Ch 21 – Electron Transfer Reactions

1998-40.	**(A) +0.54 V** $E°(cell) = E°_{red}(reduction) - E°_{red}(oxidation)$ $0.79 = 1.33 – x$ $x = 1.33 – 0.79 = +0.54$ volts
1998-41.	**(C) O_2(g)** anode = oxidation (lose electrons "LeO") (+ electrode removes electrons from chemicals) NO_3^- and H_2O are at the + electrode NO_3^- is already oxidized. O in H_2O is oxidized. $2H_2O + 4e^- \rightarrow O_2 + 4H^+$
1998-42.	**(D) Cu^+** best **reducing agent** = most easily **oxidized = smallest E° value** look at **product** in equation with the **smaller E°** value $Cu^{2+} + e^- \rightarrow Cu^+$
1998-43.	**(C) decrease the concentration of Zn^{2+}(aq) ions** anything that drives this reaction **forward** will **increase** the cell's voltage… Le Châtelier's ~~add more Zn(s)~~ [adding a **solid** will not shift the equilibrium] ~~add more Cl⁻ ions~~ [this will increase the **reverse** reaction] decrease Zn^{2+} ions [this is the answer] ~~decrease pressure of Cl₂(g)~~ [this will increase the **reverse** reaction]
1997-43.	**(B) reducing agent** if H_2O_2 is getting **reduced** it is an **oxidizing agent** if H_2O_2 is getting **oxidized**, it is a **reducing agent** $H_2O_2 \rightarrow O_2$ is **oxidation**; O's oxidation number changes from 1- to 0 a **catalyst** would have been written over the arrow. I don't know how you would recognize an **inhibitor**. catalyst and inhibitor are "distractors" …wrong answers
1997-44.	**(D) 4.4 L** Know that electrolysis of water is $2H_2O(l) ® 2H_2(g) + O_2(g)$ if you make 2.2 L of O_2, you make **twice as much** H_2(g)
1997-45.	**(C) increasing the concentration of Co^{2+}(aq)** be careful of convoluted wording… "less negative = more positive" anything that drives this reaction **forward** will **increase** the E° of the half-reaction. increasing or decreasing the **solid Co** will have no effect
1996-43.	**(D) $2 Ag^+ + Pb ® 2 Ag + Pb^{2+}$** when the E° values are written in decreasing order (most + to most -) the **upper-left & lower-right rule** applies… Ag^+ reacts with Pb° and V°, etc. Pb^{2+} reacts with V° but not Ag°
1996-44.	**(a) $Ca(OCl)_2$** You need to look at the half-reaction forming Cl_2(g). The reaction with the **least number of moles of e⁻'s** is the answer. $Ca(OCl)_2$ Cl has oxidation # of +1 in OCl⁻ $2 OCl^- + 2e^- \rightarrow Cl_2 + 2 O^{2-}$ $NaClO_2$ Cl has oxidation # of +3 in ClO_2^- $2 ClO_2^- + 6e^- \rightarrow Cl_2 + 4 O^{2-}$ $KClO_3$ Cl has oxidation # of +5 in ClO_3^- $2 ClO_3^- + 10e^- \rightarrow Cl_2 + 6 O^{2-}$ $Mg(ClO_4)_2$ Cl has oxidation # of +7 in ClO_4^- $2 ClO_4^- + 14e^- \rightarrow Cl_2 + 8 O^{2-}$
1994-46.	**(C) two** The chart is in order of decreasing reduction potentials (E° values) so the upper-left lower-right rule applies. Ag^+ reacts with Ni°, Sn^{2+} reacts with Ni°

1994-48.	**(D) Ti^{4+}** You need to visualize the half-reactions for each metal. The reaction with the most electrons involved will require the greatest length of time. $Ag^+ + e^- \rightarrow Ag°$ $Cu^{2+} + 2e^- \rightarrow Cu°$ $Fe^{3+} + 3e^- \rightarrow Fe°$ $Ti^{4+} + 4e^- \rightarrow Ti°$
1993-67.	**(C) 20** This is a line equation. Remember to begin with "amps x time" $$10\,amps \times 109\,min \times \frac{60\,sec}{1\,min} \times \frac{1\,Couloumb}{1\,amp \bullet sec} \times \frac{1\,mole\ e^-}{96,500\,C} \times \frac{1\,mole\ Co}{2\,mole\ e^-} \times \frac{58.93\,g\,Co}{1\,mole\ Co} = 19.969\,g\,Co$$
1993-66.	**(B) 1.55 V** E°(cell) = E°$_{red}$(reduction) - E°$_{red}$(oxidation) = -0.13 volts –(-1.68 volts) = +1.55 volts Note: The overall reaction is $3Pb^{2+} + 2Al \rightarrow 3Pb + 2Al^{3+}$ (6 moles of electrons are involved) but the coefficients of 2 and 3 are not used for this calculation… not like Hess's Law.
1992-59.	**(A) 0.255 g** This is a line equation. Remember to begin with "amps x time" $$2\,amps \times 1.9\,min \times \frac{60\,sec}{1\,min} \times \frac{1\,Couloumb}{1\,amp \bullet sec} \times \frac{1\,mole\ e^-}{96,500\,C} \times \frac{1\,mole\ Ag}{1\,mole\ e^-} \times \frac{107.9\,g\,Ag}{1\,mole\ Ag} = 0.2549\,g\,Ag$$
1992-60.	**(C) 0.492 V** IF this were a STANDARD cell (everything 1 <u>M</u>) the equation would be: E°(cell) = E°$_{red}$(reduction) - E°$_{red}$(oxidation) -0.250 volts – (-.763 volts) = +0.513 volts (answer D) However… The shortcut cell notation (anode || cathode) shows that the [Zn^{2+}] is only 0.5 <u>M</u>. The overall reaction is: $Zn^{2+} + Ni° \rightarrow Zn° + Ni^{2+}$ Since [Zn^{2+}] is reduced, the cell will run at a little lower voltage… best answer is 0.492 volts. You can also calculate this answer using the **Nernst equation**.