Chapter 9. ELECTROCHEMISTRY

1. C is correct.

2. A is correct. Since the reduction potentials for $Li^+(aq)$ and $Na^+(aq)$ are both negative, neither process is spontaneous. Reversing each of the reactions listed in the table supports that Li (*s*) is more easily oxidized than Na (*s*), because 3.05 V > 2.71 V.

3. B is correct.

4. D is correct.

5. C is correct. One faraday is equivalent to 1 mole of electrons. To electrolyze 5 moles of H_2O_2, 10 moles of electrons are required. 10 moles of electrons = 10 faradays.

6. D is correct. One mole of e^- has a charge of 96,500 coulombs.

Conversion factor: 1 mole e^- / 96,500 coulombs

10 moles of H_2O need 10 moles of e^-.

To calculate charge:

10 moles of e^- × 96,500 coulombs / 1 mole e^- = 965,000 coulombs

7. C is correct. 1 amp = 1 coulomb/sec.

2 moles of H_2O need 2 moles of e^-.

To calculate moles of H_2O:

$$2 \text{ moles } e^- \times \frac{96{,}500 \text{ coul}}{1 \text{ mole } e^-} \times \frac{1 \text{ sec}}{2 \text{ coul}} = 9.65 \times 10^4 \text{ sec}$$

8. A is correct. Half reaction:

$NO_3^- \rightarrow NH_4^+$

Balancing half-reaction in acidic conditions:

Step 1: Balance all atoms except for H and O

$NO_3^- \rightarrow NH_4^+$ (N is already balanced)

Step 2: To balance oxygen, add H_2O to the side with less oxygen atoms

$NO_3^- \rightarrow NH_4^+ + 3H_2O$

Step 3: To balance hydrogen, add H^+ to the opposing side of H_2O added in the previous step

$NO_3^- + 10\ H^+ \rightarrow NH_4^+ + 3H_2O$

Step 4: Balance charges by adding electrons to the side with higher/more positive total charge

Total charge on left side: 10(+1) – 1 = +9

Total charge on right side: +1

Add 8 electrons to left side:

$NO_3^- + 10\ H^+ + 8e^- \rightarrow NH_4^+ + 3H_2O$

Total electrons required: $8e^-$ on the left side

9. D is correct.

10. B is correct.

11. A is correct. The oxidizing and reducing agents are always reactants, not products, in a redox reaction. The oxidizing agent is the species that gets reduced that is, the one that gains electrons; this is Cu^{2+}.

12. D is correct.

13. B is correct.

14. B is correct. In all cells, reduction occurs at the cathode while oxidation occurs at the anode (The following mnemonic can be used: RED CAT / AN OX).

15. C is correct. Calculating mass of metal deposited in cathode.

Step 1: Calculate total charge using current and time

$Q = \text{current} \times \text{time}$

$Q = 5.2\ A \times (45.0\ \text{minutes} \times 60\ \text{s/minute})$

$Q = 14040\ A \cdot s = 14040\ C$

Step 2: Calculate moles of electron which has that same amount of charge

$\text{moles } e^- = Q / 96500\ C/mol$

$\text{moles } e^- = 14040\ C / 96500\ C/mol$

$\text{moles } e^- = 0.146\ mol$

Step 3: Calculate moles of metal deposit

The solution contains chromium (III) sulfate. Half-reaction of chromium (III) ion reduction:

$$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$$

Use the moles of electron from the previous calculation to calculate moles of Cr:

$\text{moles Cr} = (\text{coefficient Cr} / \text{coefficient electron}) \times \text{moles electron}$

$\text{moles Cr} = (1/3) \times 0.146\ mol$

$\text{moles Cr} = 0.0487\ mol$

Step 4: Calculate mass of metal deposit

$\text{mass Cr} = \text{moles Cr} \times \text{molecular mass of Cr}$

$\text{mass Cr} = 0.0487\ mol \times 52.00\ g/mol$

$\text{mass Cr} = 2.5\ g$

16. C is correct.

17. A is correct.

18. D is correct. Oxidation always occurs at the anode while reduction always occurs at the cathode. Mg^{2+} does not lose any more electrons and Cl^- does not gain any more electrons.

19. C is correct.

20. A is correct.

21. D is correct. E° tends to be negative and G positive because electrolytic cells are nonspontaneous. Electrons must be forced into the system for the reaction to proceed.

22. D is correct. Balancing half-reaction in basic conditions:

The first few steps are identical to balancing reactions in acidic conditions.

Step 1: Balance all atoms except for H and O

$C_8H_{10} \rightarrow C_8H_4O_4^{2-}$ C is already balanced

Step 2: To balance oxygen, add H_2O to the side with less oxygen atoms

$C_8H_{10} + 4H_2O \rightarrow C_8H_4O_4^{2-}$

Step 3: To balance hydrogen, add H^+ to the opposing side of H_2O added in the previous step

$C_8H_{10} + 4H_2O \rightarrow C_8H_4O_4^{2-} + 14H^+$

This next step is the unique additional step for basic conditions.

Step 4: Add equal amounts of OH^- on both sides. The number of OH^- should match the number of H^+ ions. Combine H^+ and OH^- on the same side to form H_2O. If there are H_2O molecules on both sides, subtract accordingly to end up with H_2O on one side only.

There are 14 H^+ ions on the right, so add $14OH^-$ ions on both sides:

$C_8H_{10} + 4H_2O + 14OH^- \rightarrow C_8H_4O_4^{2-} + 14H^+ + 14OH^-$

Combine H^+ and OH^- ions to form H_2O:

$C_8H_{10} + 4H_2O + 10OH^- \rightarrow C_8H_4O_4^{2-} + 14H_2O$

There are H_2O molecules on both sides. Those will cancel out and some H_2O will remain on one side:

$C_8H_{10} + 14OH^- \rightarrow C_8H_4O_4^{2-} + 10H_2O$

Step 5: Balance charges by adding electrons to the side with higher/more positive total charge

Total charge on left side: $14(-1) = -14$

Total charge on right side: -2

Add 12 electrons to right side:

$C_8H_{10} + 14OH^- \rightarrow C_8H_4O_4^{2-} + 10H_2O + 12e^-$

Total electrons required: $12e^-$ on the right side

23. A is correct.

24. C is correct.

25. B is correct. Oxidation always occurs at the anode while reduction always occurs at the cathode. In reaction 2, each I^- anion must lose one electron to produce neutral I_2 in the process of oxidation.

26. A is correct.

27. D is correct.

28. D is correct.

Oxidation of Mg (*s*): $Mg\ (s) \rightleftarrows Mg^{2+} + 2e^-$ $\quad E° = 2.37\ V$

Reduction of Pb (*s*): $Pb^{2+} + 2e^- \rightleftarrows Pb\ (s)$ $\quad E° = -1.26\ V$

Net Reaction: $Mg\ (s) + Pb^{2+} \rightleftarrows Pb\ (s) + Mg^{2+}$ $\quad E° = 2.244\ V$

Note: when a reaction reverses, the sign of E° changes.

29. B is correct.

30. D is correct.

31. B is correct.

32. B is correct. Oxidation is defined as an increase in oxidation number and is the loss of electrons.

33. C is correct.

34. D is correct.

35. C is correct. The anode in galvanic cells attracts anions. Anions in solution flow toward the anode while cations flow toward the cathode. Oxidation (i.e. loss of electrons) occurs at the anode. Positive ions are formed while negative ions are consumed at the anode. Therefore, negative ions flow toward the anode to equalize the charge.

36. A is correct. Half reaction:

$Cl_2O_7 \rightarrow HClO$

Balancing half-reaction in acidic conditions:

Step 1: Balance all atoms except for H and O

$Cl_2O_7 \rightarrow 2HClO$

Step 2: To balance oxygen, add H_2O to the side with less oxygen atoms

$Cl_2O_7 \rightarrow 2HClO + 5H_2O$

Step 3: To balance hydrogen, add H^+ to the opposing side of H_2O added in the previous step

$Cl_2O_7 + 12\ H \rightarrow 2HClO + 5H_2O$

Step 4: Balance charges by adding electrons to the side with higher/more positive total charge

Total charge on left side: 12(+1) = +12

Total charge on right side: 0