

According to calculations NaCl gives more mass (12.86 g) than CaCl_2 (11.19 g) for the same current and time, so, NaCl will give more mass of metal.

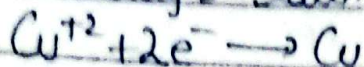
(xvi)

How many hours would electroplating have to be continued at the rate of 5 amperes if 75g of copper is to be deposited from CuSO_4 sol?

Sol) Molar mass of $\text{Cu} = 63.5 \text{ g mol}^{-1}$

To find moles of $\text{Cu} = \frac{\text{mass}}{\text{molar mass}} = \frac{75}{63.5} = 1.181 \text{ mol}$

Using Faraday's Law:



So 1 mol Cu need 2 Faradays $= 2 \times 96500$
 $= 193000 \text{ C}$

Total charge $Q = 1.181 \times 193000$
 $Q = 227370 \text{ C}$

Now to find time using $Q = It$.

$$t = \frac{Q}{I} = \frac{227370}{5}$$

$$t = 45474 \text{ s}$$

Convert to hours:

$$\frac{45474}{3600} = 12.63 \text{ hours}$$

So, electroplating should continue for about 12.63 hours with 5A current to deposit 75g of Cu .

(xvii)

Differentiate between:

Oxidation half cell
involves loss of electrons

Reduction half cell
involves gain of electrons.

It occurs at anode

It occurs at cathode.

Oxidation number increase

Oxidation number decreases

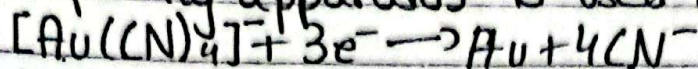
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(Xviii)

An electroplating apparatus is used to coat...



First we'll calculate time and charge.

$$t = 30 \text{ min} = 1800 \text{ seconds}$$

$$Q = It$$

$$= (5)(1800)$$

$$Q = 9000 \text{ C}$$

Moles of electrons:

$$1 \text{ Faraday} = 96500 \text{ C}$$

$$\text{Moles of } e^- = \frac{9000}{96500} = 0.09326 \text{ mol}$$

Moles of Au:

Gold requires $3e^-$ per atom

$$\text{moles of Au} = \frac{0.09326}{3}$$

$$\text{Au} = 0.0311 \text{ mol}$$

Moles of Gold:

$$\text{Molar mass of Au} = 197 \text{ g mol}^{-1}$$

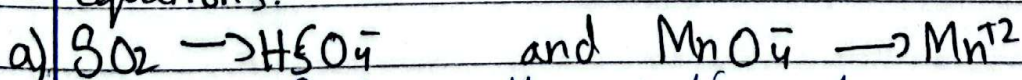
$$\text{Mass} = 0.0311 \times 197$$

$$\boxed{\text{Mass} = 6.12 \text{ g}}$$

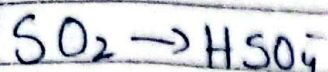
About 6.12 g of gold will be deposited in 30 minutes with 5A current.

(Xix)

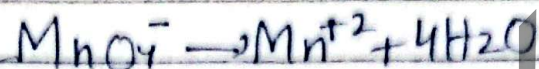
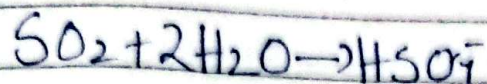
Construct redox equations using the following half equations.



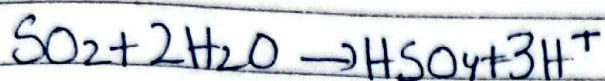
Step 1: Balance the half reactions except O and H atoms.



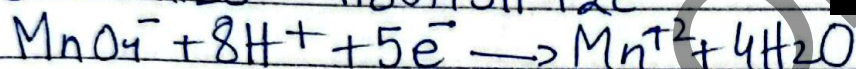
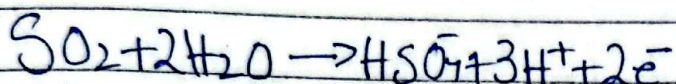
Step 2: Balance O atoms with H_2O



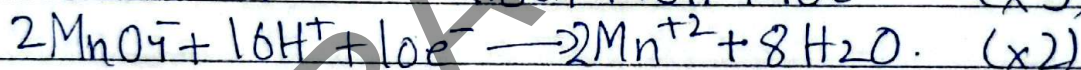
Step 3: Balance H atom by adding H^+ ions



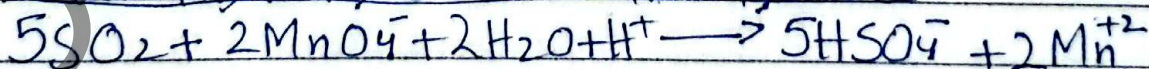
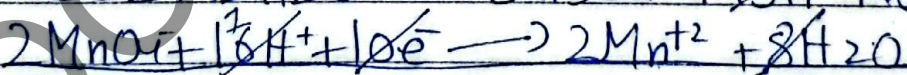
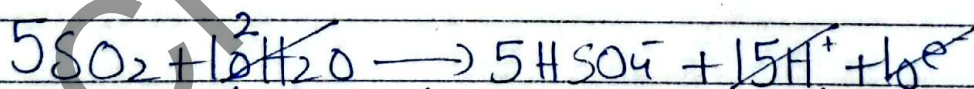
Step 4: Balance charge with e^-



Step 5: Equalize e^- by multiplication:



Step 6: Add both equations.

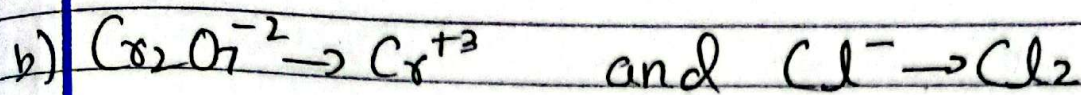


The above equation is balanced.

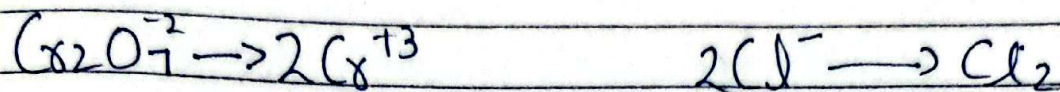
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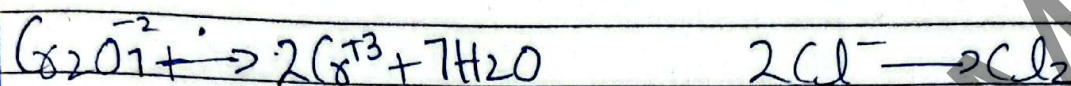
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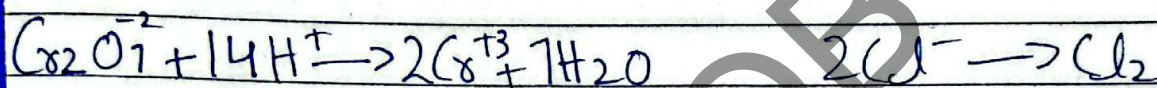
Step 1: Balance all atoms except O and H



Step 2: Balance O-atoms using H_2O



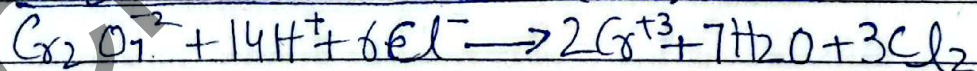
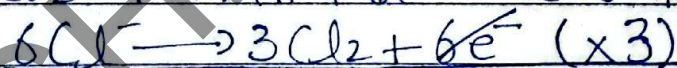
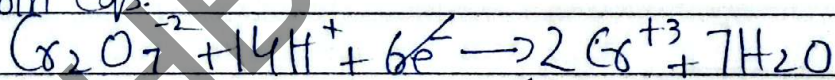
Step 3: Balance H-atoms by adding H^+ ions.



Step 4: Balance charges using e^-



Step 5: Equalize electrons by multiplication and add both eqs.

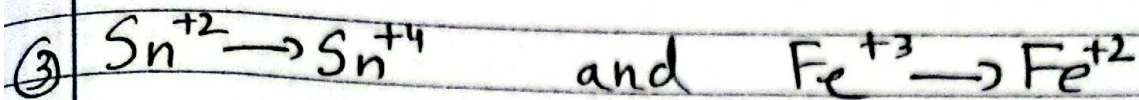


The above equation is balanced.

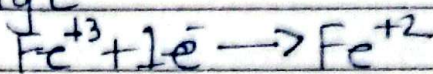
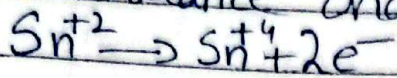
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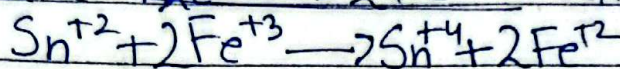
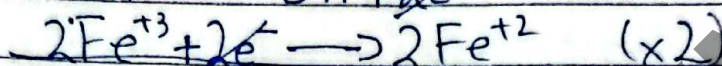
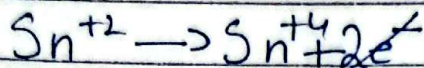
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Step 1: Balance charges using e^-



Step 2: equalize e^- by multiplication and add both equations.

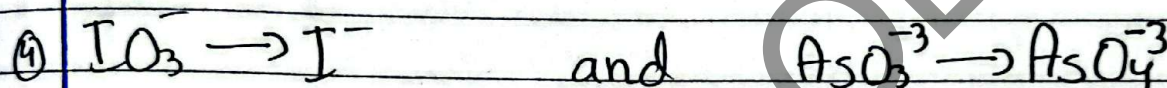


The above equation is balanced.

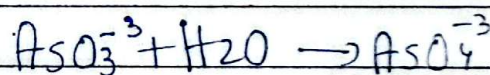
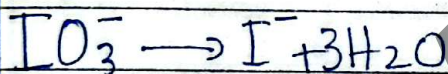
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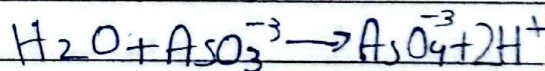
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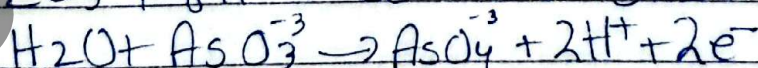
Step 1: balance O-atoms by adding H_2O



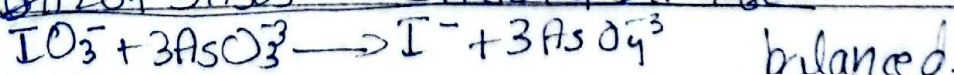
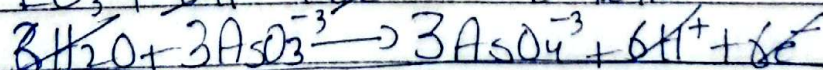
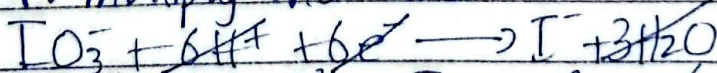
Step 2: balance H-atom by H^+ ions.

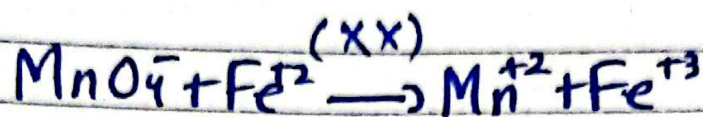


Step 3: Balance charges with e^-



Step 4: multiply and add





a) Identify the oxidation states of Mn and Fe before and after the reaction:

Mn in MnO_4^- changes from +7 to +2

Fe changes from +2 to +3

b) Explain which species is oxidized & reduced.

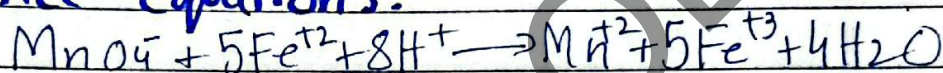
- Fe is oxidized species
- Mn is reduced species.

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c) Balance equations:



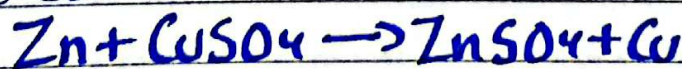
d) Oxidizing & Reducing Agents:

- MnO_4^- is oxidizing Agent
- Fe is reducing Agent.

(XXI)

Why Mg can displace Zn from ZnSO_4 solution but Cu cannot.

Ans). Mg is more reactive than Zn, so it can displace Zn from ZnSO_4 . Copper is less reactive and cannot. When Zn strip is placed in CuSO_4 solution, Zn displaces Cu.



By Myjiesha Ehsan Khan.