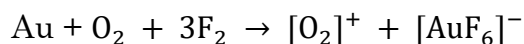


GENESIS TUTORIALS**ASSIGNMENT:1- ELECTROCHEMISTRY**

1. Consider the following reaction:



Which statement is correct about the redox changes in this reaction?

- (a) Au is oxidized; O is oxidized; F is reduced
- (b) Au is reduced; O is oxidized; F is reduced
- (c) Au is oxidized; O does not undergo a redox change; F is reduced
- (d) Au is reduced; O is oxidized; F does not undergo a redox change

2. The Daniel cell is

- (a) $\text{Pt}_I(\text{s}) | \text{Zn}(\text{s}) | \text{Zn}^{2+}(\text{aq}) || \text{Cu}^{2+}(\text{aq}) | \text{Cu}(\text{s}) | \text{Pt}_{II}(\text{s})$
- (b) $\text{Pt}_I(\text{s}) | \text{Zn}(\text{s}) | \text{Zn}^{2+}(\text{aq}) || \text{Ag}^+(\text{aq}) | \text{Ag}(\text{s}) | \text{Pt}_{II}(\text{s})$
- (c) $\text{Pt}_I(\text{s}) | \text{Fe}(\text{s}) | \text{Fe}^{2+}(\text{aq}) || \text{Cu}^{2+}(\text{aq}) | \text{Cu}(\text{s}) | \text{Pt}_{II}(\text{s})$
- (d) $\text{Pt}_I(\text{s}) | \text{H}_2(\text{g}) | \text{H}_2\text{SO}_4(\text{aq}) || \text{Cu}^{2+}(\text{aq}) | \text{Cu}(\text{s}) | \text{Pt}_{II}(\text{s})$

3. Given: A. $\text{Fe}(\text{OH})_2(\text{s}) + 2\text{e}^- \rightarrow \text{Fe}(\text{s}) + 2\text{OH}^-(\text{aq}); E^0 = -0.877\text{V}$

B. $\text{Al}^{3+}(\text{aq}) + 3\text{e}^- \rightarrow \text{Al}(\text{s}); E^0 = -1.66\text{V}$

C. $\text{AgBr}(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s}) + \text{Br}^-(\text{aq}); E^0 = 0.071\text{V}$

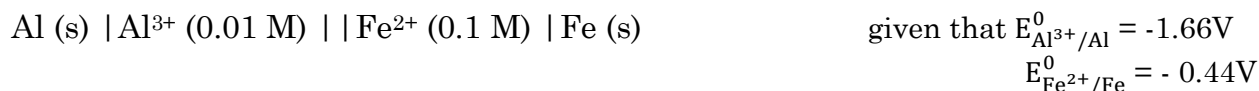
The overall reaction for the cells in the direction of spontaneous change would be

- (a) Cell with A and B : Fe reduced
Cell with A and C : Fe reduced
- (b) Cell with A and B : Fe reduced
Cell with A and C : Fe oxidized
- (c) Cell with A and B : Fe oxidized
Cell with A and C : Fe oxidized
- (d) Cell with A and B : Fe oxidized
Cell with A and C : Fe reduced

4. The cell voltage of Daniel cell $[Zn | Zn SO_4 (aq) || CuSO_4 (aq) | Cu]$ is 1.07 V. If reduction potential of $Cu^{2+} | Cu$ is 0.34 the reduction potential of $Zn^{2+} | Zn$ is –

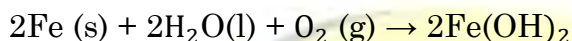
- (a) 1.41 V (b) – 1.41 V (c) 0.73 V (d) – 0.73 V

5. The cell potential for the following electrochemical system at 25 °C is:



- (a) 1.23 V (b) 1.21 V (c) 1.22 V (d) – 2.10 V

6. A 19th century iron bridge is protected from corrosion by connecting it to a block of metal (sacrificial anode), which is replaced annually. The corrosion of iron, represented by the chemical equation:



Which of the following metals is best suited as sacrificial anode?

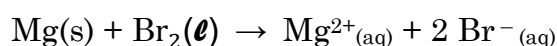
- (a) Ag : $Ag^+ + e^- \rightarrow Ag(s)$ $E^0 = +0.80V$
 (b) Cd : $Cd^{2+} + 2e^- \rightarrow Cd(s)$ $E^0 = -0.40V$
 (c) Cu : $Cu^+ + 2e^- \rightarrow Cu(s)$ $E^0 = +0.34V$
 (d) Mg : $Mg^{2+} + 2e^- \rightarrow Mg(s)$ $E^0 = -2.36V$

7. Which combination of the following half-cell reactions will produce a cell with the largest potential?

- (i) $A + e^- \rightarrow A^-$ $E^0 = -0.24 V$
 (ii) $B^- + e^- \rightarrow B^{2-}$ $E^0 = 1.25 V$
 (iii) $C^- + 2 e^- \rightarrow C^{3-}$ $E^0 = -1.25 V$
 (iv) $D + 2 e^- \rightarrow D^{2-}$ $E^0 = 0.68 V$
 (v) $E + 4 e^- \rightarrow E^{4-}$ $E^0 = 0.38 V$

- (a) (i) and (ii) (b) (ii) and (iii) (c) (iii) and (iv) (d) (iv) and (v)

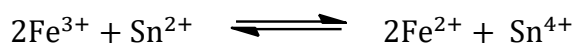
8. What is the standard cell potential for the reaction below?



The standard reduction potential is – 2.37 V for the Mg^{2+}/ Mg half-cell and + 1.09 V for the Br_2 / Br^- half cell

- (a) + 1.28 V (b) – 1.28 V (c) + 3.46 V (d) – 3.46 V

9. The equilibrium constant for an electrochemical reaction



$$[E^0 \left(\frac{\text{Fe}^{3+}}{\text{Fe}^{2+}} \right) = 0.75 \text{ V}, E^0 \left(\frac{\text{Sn}^{4+}}{\text{Sn}^{2+}} \right) = 0.15 \text{ V}, E^0 \left(\frac{2.303 RT}{F} \right) = 0.06 \text{ V}]$$

- (a) 10^{10} (b) 10^{20} (c) 10^{30} (d) 10^{40}

10. The electrochemical cell potential (E), after the reactants and products reach equilibrium, is (E^0 is the standard cell potential and n is the number of electrons involved)

- (a) $E = E^0 + \frac{nF}{RT}$ (b) $E = E^0 - \frac{RT}{nF}$ (c) $E = E^0$ (d) $E = 0$

11. The EMF of the concentration cell



- (a) 0.001 V (b) 0.025 V (c) 0.059 V (d) 0.118 V

12. The amount of copper (atomic weight 63.6 g mol^{-1}) deposited on passing a current of 2 ampere through a solution of copper sulfate for 965 seconds is –

- (a) 636 g (b) 63.6g (c) 6.36 g (d) 0.636 g

13. A current equivalent to 0.1 faraday of charge is passed through a solution of CuSO_4 to deposit pure copper at the cathode. The weight of copper deposited is:

- (a) 6.4 g (b) 12.8 g (c) 3.2 g (d) 1.6 g

14. How many faradays are required to reduce one mole of M_nO_4^- to M_n^{2+} ?

- (a) 1 (b) 2 (c) 3 (d) 5

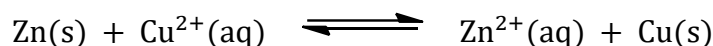
15. Passage of 30 A current for 70.2 minutes correspond to:

- (a) 1.31 C (b) 1.31 F (c) 2100 C (d) 2100 F

16. The temperature derivative of electrochemical cell potential E at constant pressure, $\left(\frac{\partial E}{\partial T} \right)_p$ is given by

- (a) $-\frac{\Delta S}{nF}$ (b) $\frac{\Delta S}{nF}$ (c) $\frac{\Delta S}{nFT}$ (d) $-\frac{\Delta S}{nFT}$

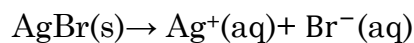
17. The standard cell potential for the reaction,



is + 1.10 V. The Gibbs free energy change during the reaction is ($F = 96500 \text{ coulomb mol}^{-1}$)

- (a) $-21.2 \text{ kJ mol}^{-1}$ (b) $+212 \text{ kJ mol}^{-1}$ (c) -212 kJ mol^{-1} (d) -212 J mol^{-1}

18. The standard free energy of the reaction:



is closest to

($E^\circ(\text{AgBr}/\text{Ag}, \text{Br}^-) = +0.07 \text{ V}$, $E^\circ(\text{Ag}^+/\text{Ag}) = 0.80 \text{ V}$; $F = 96500 \text{ C mol}^{-1}$)

- (a) 7 kJmol^{-1} (b) 70 Jmol^{-1} (c) 70 kJmol^{-1} (d) 7 Jmol^{-1}

