

## CHEM 112 Final Exam (New & Old Material)

### Practice Test – Solutions

1B

Since we're looking for the overall change in concentration over time, we can use the balanced equation to write the following:

$$-\frac{\Delta[N_2]}{\Delta T} = -\frac{1}{3} \frac{\Delta[F_2]}{\Delta T} = \frac{1}{2} \frac{\Delta[NF_3]}{\Delta T}$$

We can then solve for  $\Delta[NF_3]/\Delta T$  to get:

$$\frac{\Delta[NF_3]}{\Delta T} = -\frac{2\Delta[N_2]}{\Delta T} = -\frac{2}{3} \frac{\Delta[F_2]}{\Delta T}$$

Either answer would work, but only one option is given, B.

2D

Since the reaction has only one step, its coefficients can be used as the exponents in the reaction's rate law. Therefore, the rate law is:

$$\text{Rate} = k[A][B]^2.$$

The units of rate are M/s and M for concentration, which leads to  $M^2 s^{-1}$  for the units of k.

3D

By looking at experiments 1 and 2, we can see that tripling the concentration of  $[ClO_2]$  causes the rate to increase by a factor of 9. Thus, the order is 2 with respect to  $[ClO_2]$ .

#### 4A

To get the rate constant we first need to get the order with respect to hydroxide. By looking at experiments 2 and 3, we can see that tripling the hydroxide concentration causes the rate to increase by a factor of 3. Thus the order is 1 with respect to  $[OH]$ . At this point, you can take any experiment's data and plug in all the values that are now known. The only unknown variable is  $k$ , and so we can solve for it:

#1:

$$0.0248 = k[0.060]^2[0.030]^1$$

$$k = 229.6$$

#### 5A

A straightforward question. We can relate half-life to rate constant using the equation:

$$t_{\frac{1}{2}} = \frac{0.693}{k}$$
$$k = 0.00385/s$$

#### 6D

For a compound to be an intermediate, it must be produced during one of the first steps of the mechanism, and then completely used up in a later step. Thus intermediates never appear in the general reaction. As a rule, intermediates are first seen as products, and later as reactants. Only D fits this description.

#### 7E

This is conceptual question dealing with the application of the collision model theory.

## 8C

For a first order reaction, the rate law would look as follows:

$$\text{M/s} = k[\text{M}]$$

Thus if we solve for k, we get:

$$k = 1/\text{s}$$

These are the correct units for every first order reaction.

## 9B

The activation of the reverse reaction is the sum of the  $E_a$  of the forward reaction and the total change in energy with the opposite sign.

## 10C

Considering that after one hour the radioisotope would have experienced a little under 3 half-lives, we can expect the percentage remaining to be close to 12.5%. Remember that after one half-life 50% remains, after two half-lives 25% remains, and after three half-lives 12.5% remains. Since one hour doesn't quite make 3 half-lives, we expect there to be a little more left than 12.5%.

## 11B.

At the normal boiling point, the vapor pressure of propanol (l) equals the atmospheric pressure. Therefore, this is an equilibrium state, so  $\Delta G$  is equal to zero. Use the formula  $\Delta G = \Delta H - T\Delta S$ , plug in zero for G and the S value given to solve for enthalpy. Be sure to convert temperature to Kelvin and Joules into kJ.

12D

Since we are dealing with a weak acid alone in solution, we can skip straight to the equation:  $K_a = (x^2) / (M_i - x)$ . To get pH we'll need to solve for x, the hydronium concentration, and so we are allowed to assume the x in the denominator is negligible and to not include it. Once we get the hydronium concentration we can simply convert to pH by taking its negative log.

13E

The one requirement to be a lewis base is to have a lone pair of electrons. Since option E has absolutely no lone pairs of electrons when you draw its lewis structure, it is least likely to act as a base.

14D

The only candidates even worth considering are options D and E since the other choices are all strong acids. Since the trend for oxoacids involving halogens is that the strength increases as you go up the group, we find that HOBr will be the weakest option.

15A

This is a strong acid / strong base titration. We first need to determine the number of moles of hydronium and of hydroxide:

$$\text{H}_3\text{O}^+: 0.05\text{M} \times 0.13\text{L} = 0.0065 \text{ moles}$$

$$\text{OH}^-: (0.05\text{M} \times 0.055\text{L}) \times 2 \text{ mol OH} / 1 \text{ mol Ba(OH)}_2 = 0.0055 \text{ moles}$$

If we set up an ICE table for this reaction we'll see that all the hydroxide is used up and we are left with 0.0010 moles of  $\text{H}_3\text{O}^+$ .

Thus the pH can now be determined by calculating the final concentration of hydronium and then converting to pH:

$$[\text{H}_3\text{O}^+] = (0.0010 \text{ moles} / 0.185 \text{ L}) = 0.0054 \text{ M}, \text{ and } \text{pH} = -\log[\text{H}_3\text{O}^+] = 2.27$$

16C

For greater entropy, we first look to gas phase options. This leaves A and C. Since they both have the same number of atoms, we can use molecular weight as a quick guide to entropy. The greater weight corresponds to the greater entropy.

17C

Since all options are gas phase, we must first look at the number of atoms. C and D have the fewest atoms, and C has the lowest molecular weight, and thus is the best choice.

18A

Understand that alone in solution, HClO does not break up, and so it will be present in the net reaction. On the other hand, NaOH does break up even before it reacts with anything else, and so we really have  $\text{Na}^+$  and  $\text{OH}^-$  floating around separately. Since only the  $\text{OH}^-$  is needed for the reaction with HClO, the  $\text{Na}^+$  is considered a spectator ion and will not appear in the net reaction. Thus option A is the appropriate answer choice.

19C

By recognizing that this solution is essentially a weak acid, HF, and its conjugate base, which is the  $\text{F}^-$  ion, we can always calculate the pH using the Henderson Hasselbach equation:

$$\text{pH} = -\log(6.8 \times 10^{-4}) + \log([0.4]/[0.5]) = 3.07$$

20A

A common ion exists between HCN, a weak acid, and HI, a strong acid that will completely dissociate. Therefore, to calculate the concentration of  $\text{CN}^-$  from the  $K_a$  equation of HCN, an initial concentration of  $\text{H}^+$  exists from HI (0.10 M). Therefore, you get an equation that looks like:

$$K_a = [x(0.10 + x)]/(0.70 - x)$$

Plug in the value for  $K_a$  given and disregard the  $+ x$  and  $- x$  in the equation because the value of  $x$  is so small that part can be ignored to make for an easier calculation. The value for  $x$  equals  $[\text{CN}^-]$ .

21E

Simply use the Henderson-Hasselbalch equation because it is a buffer system of a weak acid and its conjugate base. The concentration of the acid is 0.15 M and 0.23 M for the base.

$$\text{pH} = \text{pKa} + \log [\text{base}]/[\text{acid}]$$

22A

This question can be answered easily using the Henderson-Hasselbalch equation because it involves a weak acid/base conjugate pair. The  $K_b$  was given, so simply find the  $K_a$  first by taking  $(1.0 \times 10^{-14}/K_b)$  to get  $K_a = 2.5 \times 10^{-9}$ . Now plug in the known values into:

$$\text{pH} = \text{pKa} + \log [\text{base}]/[\text{acid}]$$

Solve for [base]. THEN, multiply that number by 2.0 because the question asks for number of MOLES in a 2.0L solution, not the concentration. Tricky, tricky, tricky...

23A

No matter what temperature the reaction takes place, the product of  $[\text{OH}^-] \times [\text{H}^+]$  will always equal the value of  $K_w$ . Therefore, if you take the square root of the given  $K_w$ , you can get the concentration of  $\text{H}^+$ . Take the  $-\log$  to get the pH.

24C

Most metal oxides are basic; they form metal hydroxides in water. The other statements all follow from the periodic trends outlined in the packet and in notes.

### 25B

The main objective will be to determine the pH (and then the pOH) which will require the number of moles of acetic acid and NaOH present. That calculation is as follows:

$$(0.05 \text{ L})(0.2 \text{ M}) = 0.01 \text{ moles acetic acid}$$

$$(0.035 \text{ L})(0.1 \text{ M}) = 0.0035 \text{ moles NaOH}$$

Since there are less moles of base, it is the limiting reagent in the reaction with the acid and will run out first. Therefore,  $0.01 - 0.0035$  gives  $0.0065$  moles of acidic acid formed. The  $0.0035$  moles of OH<sup>-</sup> from NaOH led to the production of  $0.0035$  moles of acetic acid's conjugate base (the acetate ion). The total volume in this solution is  $0.085 \text{ L}$ , which the number of moles of acetic acid and its conjugate base must each be divided into to get their concentration values. Now, plug those values into the Henderson-Hasselbalch equation to determine pH, subtract that value from 14, and you get your pOH.

### 26C

The buffer zone exists between pH = 4 and pH = 6, and this curve shows a weak acid being titrated with a strong base.

### 27C

Expanding the volume of a gas container allows for more room for the gas molecules to migrate and for entry of more gas molecules, which leads to an increase in the randomness of their motion.

### 28D

The Second Law of Thermodynamics states that the entropy of the universe is always INCREASING.

### 29B

An endothermic reaction that decreases in entropy will never be spontaneous.

### 30C

Remember we can relate equilibrium reactions with free energy using:

$$\Delta G = \Delta G^0 + RT\ln(Q)$$

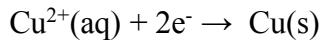
By increasing the pressure of gaseous products we cause Q to increase, which in turn causes the entire right side of the equation to increase, and so ΔG increases as well.

### 31D

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<sup>2+</sup> 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:

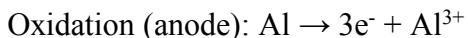
$$2.3A = \frac{2.3C}{s} \times 1500s \times \frac{1 \text{ mol e}^-}{9.65 \times 10^4 C} \times \frac{1 \text{ mol Cu(s)}}{2 \text{ mol e}^-} \times \frac{63.5 \text{ g}}{1 \text{ mol}} = 1.14 \text{ g}$$

### 32C

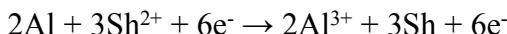
The equation used to find E<sup>o</sup><sub>cell</sub> is E<sup>o</sup><sub>cell</sub>=(0.0592/n)\*log(K<sub>eq</sub>). In this case, there are 2 electrons (n=2), which can be determined from the +2 charges. Take note that this is for the entire cell, not just the potential of the oxidation or reduction. This question is specifically asking for the oxidation part of the cell. When using this equation, E<sup>o</sup><sub>cell</sub> = (0.0592/2)\*log(4.5\*10<sup>64</sup>)=1.91 V. E<sup>o</sup><sub>cell</sub> = E<sup>o</sup><sub>red</sub> + E<sup>o</sup><sub>ox</sub>. It's necessary to look up the E<sup>o</sup><sub>red</sub> for Fe, which is "Fe<sup>2+</sup> + 2e<sup>-</sup> → 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<sup>o</sup><sub>red</sub> = -E<sup>o</sup><sub>ox</sub>). Therefore, E<sup>o</sup><sub>cell</sub> = E<sup>o</sup><sub>red</sub> + E<sup>o</sup><sub>ox</sub> turns into 1.91=-0.41 + E<sup>o</sup><sub>ox</sub>. E<sup>o</sup><sub>ox</sub> is therefore equal to 2.32 V.

33E

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.



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

$$\log Q = \log([\text{Al}^{3+}]^2 / [\text{Sh}^{2+}]^3)$$

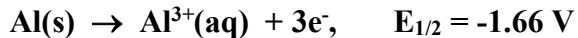
$$\log Q = \log([0.5]^2 / [3 * 10^{-4}]^3) = 6.4$$

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

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

### 34B

To get  $E^0_{\text{cell}}$  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  $E^0_{\text{cell}}$  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.

### 35C

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  $E^0_{\text{cell}}$ . **The stoichiometry of the overall reaction does not matter for  $E^0_{\text{red}}$**  (do not multiply 0.8 by 2). This will only affect the number of electrons transferred ( $n$ ).  $E^0_{\text{cell}}$  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  $E^0_{\text{cell}}$  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.  $301 * 0.8 = 240.8 \text{ kJ}$ .

## 36A

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

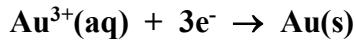
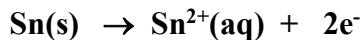
Since F is just a constant, all we need is to find  $E^0$  and n.

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:

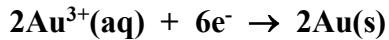
$$E^0 = E_{\text{cath}}^0 - E_{\text{an}}^0 = (+1.50) - (-0.14) = +1.64 \text{ V}$$

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:

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

37E.

The  $E^\circ_{\text{cell}}$  cannot be found using the standard reduction table because both reactants are not listed. However,  $K_{\text{eq}}$  is given, so  $E^\circ_{\text{cell}}$  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.  $E^\circ_{\text{cell}}$  comes out to 0.229V. However, we want  $E_{\text{cell}}$  because we are NOT at standard conditions.

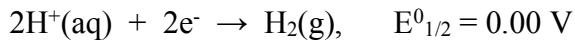
$$E_{\text{cell}} = E_{\text{cell}}(\text{standard}) - (0.059/n) * \log Q$$

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

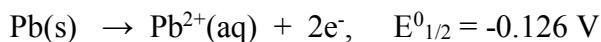
All the values will cancel except for the  $[\text{H}^+]$ , which you find by taking the inverse log of the pH given. Therefore,  $Q = 1 / (1 \times 10^{-6})^2$ . Plug this value into the above equation along with the calculated  $E_{\text{cell}}(\text{standard})$ , to get a final answer of -0.479.

38B

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  $E^\circ_{\text{cell}}$ :

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

Because  $E^\circ_{\text{cell}}$  is positive we now know that this is a voltaic cell and that  $\Delta G^\circ$  should be negative, which means we can eliminate answers (D) and (E).

Lastly we plug in our values into:

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

39C

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.

40B.

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.

41D

The question is really asking us to find out the molar solubility of Mn(OH)<sub>2</sub> 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 pH = 11.5, we know pOH = 14 – 11.5 = 2.5, and so [OH<sup>-</sup>] = 10<sup>-2.5</sup> = 0.0032 M

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

Mn(OH) <sub>2</sub> (s) $\leftrightarrow$ Mn <sup>2+</sup> (aq) + 2OH <sup>-</sup> (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<sub>sp</sub>:

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

Since K<sub>sp</sub> is so small we can assume that x is negligible and the equation becomes:

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

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

By examining the ICE table we can see that [Mn<sup>2+</sup>] = x, and so [Mn<sup>2+</sup>] = 1.6 × 10<sup>-8</sup> M

42C

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  $\text{Br}^-$ .

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

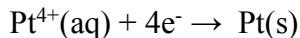
Of the available options, only  $\text{Cl}_2(\text{g})$  is above our bromide oxidation.

43A

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:

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

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

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

#### 44B

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:

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

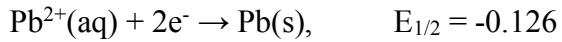
$$\text{So Molar mass} = \frac{1.26 \text{ g}}{0.0466 \text{ mol}} = 27 \text{ g/mol}$$

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

45E

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  $\text{Ag}^{2+}$  to  $\text{Ag(s)}$  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  $\text{Pb}^{2+}$  is more positive than that of  $\text{Ag}^{2+}$ , we know that  $\text{Pb}^{2+}$  is reduced and its half reaction stays as is, while the  $\text{Ag}$  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  $[\text{Ag}^{2+}] = 0.038 \text{ M}$

46A.

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.

47D.

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  $K \gg 1$  and  $\Delta G$  is negative.

48E.

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

49D.

Base this question off the formula:

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

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  $K_{\text{eq}}/Q$  expressions, so their amounts will not affect  $E_{\text{cell}}$ . Hydrogen gas is a product and it exists in the  $Q$  expression for this reaction, so increasing its value will decrease  $E_{\text{cell}}$ .

50A.

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

51C.

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

52A.

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.

53B

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.

54C

Diatom oxygen involves a total of 12 valence electrons in the bonding ( $2 \times 6$  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.

55E

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

56C

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.