1) We are given the data that a solution of an unknown salt of rhenium is electrolyzed with 95.0% efficiency, and we get 19.51 g of rhenium. We are also told that we apply 2.75 amps for 2.15 hours. We are asked a) how many faradays are used to deposit the rhenium; and b) what is the charge on the rhenium? We'll approach these in order.

a) The number of coulombs (and thereby, faradays) can be determined from the amperage and the time. \( q = (2.75 \text{ A})(2.15 \text{ h})(3600 \text{ sec/h}) = 21285 \text{ C} \), which, given that 1 Faraday (F) = 96500 C, yields 0.221 F, of which 95%, or 0.210 F is used to deposit the rhenium.

b) Rhenium has an atomic mass of 186.207 g/mol, so 19.51 g of rhenium corresponds to 0.1048 mole rhenium. We used 0.210 Faradays, which is 0.210 moles of electrons to deposit 0.1048 moles of rhenium metal. Taking the ratio of these, we get:

\[
\frac{0.210 \text{ mol}}{0.1048 \text{ mol}} = 2 \text{ mole electrons/mole rhenium}
\]

which leads us to conclude that the charge on rhenium is +2 (i.e. we have Re\(^{2+}\)).

2) We are given the following cell, and asked to get the \( \Delta G^\circ \).

Ce|Ce\(^{3+}\)(aq)||Ni\(^{2+}\)|Ni

The balanced equation for this cell is:

\[
2\text{Ce} + 3\text{Ni}^{2+} \rightarrow 2\text{Ce}^{3+} + 3\text{Ni}
\]

\[\Delta G^\circ = 0.0 \text{kJ/mol} - 46.4 \text{kJ/mol} - 671.95 \text{kJ/mol} + 0.0 \text{kJ/mol} = -1204.7 \text{kJ/mol}\]

This \( \Delta G^\circ \) value, given an \( n = 6 \), leads, via \( \Delta G^\circ = -nFE^\circ_{\text{cell}} \) to:

\[
E^\circ_{\text{cell}} = \frac{-1204700J}{-(6)(96500C)} = 2.08 \text{ V}
\]

Given that \( E^\circ_{\text{Ni/Ni}^{2+}} = -0.23 \), and that \( E^\circ_{\text{cell}} = E^\circ_{\text{cath}} - E^\circ_{\text{an}} \) leads to:

2.08 = -0.23 - (x), and x = \( E^\circ_{\text{Ce}^{3+/Ce}} = -2.31 \text{ V} \).

3) We are given:

\[\text{Ag}^+_{(aq)} + e^- \rightleftharpoons \text{Ag(s)} \quad \quad \quad \quad E^\circ = 0.80 \text{ V}\]

\[\text{Ag(NH}_3\text{)}_2^+_{(aq)} + e^- \rightleftharpoons \text{Ag(s)} + 2 \text{NH}_3(aq) \quad \quad \quad \quad E^\circ = 0.37 \text{ V}\]
and asked to find $K_f$ for the complex ion. If we reverse the second equation, we get a very good cell equation for the formation of the complex i.e.

$$\text{Ag}^{+}(aq) + 2\text{NH}_3(aq) \Leftrightarrow \text{Ag(NH}_3)_2^+ (aq)$$

and an $E^{\circ}_{cell} = 0.80-0.37 = 0.43$ V. Now, we need to realize that the Nernst Equation yields a relation between $E^{\circ}_{cell}$ and the K value: $E^{\circ}_{cell} = \frac{RT}{nF} \ln(K_{th})$, which in this case is $K_f$. Substituting all our values, we get:

$$0.43 = \frac{(8.314 J / K \cdot mol)(298.15 K)}{(1)(96500 C / mol)} \ln(K_{th})$$

which leads us to: $\ln(K_f) = 16.740$, and $K_f = 1.86 \times 10^7$ (which is in quite good agreement with the text value of $1.7 \times 10^7$)

4) We are told that this cell is: $\text{Hg}_2^{2+}(aq) + \text{H}_2(g) \rightarrow 2\text{Hg}(l) + 2\text{H}^+(aq)$ in which the concentration of mercuric ion is controlled by the $K_{sp}$ of $\text{Hg}_2\text{Cl}_2$, which we are asked to find. We need to obtain the $E^{\circ}_{cell}$ and, using the Nernst Equation, obtain the $[\text{Hg}_2^{2+}]$ and from this, the $K_{sp}$ of $\text{Hg}_2\text{Cl}_2$. We know the $E^{\circ}_{\text{S.H.E.}} = 0.0$ V (definition) and we can find $E^{\circ}_{\text{Hg}^+/\text{Hg}} = 0.80$ V. This yields the $E^{\circ}_{cell} = 0.80 - 0.00 = 0.80$ V. We are also given the $E_{cell} = 0.268$ V, and we can determine $Q = \frac{[H^+]^2}{P_{Hg}[\text{Hg}_2^{2+}]} = \frac{[1.0]^2}{1.0[\text{Hg}_2^{2+}]} = \frac{1}{[\text{Hg}_2^{2+}]}$. Lastly, $n=2$.

Substituting, we get: $0.268 V = 0.80 V -(1.285 \times 10^{-2})\ln\left(\frac{1}{[\text{Hg}_2^{2+}]}\right)$ which further leads to:

$$0.532 V = -(1.285 \times 10^{-2}) \ln\left(\frac{1}{[\text{Hg}_2^{2+}]}\right)$$

or

$$\ln\left(\frac{1}{[\text{Hg}_2^{2+}]}\right) = 41.401$$

which gives us (careful!)

$$[\text{Hg}_2^{2+}] = 1.047 \times 10^{-18}$$. With a $[\text{Cl}^-] = 1.0 \text{ M}$, we have, finally, a $K_{sp} = (1.023 \times 10^{-9})(1.0)^2 = 1.047 \times 10^{-18}$.

5) These values are not additive because the cell potential is for a given oxidation/reduction and does not take into account in any overt manner the number of electrons transferred.

How you could use the values is via the $\Delta G^{\circ}_f$, which, as state functions, are additive. Here’s the work:
\[ \text{Ge}^{2+} + 2\text{e}^- \rightarrow \text{Ge} \quad E^\circ = 0.24 \text{ V} \] which gives, via \( \Delta G^\circ_f = -nFE^\circ \),
\[ \Delta G^\circ = -(2)(96500)(0.24) = -46320 \text{ J}. \]

Similarly, \( \text{Ge}^{4+} + 2\text{e}^- \rightarrow \text{Ge}^{2+} \quad E^\circ = 0.00 \text{ V} \) gives \( \Delta G^\circ_f = 0 \text{ J}. \)

Now, we can note that the last equation is merely the sum of the first two, so if we add the \( \Delta G^\circ_f \) values and then solve for \( E^\circ \), we should get the correct answer. The sum of the \( \Delta G^\circ_f \) values is \(-46320 \text{ J} \) which, when used to solve for \( E^\circ \), gives:
\[ E^\circ = \frac{-46320 \text{ J}}{(96500)(-4)} = 0.12 \text{ V} \] which is pretty good.