

Redox Reactions — Assignment 2

Class 11 Chemistry · Higher-difficulty challenge set + integrated problems · Answer key included · Narayan Gurukul Academy

Instructions: (1) Attempt all questions; start with the topic-wise challenge set, then attempt the integrated (multi-topic) problems.
(2) Give all answers with correct SI units. $g = 10 \text{ m/s}^2$ unless stated.
(3) A final answer key is provided at the end — try the full set before checking it.

Part A — Topic-wise challenge questions

Oxidation and Reduction

1. Define oxidation and reduction in terms of electrons and identify the oxidising and reducing agents in $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$.
2. Explain the terms oxidant and reductant with one example each.
3. Show that oxidation and reduction always occur together.

Oxidation Number and Its Rules

1. State the rules for assigning oxidation numbers.
2. Calculate the oxidation number of Cr in $\text{K}_2\text{Cr}_2\text{O}_7$.
3. Calculate the oxidation number of S in H_2SO_4 .

Balancing Redox Equations — Oxidation-Number Method

1. Explain the steps of the oxidation-number method of balancing a redox equation.
2. Balance $\text{Fe}^{2+} + \text{MnO}_4^- + \text{H}^+ \rightarrow \text{Fe}^{3+} + \text{Mn}^{2+} + \text{H}_2\text{O}$ using the oxidation-number method.
3. State the principle used in the oxidation-number method of balancing.

Ion-Electron (Half-Reaction) Method

1. Explain the ion-electron method of balancing redox equations.
2. Balance the half-reaction $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$ in acidic medium.
3. How are oxygen and hydrogen atoms balanced in an acidic medium?

Types of Redox Reactions

1. Classify the following: (i) $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$, (ii) $2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$.
2. Explain disproportionation with one example.
3. Give one example each of combination, decomposition and displacement reactions.

Electrochemical Cells and Electrode Potentials

1. Describe the construction and working of a Daniell cell.
2. Define the standard electrode potential and explain the significance of a negative value.
3. What is the role of the salt bridge in an electrochemical cell?

Nernst Equation and Gibbs Energy

1. Write the Nernst equation and derive the relation $\Delta G^\circ = -nFE^\circ$.
2. Calculate the emf of a cell at 298 K given $E^\circ = 1.10$ V and $Q = 0.01$.
3. Derive the relation between the equilibrium constant and the standard cell potential.

Part B — Integrated problems (multi-topic)

1. Calculate the oxidation number of Mn in KMnO_4 and of Cr in $\text{K}_2\text{Cr}_2\text{O}_7$.
2. Balance the equation $\text{MnO}_4^- + \text{Fe}^{2+} + \text{H}^+ \rightarrow \text{Mn}^{2+} + \text{Fe}^{3+} + \text{H}_2\text{O}$ by the ion-electron method.
3. For the cell $\text{Zn}|\text{Zn}^{2+}(0.1\text{ M})||\text{Cu}^{2+}(0.01\text{ M})|\text{Cu}$ with $E^\circ = 1.10$ V, calculate the emf at 298 K.
4. Calculate ΔG° for a cell with $n = 2$ and $E^\circ = 1.10$ V. ($F = 96500\text{ C mol}^{-1}$)
5. Calculate the equilibrium constant for a cell reaction with $E^\circ = 1.10$ V and $n = 2$ at 298 K.

Answer Key

Oxidation and Reduction

Oxidation: loss of electrons; reduction: gain of electrons. In $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$, Cu^{2+} is the oxidising agent and Zn the reducing agent. Oxidation and reduction always occur together (redox).

Oxidation Number and Its Rules

Rules: free element = 0; H = +1 (except hydrides, -1); O = -2 (except peroxides, -1 and OF_2 , +2); sum of oxidation numbers equals the charge. KMnO_4 : Mn = +7; $\text{K}_2\text{Cr}_2\text{O}_7$: Cr = +6.

Balancing Redox Equations — Oxidation-Number Method

Identify atoms whose oxidation numbers change, equate the increase and decrease, then balance other atoms. $\text{Fe}^{2+} + \text{MnO}_4^- + 8\text{H}^+ \rightarrow 5\text{Fe}^{3+} + \text{Mn}^{2+} + 4\text{H}_2\text{O}$ (after balancing).

Ion-Electron (Half-Reaction) Method

Split into half-reactions, balance atoms and charge, then add. $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$. Combine with $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$ ($\times 5$).

Types of Redox Reactions

Combination ($2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$), decomposition ($2\text{H}_2\text{O}_2 \rightarrow 2\text{H}_2\text{O} + \text{O}_2$), displacement ($\text{Zn} + \text{CuSO}_4 \rightarrow \text{ZnSO}_4 + \text{Cu}$) and disproportionation ($\text{Cl}_2 + 2\text{NaOH}$).

Electrochemical Cells and Electrode Potentials

Anode: oxidation; cathode: reduction. Standard hydrogen electrode: $E^\circ = 0$ V. Salt bridge maintains neutrality. A more negative E° means a more reactive metal.

Nernst Equation and Gibbs Energy

$E = E^\circ - (0.059/n)\log Q$; $\Delta G^\circ = -nFE^\circ$; at equilibrium $E = 0$. K is obtained from $E^\circ = (0.059/n)\log K$.

Integrated problems

1: Mn = +7; Cr = +6. 2: $\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O}$. 3: $E = 1.10 - (0.059/2)\log(0.1/0.01) = 1.10 - 0.0295 = 1.0705$ V. 4: $\Delta G^\circ = -2 \times 96500 \times 1.10 = -2.123 \times 10^5\text{ J} = -212.3\text{ kJ}$. 5: $\log K = 2 \times 1.10/0.059 \approx 37.3$, $K \approx 2 \times 10^{37}$.