Designed by Tr. Joel PCM – Designer of Physics, Chemistry & Mathematics Item Banks A level

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Item 1

Redox Processes in Metal Extraction and Electrochemistry

A student in a school laboratory investigated redox reactions using copper(II) sulfate solution, zinc metal, and acidified potassium manganate(VII). In one experiment, zinc granules were added to a blue copper(II) sulfate solution and a reddish-brown deposit formed while the solution gradually lost its colour. In another experiment, iron(II) ions were titrated with acidified potassium manganate(VII) solution. The student later constructed a simple galvanic cell using zinc and copper electrodes connected through a salt bridge to study electron flow and electrode potentials. Observations showed that zinc was more reactive than copper and acted as a reducing agent in several reactions.

Tasks

A. Concepts of Redox

i. Define oxidation and reduction in terms of electron transfer.

ii. From the reaction between zinc and copper(II) sulfate, identify:

the oxidizing agent

the reducing agent

B. Half Equations and Overall Reaction

i. Write the half-equation for oxidation of zinc.

ii. Write the half-equation for reduction of copper(II) ions.

iii. Combine the half-equations to obtain the overall ionic equation.

C. Redox Titration

i. State the colour change observed when acidified potassium manganate(VII) reacts with iron(II) ions.

ii. Write the balanced ionic equation for the reaction between MnO_4^- and Fe^{2+} in acidic medium.

iii. Explain why potassium manganate(VII) is described as a self-indicator.

D. Electrochemical Cells

i. Draw a labelled diagram of the Zn/Cu galvanic cell.

ii. State the direction of electron flow in the external circuit.

iii. Explain why zinc acts as the anode.

iv. Predict what would happen to the cell voltage if copper were replaced with silver, giving a reason.

Solutions: Redox Processes in Metal Extraction and Electrochemistry

A. Concepts of Redox

i. Definitions

Oxidation: Loss of electrons.

Reduction: Gain of electrons.
(Also oxidation = increase in oxidation state, reduction = decrease.)
ii. Reaction: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$
Oxidizing agent: Copper(II) ions (Cu^{2+}) they gain electrons.
Reducing agent: Zinc (Zn) — it loses electrons.

B. Half Equations and Overall Reaction
i. Oxidation half-equation (zinc):
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
ii. Reduction half-equation (copper):
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
iii. Overall ionic equation
$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

C. Redox Titration (Fe^{2+} vs MnO_4^-)
i. Colour change
Purple potassium manganate(VII) solution becomes colourless/pale pink at end point.
ii. Balanced ionic equation
$MnO_4^- + 5Fe^{2+} + 8H^+ \rightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$
Explanation:
MnO_4^- is reduced (oxidizing agent)
Fe^{2+} is oxidized
iii. Why $KMnO_4$ is a self-indicator
Because it has a strong purple colour that disappears when reduced. The first permanent pale pink colour shows the end point — no external indicator is required.

D. Electrochemical Cell
i. Description of labelled Zn/Cu cell
Components:
Zinc electrode in Zn^{2+} solution
Copper electrode in Cu^{2+} solution
Salt bridge connecting solutions
External wire and voltmeter
ii. Direction of electron flow
Electrons flow from zinc → copper through the external circuit.
Reason: Zinc is more reactive and loses electrons.
iii. Why zinc is the anode
Anode = oxidation site.
Zinc loses electrons:
$Zn \rightarrow Zn^{2+} + 2e^-$
Therefore zinc is the anode.
iv. Replacing copper with silver
Prediction: Cell voltage increases.
Reason:

Silver has a higher reduction potential than copper, so the potential difference between zinc and the second electrode becomes larger.
Result → stronger cell.

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