

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

26. Answer: (a) $-ve, >1, +ve$

Explanation:

- Spontaneous reaction $\rightarrow \Delta G < 0$
- $\Delta G < 0 \rightarrow$ Equilibrium constant $K > 1$
- $\Delta G < 0 \rightarrow E^\circ_{\text{Cell}} > 0$

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) $A | A^+(a=1) || B^+(a=1) | B$

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

30. (d) $E^\circ = -3.05 \text{ V}$ for 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 reduction potential than that of Fe^{2+} .

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

On similar reason Mg and Zn cannot oxidize Fe .



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

$$E_{cell}^{\circ} = E_{cathode}^{\circ} - E_{anode}^{\circ} = -0.44 - (-0.14) = -0.30V$$

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

34. (a) $E_{cell}^{\circ} = E_{cathode}^{\circ} - E_{anode}^{\circ} = 0.34 - (-0.76) = 1.10 V$.

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

$$E_{cell}^{\circ} = 2.71 V$$

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

$$E_{cell}^{\circ} = E_{Cu^{++}/Cu}^{\circ} - E_{Mg^{++}/Mg}^{\circ} \quad 2.70 V = 0.34 - E_{Mg^{++}/Mg}^{\circ}; \quad E_{Mg^{++}/Mg}^{\circ} = -2.36 V$$

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:
 $Fe(s) + Pb^{2+}(aq) \rightarrow Fe^{2+}(aq) + Pb(s)$
- Lead being less reactive does **not displace Fe^{2+}** , so *Pb* is deposited and Fe^{2+} is formed.

39. (d) $\Delta G^{\circ} = -nE^{\circ}F$

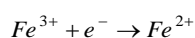


$$\Delta G^{\circ} = -2 \times F \times (-0.440 V) = 0.880 F$$



$$\Delta G^{\circ} = -3 \times F \times (-0.036) = 0.108 F$$

On subtracting equation (i) from (ii)



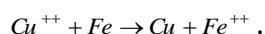
$$\Delta G^{\circ} = 0.108 F - 0.880 F = -0.772 F$$



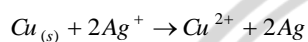
$$E^{\circ} \text{ for the reaction} = -\frac{\Delta G^{\circ}}{nF} = -\frac{(-0.772 F)}{1 \times F} = +0.772 V.$$

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

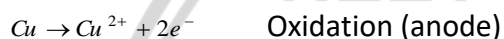
41. (b) Cu^{++} will be reduced and Fe will be oxidized.



42. (c) Cell reaction is



Two half cell reaction is



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

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

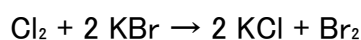
(c) Anode is made of Zinc

$$EMF = 0.34 - (-0.76) = 1.1 \text{ volt.}$$

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 :



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 , $25^\circ C$).

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

47. (c) $E_{cell}^o = E_{cathode} - E_{anode}$.

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 \rightarrow 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** \rightarrow oxidizes first.
- It protects iron by acting as **sacrificial anode**.
- Iron remains unoxidized while zinc slowly dissolves.

