Electrode potential, E_{Cell}, Nernt equation and ECS

26. Answer: (a) –ve, >1, +ve

Explanation:

- **Spontaneous reaction → ΔG < 0**
- ΔG < 0 → Equilibrium constant K > 1
- ΔG < 0 → E° -Cell > 0

27. Answer: (c) Hg_2Cl_2

Explanation:

- The **calomel electrode** (common reference electrode) uses **Hg_2Cl_2 (mercurous chloride)** with mercury.
- It provides a stable and well-defined electrode potential.

28. (a) In galvanic cell anode always made up of negative electrode.

29. (d) $A|\overset{\text{Anode}}{A^+(a=1)}||\overset{\text{Cathode}}{B^+(a=1)}|B$

$EMF = E_{cathode} - E_{anode} = 0.75 - (0.5)$; $EMF = 0.25\ V$.

30. (d) $E^o = -3.05\ Li^+/Li$ is most negative (minimum) and hence Li has maximum tendency to lose electrons or it is the strongest reducing agent.

31. (b) Brown layer is deposited on iron rod because *Cu* has greater than reduction potential than that of Fe^{2+}.

32. (a) $E^o_{Zn^{++}/Zn} < E^o_{Fe^{++}/Fe}$, so *Zn* will reduce Fe^{++}. Zn cannot reduce Mg^{2+} because $E^o_{Zn^{++}/Zn} > E^o_{Mg^{++}/Mg}$

On similar reason *Mg* and Zn cannot oxidize Fe.

33. (d) For the cell reaction, *Fe* acts as cathode and *Sn* as anode. Hence,

$E^o_{cell} = E^o_{\text{cathode}} - E^o_{\text{anode}}$ = – 0.44 – (– 0.14)= – 0.30*V*

The negative *EMF* suggests that the reaction goes spontaneously in reversed direction.

34. (a) $E^o_{\text{cell}} = E^o_{\text{cathode}} - E^o_{\text{anode}} = 0.34 - (-0.76) = 1.10\ V$.

35. (c) $E^o_{\text{cell}} = E^o_{\text{cathode}} - E_{\text{anode}}$; $E^o_{\text{cell}} = 0.34 - (-2.37)$

$E^o_{\text{cell}} = 2.71\ V$.

36. (d) *Mg* lies above *Cu* in electrochemical series and hence *Cu* electrode acts as cathode

$E^o_{cell} = E^o_{Cu^{++}/Cu} - E^o_{Mg^{++}/Mg}$ $2.70\ V = 0.34 - E^o_{Mg^{++}/Mg}$; $E^o_{Mg^{++}/Mg} = -2.36\ V$.

37. (a) Because H_2 has greater reduction potential so it reduced the Ag^+.

38. **Answer:** (b) More lead and Fe^{2+} ions are formed

Explanation:

- Based on the **reactivity series**, Fe is more reactive than Pb.
- Fe displaces Pb^{2+} from solution:
  $Fe(s) + Pb^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Pb(s)$
- Lead being less reactive does **not displace Fe^{2+}**, so Pb is deposited and Fe^{2+} is formed.

39. (d) $\Delta G^o = -nE^oF$

$Fe^{2+} + 2e^- \rightarrow Fe$ .....(i)

$\Delta G^o = -2 \times F \times (-0.440\ V) = 0.880\ F$

$Fe^{3+} + 3e^- \rightarrow Fe$ .....(ii)

$\Delta G^o = -3 \times F \times (-0.036) = 0.108\ F$

On subtracting equation (i) from (ii)

$Fe^{3+} + e^- \rightarrow Fe^{2+}$

$\Delta G^o = 0.108\ F - 0.880\ F = -0.772\ F$

E^o for the reaction $= -\frac{\Delta G^o}{nF} = -\frac{(-0.772\ F)}{1 \times F} = +0.772\ V$.

40. (d) Reducing power *i.e.* the tendency to lose electrons increases as the reduction potential decreases.

41. (b) Cu^{++} will be reduced and Fe will be oxidized.

$Cu^{++} + Fe \rightarrow Cu + Fe^{++}$.

42. (c) Cell reaction is

$Cu_{(s)} + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$

Two half cell reaction is

$Cu \rightarrow Cu^{2+} + 2e^-$ Oxidation (anode)

$Ag^+ + e^- \rightarrow Ag$ Reduction (cathode)

$E_{Cell} = E_{ox} - E_{Red} = 0.80 - 0.34 = +0.46\ V$

43. (a) EMF = [s.r.p. of cathode–s.r.p of anode]

Where s.r.p. = Standard reduction potential

If EMF is positive then the reaction is spontaneous

For *e.g.* in Galvanic cell

(a) EMF = 1.1 *volt*

(b) Cathode is made of copper

(c) Anode is made of Zinc

EMF = 0.34 –(– 0.76) = 1.1 *volt.*

44. **Answer:** (b) Cl_2

Explanation:

- A more reactive halogen displaces a less reactive halogen from its salt.
- Chlorine (Cl_2) displaces Br^- from KBr:
  $Cl_2 + 2\ KBr \rightarrow 2\ KCl + Br_2$

45. **Answer:** (b) The electrode potential is assumed to be zero

Explanation:

- SHE is the **reference electrode**, and its potential is **defined as 0 V** under standard conditions (1 M H^+, 1 atm H_2, 25° C).

46. (d) H_2 is anode because oxidation takes place.

Cu is cathode because reduction is takes place.

47. (c) $E^o_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$.

48. **Answer:** (b) The standard reduction potential of iron is less than that of copper

Explanation:

- A metal with **lower standard reduction potential** is more easily oxidized.
- Fe is oxidized to Fe^{2+}, displacing Cu^{2+} from its solution.

49. **Answer:** (c) Zn > Cu > Ag

Explanation:

- **Reducing agent strength** increases with ease of losing electrons.
- Zn loses electrons most easily → strongest reducing agent.
- Cu is intermediate, Ag is least.

50. **Answer:** (b) Since (i) is less than (ii), zinc becomes the anode and iron the cathode

Explanation:

- Zinc has **lower reduction potential** → oxidizes first.
- It protects iron by acting as **sacrificial anode**.
- Iron remains unoxidized while zinc slowly dissolves.