60. Half reaction 1: Br_2 (l) + $2e^-$ = $2Br^-$: E_0 = 1.07 V
Half reaction 2: Al^{3+} + $3e^-$ = Al(s) : E_0 = −1.66 V
Half reaction 1 has a higher E_0, so that is the reduction half-reaction, and the reverse of half reaction 2 is the oxidation half-reaction. Oxidation occurs at the anode, which is written first in line notation:
Al (s) | Al^{3+} (aq, 0.10M) || Br^- (aq, 0.10M) | Br_2 (l) | Pt (s)
For evaluating the cell potential, we take the sum of the half reactions to get E_0 at standard conditions, which is 1.07V + 1.66V = 2.73V. To evaluate at 25°C, we use the Nernst Equation after obtaining the balanced chemical equation of **2Al(s) + 3Br_2 (l) = 2Al^{3+} + 6Br^-**, then getting a reaction quotient $Q=[Al^{3+}]^2\ [Br^-]^6 = [0.10]^8 = 10^{-8}$. $E = E_0 - 0.0591/n \times \log Q = 2.73 - 0.0591/6 \times \log(10^{-8})$ = **2.81 V**, which is galvanic because it is positive.

61. Pt (s) , Pb(s), $PbCl_2$ (s)| Cl^- (aq) || Pb^{2+} (aq) Pb(s) , Pt (s)
The two half reactions are $Pb^{2+}+2e^- = Pb$ and $Pb + 2Cl^- = PbCl_2 + 2e^-$, which are the reduction and oxidation respectively. E_0=-0.13+0.27=0.14. To find the solubility product, we find the ion product assuming equilibrium, which means we can plug values into the Nernst Equation and solve for Q, which in this case is the reciprocal of the ion product since the ions are reactants for the reaction. $0 = 0.14 - 0.0591/2 \times \log(1/K_{sp}) \Rightarrow K_{sp} = 1.85 \times 10^{-5}$

68. Au | Fe^{2+} (aq, 0.1 M), Fe^{3+} (aq, 0.1 M) || Mn^{2+} (0.1M, aq), MnO_4^- (aq, 0.1M) , H^+ (aq 0.1 M) | Au
Half reaction 1: $Fe^{3+} + e^- = Fe^{2+}$, E_0=0.77 V
Half reaction 2: $MnO_4^- + 8H^+ + 5e^- = Mn^{2+} + 4H_2O$, E_0 = 1.51 V.
Half reaction 2 is the reduction, and half reaction 1 is the oxidation.
E_0 = 1.51 V - 0.77 V = 0.74 V
Multiplying the first reaction through by 5 gets both half reactions to an n of 5, so they can be added. This yields a reaction of $5Fe^{2+} + MnO_4^- + 8\ H^+ \longrightarrow 5Fe^{3+} + Mn^{2+} + 4\ H_2O$. Given the concentrations of all are 0.1M, $Q = [0.1]^5\ [0.1]^1/[0.1]^5\ [0.1]^1\ [0.1]^8 = 10^8$. Thus using Nernst Equation,
$\Delta E = E_0 - 0.0591 / n \times \log Q = 0.74\text{ V} - 0.0591/5 \times \log(8) = +0.65\text{ V}$ which is a galvanic cell.

69. The potentials for those options are (respectively) 524mV + 246 mV = 770mV, 39mV + 640 mV = 679 mV, and 438 mV. The last option has the lowest reduction potential, which means it is least likely to be reduced, which means that is least likely to function as an oxidant.

70. a) −0.111V + 0.197 V = 0.086 V
b) 0.023V + 0.197 V - 0.241 V= −0.021 V
c) −0.023 V +0.241 V- 0.197 V = 0.021 V