ELECTROCHEMISTRY – NERNST EQUATION

A. Objective Questions (5 marks)

1. Determine the equilibrium constant, K_{eq} at 25°C for the reaction

$$2Br^-(aq) + I_2(s) \longrightarrow Br_2(l) + 2I^-(aq)$$

[$E°(Br_2|Br^-) = +1.07$ V ; $E°(I_2|I^-) = +0.53$ V]

A. 5.7×10^{-19} B. 1.7×10^{54}

C. 5.7×10^{-55} D. 1.9×10^{18}

2. What is the electrode potential for the half-reaction below at pH 4?

$$H_2O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O \qquad E° = +1.77 \text{ V}$$

A. +1.48 V B. +1.53 V

C. +1.62 V D. +1.69 V

3. The electromotive force (e.m.f) of a Daniell cell, $Zn(s) | Zn^{2+}(aq) || Cu^{2+} | Cu(s)$ can be increased by

A. increasing the concentration of copper ions.

B. increasing the concentration of zinc ions.

C. Enlarging the surface area of the electrodes.

D. increase the pressure of the hydrogen gas.

4. The following redox reaction occurs in the voltaic cell shown below.

$$2MnO_4^-(aq) + 8H^+(aq) + 3Cu(s) \longrightarrow 2MnO_2(s) + 4H_2O(l) + 3Cu^{2+}(aq)$$

What is the potential of the cell at 25°C when $[MnO_4^-] = 0.20$ M, $[Cu^{2+}] = 0.003$ M and $[H^+] = 0.10$ M?

A. +1.15 V B. +1.00 V

C. +0.23 V D. +1.50 V

5. The relationship between electrode potential and the concentration of ions is given by the following Nernst Equation:

$$\mathbf{E_{cell} = E°_{cell} - \frac{0.0592}{n} \log \frac{[oxidised\ ion]}{[reduced\ ion]}}$$

The standard e.m.f for the cell $Pt(s) | Pt^{2+}(aq, 1\ M) || Au^{3+}(aq, 1M) | Au(s)$ is +0.20 V. If the concentration of Pt^{2+} ions is increased by 2 M whilst reducing the concentration of Au^{3+} by 0.80 M, what will be the new cell potential?

A. +0.12 V B. +0.17 V

C. +0.19 V D. +0.22 V

B. Subjective Questions

1. Refer to the reaction below

$$3H_2O_2(aq) + 6H^+(aq) + 2Au(s) \longrightarrow 2Au^{3+}(aq) + 6H_2O(l)$$

(i) Identify the reducing agent and oxiding agent for the reaction.

(ii) Determine the pH of the half-cell if emf of the cell is + 0.2711 V
$[Au^{3+}] = 0.25$ M, and $[H_2O_2] = 1.50$ M

(iii) If the the concentration of $[Au^{3+}]$ is increased, what would happen to the emf of the cell ? Explain

[Given $\mathbf{E^\circ_{Au^{3+}|Au}} = +\mathbf{1.50\,V}$; $\mathbf{E^\circ_{H_2O_2|H_2O}} = +\mathbf{1.77\,V}$]

[10 marks]

ANSWERS

1.	2.	3.	4.	5.
C	**B**	**A**	**A**	**C**

NO.	PART	SUGGESTED ANSWER
1.	**(i)**	**Reducing agent : Au** 1 mark **Oxidising agent : H_2O_2** 1 mark
	(ii)	$\mathbf{E°_{cell} = E°_{cathode} - E°_{anode}}$ 1 mark $= \mathbf{E°_{H_2O_2|H_2O} - E°_{Au^{3+}|Au}}$ $= +1.77\ V - 1.50\ V$ $= \mathbf{+0.27\ V}$ 1 mark
	(iii)	$E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{n}\log Q$ $\mathbf{E_{cell} = E^{\circ}_{cell} - \frac{0.0592}{n}\log \frac{[Au^{3+}]^2}{[H_2O_2]^3[H^+]^6}}$ 1 mark $0.2711\ V = 0.27\ V - \frac{0.0592}{6}\log \frac{(0.25)^2}{(1.50)^3[H^+]^6}$ **n = 6** 1 mark $\log \frac{(0.25)^2}{(1.50)^3[H^+]^6} = -0.1115$ $\frac{(0.25)^2}{(1.50)^3[H^+]^6} = 10^{-0.1115}$ $[H^+] = \mathbf{0.5368\ M}$ 1 mark $pH = -\log [H^+]$ $= -\log 0.5368$ $= \mathbf{0.27}$ 1 mark
	(iv)	If the concentration of Au^{3+} is increased, Q increases, log Q increases. 1 mark Therefore, emf or E_{cell} decreases. 1 mark
		TOTAL = 10