Redox Reactions and Electrochemistry JEE Main PYQ - 3

Total Time: 25 Minute **Total Marks:** 40

Instructions

Instructions

1. Test will auto submit when the Time is up.
2. The Test comprises of multiple choice questions (MCQ) with one or more correct answers.
3. The clock in the top right corner will display the remaining time available for you to complete the examination.

Navigating & Answering a Question

1. The answer will be saved automatically upon clicking on an option amongst the given choices of answer.
2. To deselect your chosen answer, click on the clear response button.
3. The marking scheme will be displayed for each question on the top right corner of the test window.

Redox Reactions and Electrochemistry

1. The oxidation states of Cr in $[Cr(H_2O)_6]Cl_3, [Cr(C_6H_6)_2]$, and $K_2[Cr(CN)_2(O)_2(O_2)(NH_3)]$ respectively are : [2018] **(+4, -1)**

a. +3, +4, and +6

b. +3, +2, and +4

c. +3, 0, and +6

d. +3, 0, and +4

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2. Oxidation state of sulphur in anions $SO_3^{2-}.S_2O_4^{2-}$ and $S_2O_6^{2-}$ increases in the orders: **(+4, -1)**

22 Apr 2018

a. $S_2O_6^{2-} < S_2O_4^{2-} < SO_3^{2-}$

b. $SO_6^{2-} < S_2O_4^{2-} < SO_6^{2-}$

c. $S_2O_4^{2-} < SO_3^{2-} < S_2O_6^{2-}$

d. $S_2O_4^{2-} < S_2O_6^{2-} < SO_3^{2-}$

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3. Highest oxidation state of Mn is exhibited in Mn_2O_7 The correct statements about Mn_2O_7 are **(+4, -1)**

7 Jan 2020 Shift I

(A) Mn is tetrahedrally surrounded by oxygen atoms
(B) Mn is octahedrally surrounded by oxygen atoms
(C) Contains $Mn - O - Mn$ bridge
(D) Contains $Mn - Mn$ bond
Choose the correct answer from the options given below:

a. A and D only

b. A and C only

c. B and C only

d. B and D only

4. Sum of oxidation states of bromine in bromic acid and perbromic acid is (+4, -1)

Apr 8 2019 Shift II

5. H_2O_2 acts as a reducing agent in (+4, -1)

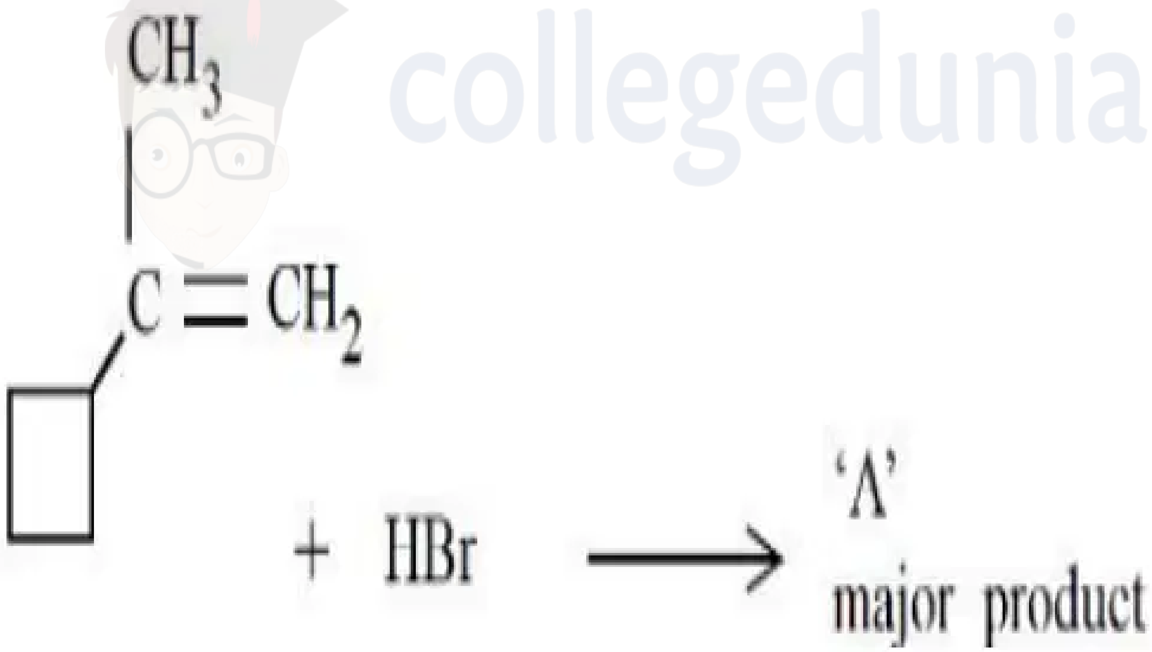
   a. $Na_2S + 4H_2O_2 \rightarrow Na_2SO_4 + 4H_2O$

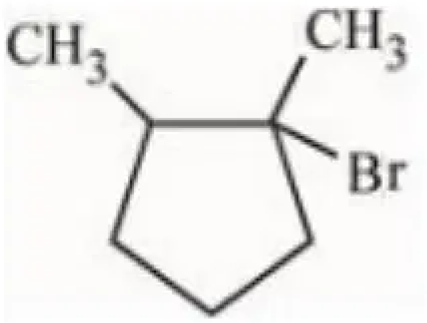
   b. $Mn^{2+} + 2H_2O_2 \rightarrow MnO_2 + 2H_2O$

   c. $2Fe^{2+} + 2H^{+} + H_2O_2 \rightarrow 2Fe^{3+} + 2H_2O$

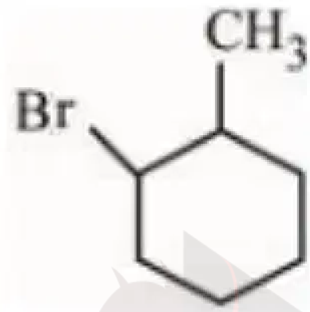
   d. $2NaOCl + H_2O_2 \rightarrow 2NaCl + H_2O + O_2$

6. In the following given reaction, ' A ' is (+4, -1)

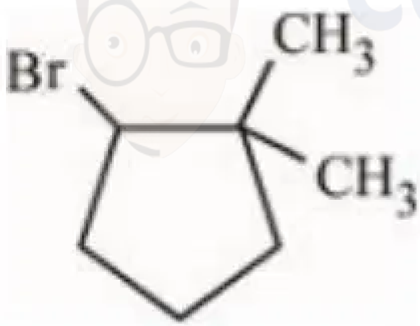




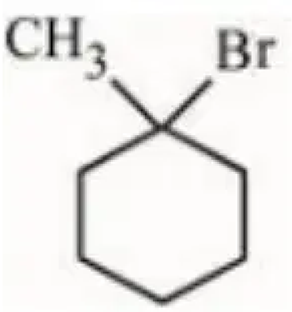
a.



b.



c.



d.

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7. In the cell $Pt(s)|H_2(g, 1bar|HCl(aq)|Ag(s)|Pt(s)$ the cell potential is 0.92 when a 10^{-6} molal HCl solution is used. THe standard electrode potential of $(AgCl/Ag, Cl^-)$ electrode is : $\{\text{given}, \frac{2.303RT}{F} = 0.06V\ at\ 298K\}$ **(+4, -1)**

10 Jan 2019 Shift II

a. 0.20 V

b. 0.76 V

c. 0.40 V

d. 0.94 V

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8. At what pH, given half cell $MnO_4^-(01M) \mid Mn^{2+}(0001M)$ will have electrode potential of $1282\,V$? (Nearest Integer) (Given: $E^o_{MnO_4^-|Mn^{2+}} = 154\,V$, $\frac{2303RT}{F} = 0059\,V$) **(+4, -1)**

1 Feb 2023 Shift I

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9. $Pt(s) \mid H2(g)(1atm) \mid H+(aq,\ [H+]=1) \parallel Fe3+(aq), Fe2+(aq) \mid Pt(s)$ **(+4, -1)**
Given $E^\circ_{Fe^{3+}Fe^{2+}} = 0.771V$ and $E^\circ_{H^{+1/2}H_2} = 0\,V, T = 298K$

25 Jan 2023 Shift I

If the potential of the cell is 0.712V, the ratio of concentration of $Fe2+$ to $Fe3+$ is

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10. The anodic half-cell of lead-acid battery is recharged unsing electricity of 0.05 Faraday. The amount of $PbSO_4$ electrolyzed in g during the process in : (Molar mass of $PbSO_4 = 303\ g\ mol^{-1}$) **(+4, -1)**

9 Jan 2019 Shift I

a. 22.8

b. 15.2

c. 7.6

d. 11.4

Answers

1. **Answer: c**

Explanation:

$$[Cr\,(H_2O)_6]\,Cl_3$$
$$\Rightarrow x + 0 \times 6 - 1 \times 3 = 0$$
$$\therefore x = +3$$
$$[Cr\,(C_6H_6)_2]$$
$$\Rightarrow x + 2 \times 0 = 0$$
$$x = 0$$
$$K_2[Cr\,(CN)_2\,(O_2)\,(O_2)\,NH_3]$$
$$\Rightarrow 1 \times 2 + x - 1 \times 2 - 2 \times 2 - 2 \times 1 = 0$$
$$\Rightarrow x - 6 = 0$$
$$x = +6$$

Concepts:

1. Oxidation Number:

Oxidation number, also called **oxidation state**, the total number of electrons that an at om either gains or loses in order to form a chemical bond with another atom.

Oxidation number of an atom is defined as the charge that an atom appears to have on forming ionic bonds with other heteroatoms. An atom having higher electronegativity (even if it forms a covalent bond) is given a negative oxidation state.

The definition, assigns oxidation state to an atom on conditions, that the atom –

1. Bonds with heteroatoms.
2. Always form ionic bonding by either gaining or losing electrons, irrespective of the actual nature of bonding.

Oxidation number is a formalized way of keeping track of oxidation state.

Read More: Oxidation and Reduction

Way To Find Oxidation Number Of An Atom?

Oxidation number or state of an atom/ion is the number of electrons an atom/ion that the molecule has either gained or lost compared to the neutral atom. Electropositive metal atoms, of group I, 2 and 3 lose a specific number of electrons and have always constant positive oxidation numbers.

In molecules, more electronegative atom gain electrons from a less electronegative atom and have negative oxidation states. The numerical value of the oxidation state is equal to the number of electrons lost or gained.

Oxidation number or oxidation state of an atom or ion in a molecule/ion is assigned by:

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2. **Answer: c**

Explanation:

In SO_3^{--} $x+3(-2)=-2; x=+4$ In S_2O_4 $2x+4(-2)=-2$ $2x-8=-2$ $2x=6; x=+3$ In $S_2O_6^{2-}$ $2x+6(-2)-2$ $2x=10; x=+5$ hence the correct order is $S_2O_4^{--} < SO_3^{--} < S_2O_6^{--}$

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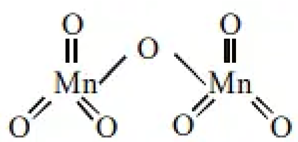
3. **Answer: b**

Explanation:

The correct option is (b) A and C only



Concepts:

1. **Redox Reactions:**

Redox Reaction:

Redox reactions are chemical reactions where oxidation and reduction take place simultaneously. In this type of reaction, there is a gain of electrons for one chemical species while the other loses electrons or simply involves transfer of electrons. The species that loses electrons is oxidized while the one that gains electrons is reduced.

Types of Redox Reactions:

Redox reactions can be differentiated into 4 categories namely combination reactions, decomposition reactions, displacement reactions, and disproportionation reactions. Each is explained separately below:

Combination Reaction:

In this, the molecules combine to form new compounds. For example, when magnesium reacts to nitrogen.

Decomposition Reaction:

Opposite to the combination reaction, here there is a breakdown of compounds to simpler substances. For example, electrolysis of water.

Displacement Reaction:

In this, the more reactive metal will displace the less reactive one in a chemical reaction. The reactivity of an element is represented in a series called the reactivity series (arranged in decreasing order of reactivity) which makes it easier to determine the chemical reaction and its products.

Disproportionation Reaction:

This is a peculiar type of reaction where an element showing a particular oxidation state will be oxidized and reduced simultaneously. Another thing to note is that these reactions will always have an element that can exhibit three oxidation states.

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4. Answer: 12 - 12

Explanation:

The correct answer is 12.
$HBrO_3$ (Bromic acid)
Ox. State of $Br = +5$
$HBrO_4$ (per bromic acid)
OX. State of $Br = +7$
Sum of Ox. State $= 12$

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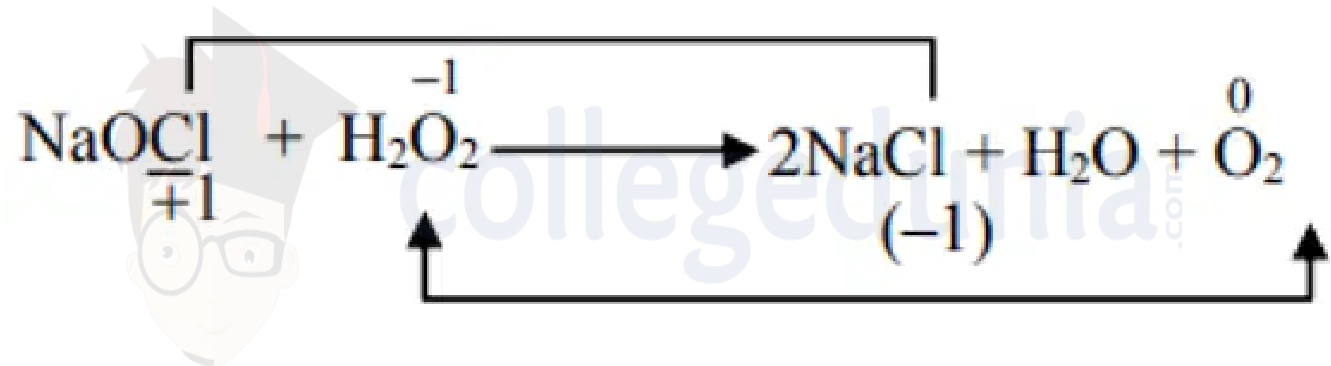
5. **Answer: d**

Explanation:

Correct answer is (d) $2NaOCl + H_2O_2 \rightarrow 2NaCl + H_2O + O_2$



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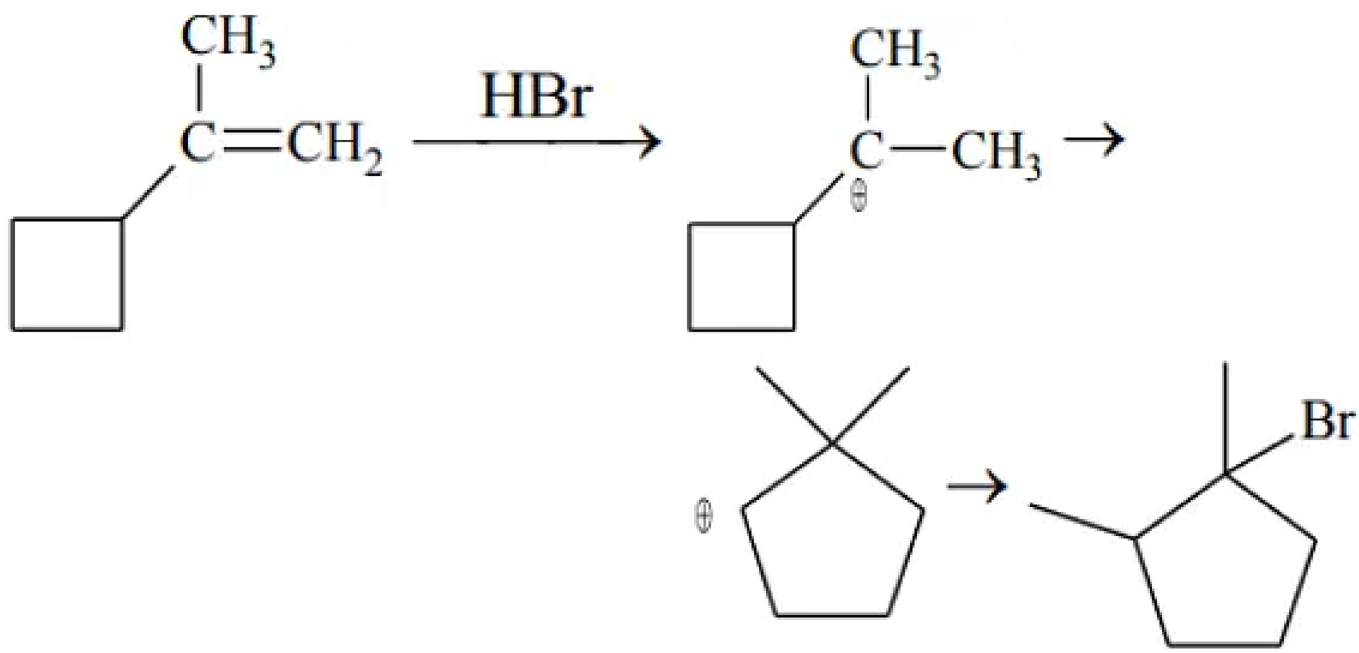
6. **Answer: d**

Explanation:

Correct answer is (d)



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---

7. **Answer: a**

Explanation:

$Pt(s)H_2(g, 1bar)HCl(aq)AgCl(s)Ag(s)|Pt(s)$

Anode: $H_2 -> [10^{-6} m]2H^+ + 2e \times 1$

Cathode : $e^- + AgCl(s) -> Ag(s) + Cl^-(aq)$

$H_2(g)l + AgCl(s) -> 2H^+ + 2Ag(s) + 2Cl^-(aq)$

$E_{cell} = E^0_{cell} - \frac{0.06}{2} \log_{10}((H^+)^2.(Cl^-)^2)$

${.925 = (E^{0}_{H_2 /H^{+}}} + E^{0}_{AgCl / Ag, Cl^{-}}) - \frac{0.06}{2} \log_{10} ((10^{-6})^{2} (10^{-6})^{2})}$

$.92 = 0 + E^0_{AgCl/Ag,Cl^-} - 0.03 \log_{10}(10^{-6})^4$

$E^0_{AgCl}/Ag, Cl^- = .9.2 + .03 \times -24 = 0.2\ V$

Concepts:

1. Galvanic Cells:

Galvanic cells, also known as voltaic cells, are electrochemical cells in which spontaneous oxidation-reduction reactions produce electrical energy. It converts chemical energy to electrical energy.

It consists of two half cells and in each half cell, a suitable electrode is immersed. The two half cells are connected through a salt bridge. The need for the salt bridge is to keep the oxidation and reduction processes running simultaneously. Without it, the electrons liberated at the anode would get attracted to the cathode thereby stopping the reaction on the whole.

Working Principle Of Galvanic Cell:

1. Take two beakers containing electrolytic solutions of copper sulphate and zinc sulphate are taken. It is connected via a salt bridge containing an aqueous solution of potassium chloride.
2. Zinc and copper electrodes are immersed in the respective electrodes and connected through a voltmeter to measure the electrical potential.
3. Zinc which acts as the anode readily undergoes an oxidation process and acquires a negative charge.
4. The electrons travel through the salt bridge and undergo a reduction process at the copper cathode.
5. Thus the cathode would acquire a positive charge.
6. This flow of electrons from the anode to the cathode induces a flow of electric current in the opposite direction which shall be measured by the voltmeter.

Types of Voltaic Cell:

- **Primary Cell**
  - Dry Cell
  - Mercury Cell
  - Alkaline Cell
- **Secondary Cell**
  - Nickel-Cadmium Cell
  - Lead-Acid Cell
  - Lithium-Ion Cell

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8. Answer: 3 - 3

Explanation:

$$MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$$

$$E = E^\circ - \frac{0.059}{5}\log\frac{[Mn^{2+}]}{[MnO_4^-][H^+]^8}$$

$$1.282 = 1.54 - \frac{0.059}{5}\log\frac{10^{-3}}{10^{-1}\times[H^+]^8}$$

$\frac{0.258 \times 5}{0.059} = \log \frac{10^{-2}}{[H^+]^8}$
$\Rightarrow 21.86 = -2 + 8pH$
$\therefore pH = 2.98$
$\simeq 3$

So, the correct answer is 3.

Concepts:

1. Electrochemical Cells:

An electrochemical cell is a device that is used to create electrical energy through the chemical reactions which are involved in it. The electrical energy supplied to electrochemical cells is used to smooth the chemical reactions. In the electrochemical cell, the involved devices have the ability to convert the chemical energy to electrical energy or vice-versa.

Classification of Electrochemical Cell:

Cathode

- Denoted by a positive sign since electrons are consumed here
- A reduction reaction occurs in the cathode of an electrochemical cell
- Electrons move into the cathode

Anode

- Denoted by a negative sign since electrons are liberated here
- An oxidation reaction occurs here
- Electrons move out of the anode

Types of Electrochemical Cells:

Galvanic cells (also known as Voltaic cells)

- Chemical energy is transformed into electrical energy.
- The redox reactions are spontaneous in nature.
- The anode is negatively charged and the cathode is positively charged.
- The electrons originate from the species that undergo oxidation.

Electrolytic cells

- Electrical energy is transformed into chemical energy.
- The redox reactions are non-spontaneous.
- These cells are positively charged anode and negatively charged cathode.
- Electrons originate from an external source.

---

9. **Answer: 10 - 10**

Explanation:

The correct answer is 10.

$\frac{1}{2}H_2(g) + Fe^{3+}(aq.) \longrightarrow H^+(aq) + Fe^{2+}(aq.)$

$E = E^o - \frac{0.059}{1}\log\frac{[Fe^{2+}]}{[Fe^{3+}]}$

$\Rightarrow 0.712 = (0.771 - 0) - \frac{0.059}{1}\log\frac{[Fe^{2+}]}{[Fe^{3+}]}$

$\Rightarrow \log\frac{[Fe^{2+}]}{[Fe^{3+}]} = \frac{(0.771-0712)}{0.059} = 1$

$\Rightarrow \frac{[Fe^{2+}]}{[Fe^{3+}]} = 10$

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---

10. **Answer: b**

Explanation:

(A) $ { $\underset{\text{ 0.05/2 mole}}{ { PbSO4(s) }}$+2OH^{-} -> PbO_2 + H2SO4+$ \underset{\text{ 0.05F}}{ { 2e^{-}}}$ }$ (C) $ { $\underset{\text{ 0.05/2 mole 0.05 F}}{ { PbSO4 + 2e^{-}}}$ + 2H^{+} -> Pb(s) + H2SO4}$ $n_T(PbSO_4) = 0.05\, mole$ $m_{PbSO_4} = 0.05 \times 303 = 15.2\ gm$

Concepts:

1. Nernst Equation:

This equation relates the equilibrium cell potential (also called the Nernst potential) to its concentration gradient across a membrane. If there is a concentration gradient for the ion across the membrane, an electric potential will form, and if selective ion channels exist the ion can cross the membrane.

The equation may be written:

$E_{cell} = E^o_{cell} - (RT/nF)\ln Q$

E_{cell} = cell potential under nonstandard conditions (V)
E^o_{cell} = cell potential under standard conditions
R = gas constant, which is 8.31 (volt-coulomb)/(mol-K)
T = temperature (K)
n = number of moles of electrons exchanged in the electrochemical reaction (mol)
F = Faraday's constant, 96500 coulombs/mol
Q = reaction quotient, which is the equilibrium expression with initial concentrations rather than equilibrium concentrations

Sometimes it is helpful to express the Nernst equation differently:

$E_{cell} = E^o_{cell} - (2.303*RT/nF)\log Q$

at 298K, $E_{cell} = E^o_{cell} - (0.0591\ V/n)\log Q$