

CHEM 112 Final Exam (New Material)

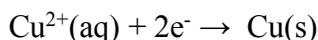
Practice Test – Solutions

1D

Another electrolysis problem. This time we're solving for mass, which almost always means solving for number of moles and then converting to mass.

Before we begin calculations, make sure you understand that the 0.50M concentration is irrelevant for solving this problem, and is just designed to throw you off.

So firstly, we should look at the type of solution we're dealing with to determine the reaction (reduction) that must take place. Copper(II) sulfate means we have a Cu^{2+} ion in solution and so the reaction we are looking for is:



Now we'll use the information given to get the moles of Cu(s) produced:

$$2.3\text{A} = \frac{2.3\text{C}}{\text{s}} \times 1500\text{s} \times \frac{1 \text{ mol e}^{-}}{9.65 \times 10^4 \text{ C}} \times \frac{1 \text{ mol Cu(s)}}{2 \text{ mol e}^{-}} \times \frac{63.5 \text{ g}}{1 \text{ mol}} = 1.14\text{g}$$

2C

The equation used to find E°_{cell} is $E^{\circ}_{\text{cell}} = (0.0592/n) \cdot \log(K_{\text{eq}})$. In this case, there are 2 electrons ($n=2$), which can be determined from the +2 charges. Take note that this is for the entire cell, not just the potential of the oxidation or reduction. This question is specifically asking for the oxidation part of the cell. When using this equation, $E^{\circ}_{\text{cell}} = (0.0592/2) \cdot \log(4.5 \cdot 10^{64}) = 1.91 \text{ V}$. $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{red}} + E^{\circ}_{\text{ox}}$. It's necessary to look up the E°_{red} for Fe, which is " $\text{Fe}^{2+} + 2\text{e}^{-} \rightarrow \text{Fe}$ " and equal to -0.41. This reaction is in the correct direction, since iron is being reduced (Remember that is the equation is reversed $E^{\circ}_{\text{red}} = -E^{\circ}_{\text{ox}}$). Therefore, $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{red}} + E^{\circ}_{\text{ox}}$ turns into $1.91 = -0.41 + E^{\circ}_{\text{ox}}$. E°_{ox} is therefore equal to 2.32 V.

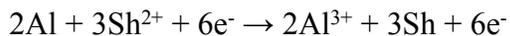
3E

For this problem, it's necessary to understand which electrode is the anode and which is the cathode. In a voltaic cell diagram, the anode will be on the left and the cathode will be on the right. Next, write out the equations associated with oxidation and reduction.

Oxidation (anode): $\text{Al} \rightarrow 3\text{e}^- + \text{Al}^{3+}$

Reduction (cathode): $\text{Sh}^{2+} + 2\text{e}^- \rightarrow \text{Sh}$

The overall equation can be determined. Since the anode and cathode were already determined, the overall equation will have the reactants/products on the same sides of the equation as written above. The aluminum half reaction has to be multiplied by 2 and the Shaiium reaction has to be multiplied by 3 to obtain an equal amount of electrons for the overall reaction (6).



This question is about assessing E_{cell} when not at equilibrium. Therefore, Q will be used instead of K. The equation for this is $E_{\text{cell}} = E^0_{\text{cell}} - (0.0592/n) \cdot \log Q$. Q is determined from the concentrations of Sh^{2+} and Al^{3+} . Remember to raise the concentrations to the power of their respective stoichiometric coefficients. The product, Al^{3+} will be in the numerator and the reactant, Sh^{2+} , will be in the denominator. Only the aqueous components go into the equation for Q and the spectator ions (Cl^-) are not a part of the Q equation.

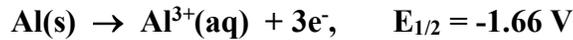
$$\begin{aligned}\log Q &= \log([\text{Al}^{3+}]^2/[\text{Sh}^{2+}]^3) \\ \log Q &= \log([0.5]^2/[3 \cdot 10^{-4}]^3) = 6.4\end{aligned}$$

$$E_{\text{cell}} = E^0_{\text{cell}} - (0.0592/n) \cdot \log Q$$

$$E_{\text{cell}} = 3.07 - (0.0592/6) \cdot 6.4 = 3.01 \text{ V}$$

4B

To get E^0_{cell} we need to find the half reactions that are taking place and look up their standard electrode potential values:



Keep in mind that the top reaction involving Al will appear in the reverse direction as a reduction but we don't have to change the sign of its half-cell potential if we use the equation:

$$E^0_{\text{cell}} = E^0_{\text{cathode}} - E^0_{\text{anode}} = (-0.136) - (-1.66) = 1.524 \text{ V}$$

Remember that reduction occurs at the cathode and oxidation at the anode which is why the half-cell values go where they do, but even if you messed up and got them in the wrong order you'd still get the same number just as a negative instead of a positive. If that happens you can just take the absolute value since most questions will indicate that this is a voltaic or galvanic cell, which means that E^0_{cell} must be positive. However, this question makes no mention as to whether or not the electrochemical cell is indeed voltaic, so technically you have to make sure that you have the right reaction for each half cell before you subtract the values.

5C

The maximum amount of work extracted is the Gibbs free energy. This can be obtained from the equation $\Delta G^0 = -nFE^0_{\text{cell}}$. Look up the standard reduction potentials for the half reactions.



Both standard electrode potentials are for *reduction*. However, taking the negative will give you the oxidation standard potential. In the overall reaction, silver is reduced while zinc is oxidized. Therefore, the sign of the zinc reduction potential must be made positive to obtain the oxidation potential. Using the equation $E^0_{\text{cell}} = E^0_{\text{red}} + E^0_{\text{ox}}$ gives you the E^0_{cell} . **The stoichiometry of the overall reaction does not matter for E^0_{red}** (do not multiply 0.8 by 2). This will only affect the number of electrons transferred (n). E^0_{cell} therefore is equal to 1.56 V. Using $\Delta G^0 = -nFE^0_{\text{cell}}$ with $n=2$ and $F=96485$, $\Delta G^0 = -301,033 \text{ J}$ (-301 kJ). The negative sign makes sense because the positive E^0_{cell} shows that the reaction is spontaneous. So the work that can be extracted is 301 kJ. Multiply this by the efficiency (0.8) to obtain the work for the process. $301 * 0.8 = 240.8 \text{ kJ}$.

6A

Looking at the information provided, it would seem that the best way to solve for ΔG would be to use: $\Delta G^0 = -nFE^0$.

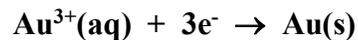
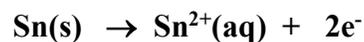
Since F is just a constant, all we need is to find E^0 and n.

We know that the reduction reaction with the most positive half-cell potential is the reaction that actually is the reduction reaction and that the other reaction must be the oxidation reaction. From that we can solve for E^0 by plugging in the values:

$$E^0 = E^0_{\text{cath}} - E^0_{\text{an}} = (+1.50) - (-0.14) = +1.64 \text{ V}$$

The positive value confirms that we have the right order since we're told this is a voltaic cell.

Next we have to calculate n. This requires us to balance the two reactions so that the number of electrons lost equals the number of electrons gained. The two reactions are:



The only way to balance these efficiently is multiply the first reaction by 3 and the second by 2:



Thus we see that number of electrons transferred is actually 6, and that is our n value.

Finally we plug in everything to get:

$$\Delta G^0 = -nFE^0 = -(6)(9.65 \times 10^4)(1.64) = -949560 \text{ J} = -950 \text{ kJ}$$

7E.

The E°_{cell} cannot be found using the standard reduction table because both reactants are not listed. However, K_{eq} is given, so E°_{cell} can be found using the formula

$$E^\circ_{\text{cell}} = (0.059/n) * \log K$$

Plug in for K and 1 for n because only one electron is involved in this redox reaction as shown in the change in oxidation states. E°_{cell} comes out to 0.229V. However, we want E_{cell} because we are NOT at standard conditions.

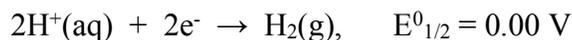
$$E_{\text{cell}} = E_{\text{cell}}(\text{standard}) - (0.059/n)*\log Q$$

$$Q = [\text{Fe}^{3+}][\text{VO}_2^+] / [\text{Fe}^{2+}][\text{VO}_2][\text{H}^+]^2$$

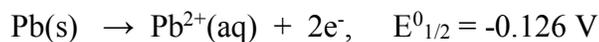
All the values will cancel except for the $[\text{H}^+]$, which you find by taking the inverse log of the pH given. Therefore, $Q = 1 / (1 \times 10^{-6})^2$. Plug this value into the above equation along with the calculated $E_{\text{cell}}(\text{standard})$, to get a final answer of -0.479.

8B

Since we're not told if this is a voltaic cell or an electrolytic cell, it's safer to make sure that we know where each half reaction happens:



This is clearly a reduction and therefore occurs at the cathode.



This is clearly an oxidation and therefore occurs at the anode.

Now we can calculate E°_{cell} :

$$E^\circ = E^\circ_{\text{cath}} - E_{\text{an}} = (0.00) - (-0.126) = +0.126 \text{ V}$$

Because E°_{cell} is positive we now know that this is a voltaic cell and that ΔG° should be negative, which means we can eliminate answers (D) and (E).

Lastly we plug in our values into:

$$\Delta G^\circ = -nFE^\circ_{\text{cell}} = -(2)(9.65 \times 10^4)(0.126) = -24318 \text{ J/mol} = -24 \text{ kJ/mol}$$

9C

Remember that the salt bridge is designed to provide ions (not electrons) to both half cells so as to maintain charge neutrality. Thus (C) is the answer. Keep in mind that these ions are usually non-reactive and certainly do not react with anything, especially at the anode/cathode.

10B.

The best reducing agents can be easily identified on the Standard Reduction Potential Chart in the book or on the exam data sheet. They are the products listed on the bottom of the chart; the lower on the chart the better the reducing agent. Since Zn (s) is the PRODUCT lowest on the chart, then it is the best reducing agent.

11D

The question is really asking us to find out the molar solubility of $\text{Mn}(\text{OH})_2$ in a solution that already has hydroxide ions floating around. This is the common ion effect, and means that the initial concentration of hydroxide has to be taken into account when calculating solubility at equilibrium.

Since $\text{pH} = 11.5$, we know $\text{pOH} = 14 - 11.5 = 2.5$, and so $[\text{OH}^-] = 10^{-2.5} = 0.0032 \text{ M}$

To help calculate solubility we'll use the following ICE table:

		$\text{Mn}(\text{OH})_2(\text{s}) \leftrightarrow \text{Mn}^{2+}(\text{aq}) + 2\text{OH}^-(\text{aq})$	
I	-	0	0.0032
C	-	+x	+2x
E	-	x	0.0032 +2x

We then set up the equation for the solubility product constant, K_{sp} :

$$K_{\text{sp}} = [\text{Mn}^{2+}][\text{OH}^-]^2 = [x][0.0032 + 2x]^2, \text{ where } x \text{ is the solubility}$$

Since K_{sp} is so small we can assume that x is negligible and the equation becomes:

$$K_{\text{sp}} = [x][0.0032]^2$$

$$x = (1.6 \times 10^{-13}) / (0.0032)^2 = 1.6 \times 10^{-8} \text{ M}$$

By examining the ICE table we can see that $[\text{Mn}^{2+}] = x$, and so $[\text{Mn}^{2+}] = 1.6 \times 10^{-8} \text{ M}$

12C

For something to spontaneously oxidize something else, we need to look for the best oxidizing agent which is above (and to the left) the reaction dealing with Br^- .

Remember that the best oxidizing agents are always at the top left of the cell potential chart.

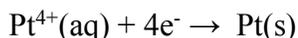
Of the available options, only $\text{Cl}_2(\text{g})$ is above our bromide oxidation.

13A

This is another equation dealing with Faraday's law, except this time we need to solve for current. Remember that the units of current are Amps = C/s, so really we're looking for Coulombs and time (in seconds).

Well, time is easy, since we can just convert 25 minutes to 1500 seconds.

Next we have to consider the reaction that has to take place in order to end up with $\text{Pt}(\text{s})$, so we must first look at the solution it comes from, $\text{PtCl}_4(\text{aq})$, which tells us that the charge on the Pt ion must be +4 and so the reaction (reduction) that must occur is:



Now we use the information provided to determine the charge needed:

$$15\text{g} \times \frac{1 \text{ mol Pt}}{195 \text{ g}} \times \frac{4 \text{ mole}^-}{1 \text{ mol Pt}} \times \frac{9.65 \times 10^4 \text{ C}}{1 \text{ mole}^-} = 29692 \text{ C}$$

All we do now is plug in our values for charge and time to get current:

$$\text{current} = \frac{\text{C}}{\text{s}} = \frac{29692\text{C}}{1500\text{s}} = 19.8 \text{ A}$$

14B

This is about as difficult a question you'll likely see regarding electrolysis. To determine what M is, we'll probably have to figure out its molar mass and then see which answer choice has the same value. Molar mass = (mass / moles), and since we already have the mass, we just need to know how many moles of the metal, M(s), are going to be formed during the electrolysis.

Since the metal oxide is M_2O_3 , we know that the metal must have a +3 charge and will undergo the following reaction: $M^{3+} + 3e^- \rightarrow M(s)$.

Now let's use the information provided:

$$7.5A = \frac{7.5C}{s} \times 1800s \times \frac{1 \text{ mol } e^-}{9.65 \times 10^4 C} \times \frac{1 \text{ mol } M(s)}{3 \text{ mol } e^-} = 0.0466 \text{ mol } M(s)$$

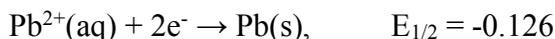
$$\text{So Molar mass} = \frac{1.26 \text{ g}}{0.0466 \text{ mol}} = 27 \text{ g / mol}$$

Of the answer choices available, the only one with a molar mass close to 27g/mol is Al.

15E

Since the cell potential, E^0 , is positive, we know we're dealing with a voltaic cell and thus we can predict which reactions are taking place. This is important because the equation we'll need to solve this question requires us to know which species are reactants and which are products. However, the reduction of Ag^{2+} to Ag(s) is not given on your data sheet and so technically this problem can't be solved for sure without that piece of information (sorry, my bad).

Nonetheless, were the data sheet to have that reaction, the reductions would be:



Since the reduction of Pb^{2+} is more positive than that of Ag^{2+} , we know that Pb^{2+} is reduced and its half reaction stays as is, while the Ag half reaction must be flipped since it has to now be an oxidation. Thus the overall redox reaction is:



Thus we use the following equation to solve for $[\text{Ag}^{2+}]$:

$$E_{cell} = E_{cell}^0 - \frac{0.059V}{n} \log Q$$
$$0.65 = 0.62 - \frac{0.059}{2} \log \left(\frac{[\text{Ag}^{2+}]}{[0.35]} \right)$$

Note that we use $n = 2$ as there are 2 electrons transferred during the balanced reaction.

We can now solve for $[\text{Ag}^{2+}] = 0.038 \text{ M}$

16A.

As electrons move from anode to cathode, a difference in charge builds up at each compartment. The reaction will eventually stop if the cathode becomes too negatively charged. The anions and cations migrate to the anode and cathode, respectively, to neutralize this charge build up and keep the reaction going.

17D.

The standard cell potential of this redox reaction can be found using the table on your data sheet. $E(\text{cathode}) - E(\text{anode})$ gives a positive value. Therefore, the reaction is spontaneous, so $K \gg \gg 1$ and ΔG is negative.

18E.

The last statement is true. This process is known as electroplating.

19D.

Base this question off the formula:

$$E_{\text{cell}} = E_{\text{cell}}(\text{standard}) - (0.059/n) \cdot \log Q$$

If you increase the concentration or pressure of a product in a reaction, it will drive the reaction back into the formation of the reactants. Driving a reaction back to the left will DECREASE the cell potential of a redox reaction. Note that solids are NEVER included in the K_{eq}/Q expressions, so their amounts will not affect E_{cell} . Hydrogen gas is a product and it exists in the Q expression for this reaction, so increasing its value will decrease E_{cell} .

20A.

The reduction process takes place at the cathode. This will involve adding electrons to the cation in the solution to make the metal form.

21C.

When a battery dies, its redox reaction stops because the cell has reached an equilibrium state. At equilibrium, the cell potential is ZERO.

22A.

The metal least likely to corrode is the one that would be least likely to be OXIDIZED. Therefore, the metal HIGHEST on the standard reduction potential list will be the most likely to be REDUCED, and therefore best at avoiding being oxidized relative to the other metals below it on the standard reduction potential list.

23B

The valence electrons in ionic solids predominantly inhabit the anion, leading to an insulation behavior. Covalent-network solid contain covalent bonds, so the electrons are shared between two atoms. This electron localization leads to a semi-conductor behavior. Lastly, metallic solid allow free movement of the valence electrons throughout all the material, giving it high electrical conductivity.

24C

Diatomic oxygen involves a total of 12 valence electrons in the bonding (2×6 from each O). There are 12 electrons in the diagram, so it is the only possibility; all the other answers have more or less total electrons.

25E

Antibonding molecular orbital are represented with a line between the molecular orbitals to designate that they do not overlap to form a bond.

26C

Bond order is calculated by taking the number of electrons in BONDING molecular orbitals, subtracting the number of electrons in ANTIBONDING molecular orbitals, and then dividing that number by two. First, determine the total number of valence electrons in each molecule/compound. F_2 has 14 valence electrons; F_2^- has 15, and F_2^{2-} has 16. When you fill the molecular orbital energy diagram, you add ONE electron at a time to each box, starting with the one at the bottom and working your way up. Once all 14 electrons occupy orbitals for F_2 , then its clear that adding one or two more electrons will only add more electrons in to ANTIBONDING orbitals, which leads to a smaller bond order. F_2 has the highest bond order of 1 ($(8 - 6)/2$), and the others are both lower than 1.