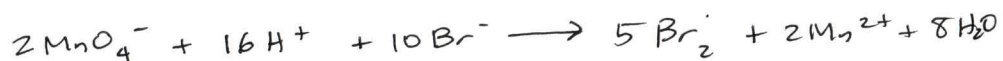
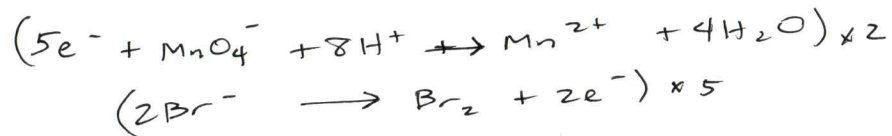
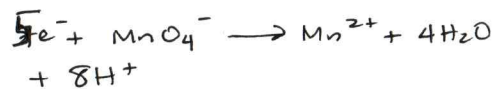
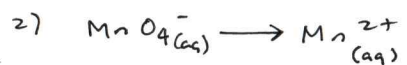
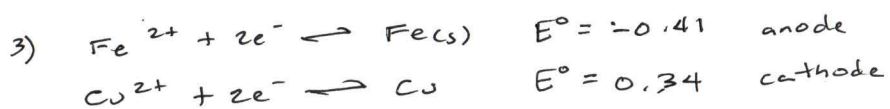


Answers to Chapter 19

1) H_2O is added to balance O in redox



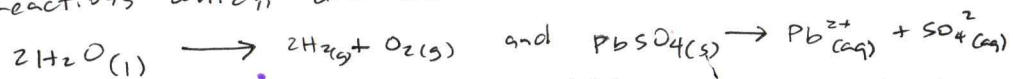
so answer is $16H^+$ reactant side



since electrons move from anode to cathode, there is accumulation of negative charge, forcing Cl^- in the Cu half cell to move to the Fe half cell through the salt bridge

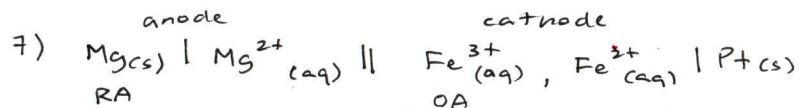
4) Anode is where oxidation happens while cathode is where reduction takes place whether it is a galvanic or electrolytic cell (2 types of electrochemical cells)

5) Looking at each reaction, 2 of them are simple reactions which are not found in reduction table

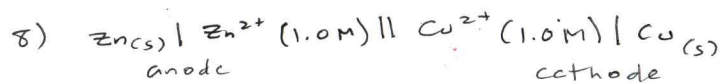


Two are reduction reaction which occur in cathode and one is oxidation reaction (E) that occur at the anode

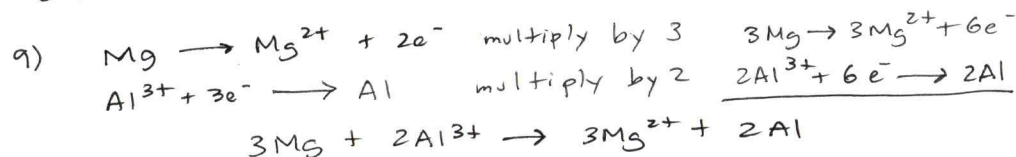
6) Half cell 1 is the anode where $Zn \rightarrow Zn^{2+}$ hence increasing Zn^{2+} resulting to increase Cl^- from salt bridge to maintain electroneutrality



LEORA & GERDA



All statements are true except B since Cu is the cathode.



10. oxidized species + $e^- \rightarrow$ reduced species $E^\circ =$ value more + E° values (tend to be reduced here responsibly for oxidizing the other component)

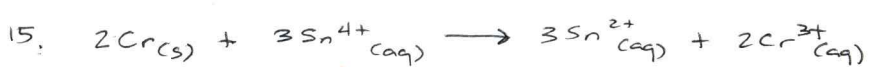
\therefore oxidizing agents are oxidized species with most positive E° values

11. Following the statement in #10, reducing agent is the reduced species with most negative E° values hence it is Mn.

12. The $E^\circ_{As^3+/As}$ = 0.80V while $E^\circ_{Ni^{2+}/Ni}$ = -0.23V. You need a E° that is in between the said E° values hence Pb is the answer.

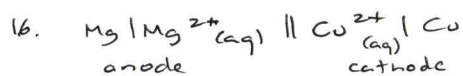
13. with $E^\circ_{cell} = 1.00V$, it means it is a galvanic cell so it happens or reaction is spontaneously.

14. You have to check each choices. For spontaneous rxn to happen, the E° value that is higher must be reduced while the lower E° value will be oxidized. Choice D is the correct one as Mg is oxidized by Cu^{2+} .



Based on the reaction Cr is oxidized while Sn^{4+} is reduced. so

$$E_{cell}^{\circ} = E_{cathode}^{\circ} - E_{anode}^{\circ} = 0.15 - (-0.74) = 0.89V$$



Reaction is spontaneous

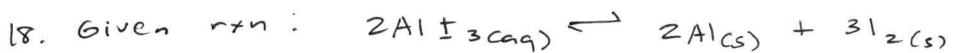
$$E_{cell} = 0.34 - (-2.38) = 2.72V$$

$$17. E_{cell} = 0.34 - (-0.76) = 1.10V$$

$$\Delta G^{\circ} = -nFE^{\circ} = -(2)(96500)(1.10) = -212300 J$$

or -212 kJ

It will become more negative with time since reaction happens spontaneously



Al is ~~oxidized~~ reduced hence happening in cathode while I is oxidized hence happening in anode

$$E_{cell}^{\circ} = E_{cathode}^{\circ} - E_{anode}^{\circ} = -1.66 - 0.54 = -2.20V$$

$$\Delta G^{\circ} = -nFE^{\circ} = -(6 \times 96500 \times -2.20) = 1,273,800 J \text{ or } 1.3 \times 10^6 J$$

6e⁻ involved in the reaction

19. When both ΔH° & ΔS° are positive, temperature is the determining factor to the sign of ΔG° . If temperature is high enough, ΔG° will be negative & spontaneous hence E_{cell}° will increase with increase in temperature.

$$20. \Delta G^{\circ} = \sum \Delta G^{\circ}_{pds} - \sum \Delta G^{\circ}_{rts}$$

$$(-163.0 + -237.0) - (-350.9) = -49.1 kJ$$

$$E_{cell} = \frac{\Delta G}{-nF} = \frac{-49.1 kJ (1000 J/kJ)}{4 \times 96500} = 0.127 V$$

$$21. E = E^{\circ} - \frac{0.0257}{n} \ln Q \quad \ln Q = \frac{[Ag^+]^2}{[Cu^{2+}]}$$

$$= 0.460V - \frac{0.0257}{2} \ln \frac{0.11^2}{0.012}$$

$$= 0.460 - \frac{0.0257}{2} (8.3 \times 10^{-3})$$

$$= 0.460 - 1.1 \times 10^{-4} = 0.460V$$

$$22. E^{\circ} = \frac{0.0257}{n} \ln K$$

$$1.10V = \frac{0.0257}{2} \ln K = 0.01285 \ln K$$

$$\frac{1.10}{0.01285} = \ln K = 85.60$$

$$K = e^{85.60} = 1.50 \times 10^{37} \text{ \& will stay the same with time.}$$

$$23. E = E^{\circ} - \frac{0.0257}{n} \ln Q \quad \ln Q = \frac{[Mn^{2+}]^2}{[MnO_4^-]}$$

$$\ln Q = \frac{[Mn^{2+}]^2 [SO_4^{2-}]^5 [H^+]^4}{[MnO_4^-]}$$

Both 1 & 3 will favor the forward rxns that will increase cell potential.

$$24. E = E^{\circ} - \frac{0.0257}{n} \ln Q$$

$$1.04 = 1.10 - \frac{0.0257}{n} \ln \frac{[Zn^{2+}]}{[Cu^{2+}]}$$

$$1.04 - 1.10 = -\frac{0.0257}{2} \ln \frac{1.0}{[Cu^{2+}]}$$

$$\frac{-0.06(2)}{-0.0257} = \ln \frac{1.0}{[Cu^{2+}]}$$

$$4.7 = \ln \frac{1.0}{[Cu^{2+}]}$$

$$e^{4.7} = \frac{1.0}{[Cu^{2+}]} = 106.6$$

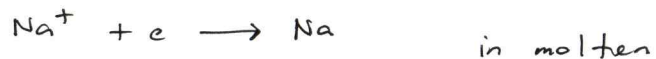
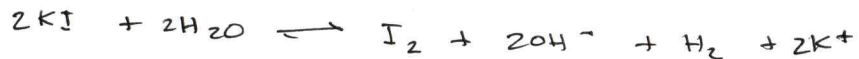
$$[Cu^{2+}] = \frac{1.0}{106.6} = 9.40 \times 10^{-3}$$

25. Cathodic protection is one where an active metal is placed to protect another metal like when iron is attached to a more active metal like Mg. Instead of Fe being oxidized, Mg which is more active, oxidized and protect Fe which act as cathode.

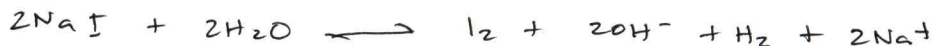


K^+ is reduced to K hence this is the one that occurs in the cathode

27. This is our lab experiment



in aqueous, similar to 27

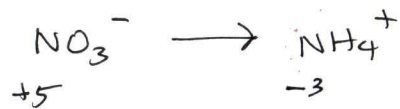


Since OH^- is produced solution becomes basic.

$$29. 1.42 \text{ mol Au} \left(\frac{1 \text{ mol } e^-}{1 \text{ mol Au}} \right) \left(\frac{96500 \text{ C}}{1 \text{ mole } e^-} \right) = 1.37 \times 10^5$$



30 There some error in the question but the main point here is how many electrons involve. oxidation # assignment is needed



so 8 electrons involved.