Practice MC Unit B (Parts 1 & 2) **Rx's, Sol'n Stoich & Galvanic** Name________________________

- ***This is practice - Do NOT cheat yourself of finding out what you are capable of doing. Be sure you follow the testing conditions outlined below.***
- ***DO NOT USE A CALCULATOR. You may use ONLY your periodic table and formula sheets.***
- ***Work only for 1.5 minutes per question (51 min total). Time yourself. It is important that you practice working for speed. Then when time is up, continue working, with a calculator if you must, but use a different color writing utensil so you will have a sense of how many you were able to complete under the time constraints and without the calculator.***

1. Consider this reaction: $3Au_{(s)} + 8H^+ + 2NO_3^- \rightarrow 3Au^{2+}_{(aq)} + 2NO_{(g)} + 4H_2O_{(L)}$

Mark each of the following statements (about the reaction above) as true or false.

a. T F $Au_{(s)}$ is reduced during the reaction.
b. T F The oxidation state of nitrogen changes from +6 to +2.
c. T F Hydrogen ions are oxidized when they form $H_2O_{(L)}$.
d. T F 8 electrons are lost and 8 electrons are gained during this reaction.

2. Which substance contains the element that is reduced in the following reaction:

$Cr_2O_7^{2-} + 6S_2O_3^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 3S_4O_6^{2-} + 7H_2O$

a. $Cr_2O_7^{2-}$
b. $S_2O_3^{2-}$
c. H^+
d. Cr^{3+}
e. $S_4O_6^{2-}$

3. What is the oxidation number of manganese in the $KMnO_4$

a. +1
b. +2
c. +5
d. +4
e. +7

4. How many electrons appear in the following *half-reaction* when it is balanced.

$Cr_2O_7^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 7H_2O$

a. 1
b. 2
c. 3
d. 4
e. 6

5. How many electrons appear in the following *half-reaction* when it is balanced.

$S_4O_6^{2-} \rightarrow 2S_2O_3^{2-}$

a. 6
b. 2
c. 4
d. 1
e. 3

6. The *half-reaction* occurring at the anode in the balanced reaction shown below is

$3MnO_4^- + 24H^+ + 5Fe \rightarrow 3Mn^{2+} + 5Fe^{3+} + 12H_2O$

a. $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$
b. $MnO_4^- + 5e^- \rightarrow Mn^{2+}$
c. $Fe \rightarrow Fe^{3+} + 3e^-$
d. $Fe + 3e- \rightarrow Fe^{3+}$
e. $MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O + 5e^-$

7. The oxidation *half reaction* occurring in the standard hydrogen electrode is

a. $H_2 \rightarrow 2H^+ + 2e^-$
b. $2H^+ + 2OH^- \rightarrow H_2O$
c. $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$
d. $2H^+ + 2e^- \rightarrow H_2$
e. $2H^+ + Cl_2 \rightarrow 2HCl$

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8. $6H^+ + 5H_2O_2 + 2MnO_4^- \rightarrow 5O_2 + 2Mn^{2+} + 8H_2O$

   According to the balanced equation above, how many moles of the permanganate ion are required to react completely with 25.0 ml of 0.100 M hydrogen peroxide?

   a. 0.000500 mol
   b. 0.00100 mol
   c. 0.00500 mol
   d. 0.00625 mol
   e. 0.0500 mol

9. A 1.0 L sample of an aqueous solution contains 0.10 mol of NaCl and 0.10 mol of $CaCl_2$. What is the minimum number of moles of $AgNO_3$ that must be added to the solution in order to precipitate all of the Cl^- as $AgCl_{(s)}$? (Assume that AgCl is insoluble.)

   a. 0.10 mol
   b. 0.20 mol
   c. 0.30 mol
   d. 0.40 mol
   e. 0.60 mol

10. When 400. milliliter of 0.10-molar sodium chloride is added to 200. milliliters of 0.10-molar aluminum chloride the number of moles of Pb^{2+} that must be added to precipitate out all of the Cl^- would be

    a. 0.020 moles
    b. 0.040 moles
    c. 0.050 moles
    d. 0.060 moles
    e. 0.10 moles

11. A 20.0-milliliter sample of 0.200-molar K_2CO_3 solution is added to 30.0 milliliters of 0.300-molar $Ba(NO_3)_2$ solution. Barium carbonate precipitates. The concentration of barium ion, Ba^{2+}, in solution after reaction is

    a. 0.00500 M
    b. 0.00900 M
    c. 0.0500 M
    d. 0.100 M
    e. 0.00250 M

12. How many moles of solid $Ba(NO_3)_2$ should be added to 300. milliliters of 0.20-molar $Fe(NO_3)_3$ to increase the concentration of the NO_3^- ion to 1.0-molar? (Assume that the volume of the solution remains constant.)

    a. 0.060 mole
    b. 0.12 mole
    c. 0.24 mole
    d. 0.30 mole
    e. 0.40 mole

13. Given the following half reactions:

    $Sn^{4+} + 2e^- \rightarrow Sn^{2+}$ $E^\circ = 0.15$ V
    $Fe^{3+} + 1e^- \rightarrow Fe^{2+}$ $E^\circ = 0.77$ V

    Determine the standard cell potential (E°_{cell}) for the voltaic cell based on the reaction

    $$Sn^{2+} + 2Fe^{3+} \rightarrow 2Fe^{2+} + Sn^{4+}$$

    a. +0.47 V
    b. +0.62 V
    c. +0.92 V
    d. +1.39 V
    e. +1.69 V

14. When this reaction is balanced, the coefficient on the Sn^{2+} is.

    $$Sn^{4+} + Cr \rightarrow Cr^{3+} + Sn^{2+}$$

    a. 1
    b. 2
    c. 3
    d. 4
    e. 6

15. Identify the redox reactions below.

    I. $K_2CrO_4 + BaCl_2 \rightarrow BaCrO_4 + 2KCl$
    II. $Pb^{2+} + 2Br^- \rightarrow PbBr_2$
    III. $Cu + S \rightarrow CuS$

    a. III only
    b. II only
    c. I only
    d. I and III
    e. II and III

16. Which one of the following reactions is a redox reaction?
    a. $NaOH + HCl \rightarrow NaCl + H_2O$
    b. $Pb^{2+} + 2Cl^{-} \rightarrow PbCl_2$
    c. $AgNO_3 + HCl \rightarrow HNO_3 + AgCl$
    d. $2HC_2H_3O_2 + Ca(OH)_2 \rightarrow 2H_2O + Ca(C_2H_3O_2)_2$
    e. None of the above are redox reactions.

17. Which substance contains the element that is oxidized in the following reaction?

    $Fe_2S_3 + 12HNO_3 \rightarrow 2Fe(NO_3)_3 + 3S + 6NO_2 + 6H_2O$

    a. HNO_3
    b. S
    c. NO_2
    d. Fe_2S_3
    e. H_2O

18. What is the coefficient of the bromide ion when the following equation is balanced?

    $$MnO_4^{-} + Br^{-} \rightarrow Mn^{2+} + Br_2$$

    a. 1
    b. 2
    c. 4
    d. 5
    e. 10

19. Consider the half reactions shown below to answer the following question:

    $Fe^{3+} + 1e^{-} \rightarrow Fe^{2+}$ $E^o = 0.77$ V
    $Sn^{4+} + 2e^{-} \rightarrow Sn^{2+}$ $E^o = 0.15$ V
    $Fe^{2+} + 2e^{-} \rightarrow Fe$ $E^o = -0.44$ V
    $Cr^{3+} + 3e^{-} \rightarrow Cr$ $E^o = -0.74$ V

    Which of the following reactions will occur spontaneously as written?

    a. $Sn^{4+} + Fe^{3+} \rightarrow Sn^{2+} + Fe^{2+}$
    b. $3Fe + 2Cr^{3+} \rightarrow 2Cr + 3Fe^{2+}$
    c. $Sn^{2+} + Fe \rightarrow Sn^{4+} + Fe^{2+}$
    d. $3Sn^{4+} + 2Cr \rightarrow 2Cr^{3+} + 3Sn^{2+}$
    e. $3Fe^{2+} \rightarrow Fe + 2Fe^{3+}$

20. In the oxidation-reduction reaction

    $$Sn^{4+} + 2\,Fe^{2+} \rightarrow 2\,Fe^{3+} + Sn^{2+}$$

    a. Sn^{4+} is reduced and Fe^{2+} is oxidized.
    b. Sn^{4+} is oxidized and Fe^{2+} is reduced.
    c. Sn^{4+} is oxidized and Fe^{3+} is reduced.
    d. Fe^{3+} is reduced and Sn^{2+} is oxidized.
    e. Fe^{2+} is oxidized and Fe^{3+} is reduced.

21. Given the standard reduction potentials

    $Cu^{2+} + 2e^{-} \rightarrow Cu$ $E° = +0.34$ Volt
    $Al^{3+} + 3e^{-} \rightarrow Al$ $E° = -1.66$ Volt

    Calculate the standard voltage for the reaction

    $$2Al_{(s)} + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu_{(s)}$$

    a. −1.22 Volts
    b. +1.22 Volts
    c. +2.00 Volts
    d. +2.30 Volts
    e. +4.34 Volts

22. Given the standard electrode (reduction) potentials:

    $Cd^{2+} + 2e^{-} \rightarrow Cd$ $E° = -0.40$ v
    $Ag^{+} + e^{-} \rightarrow Ag$ $E° = +0.80$ v

    What would be the E° for a cadmium-silver cell?

    a. 0.0 V
    b. 0.4 V
    c. 0.5 V
    d. 1.2 V
    e. 2.0 V

23. In the reaction

    $$SO_2 + 2\,H_2S \rightarrow 3\,S + 2\,H_2O$$

    a. sulfur is oxidized and hydrogen is reduced
    b. sulfur is reduced and there is no oxidation
    c. sulfur is oxidized and there is no reduction
    d. sulfur is reduced and hydrogen is oxidized
    e. sulfur is both reduced and oxidized

24. Zinc reacts with dilute acid to produce H_2 and Zn^{2+} but silver does not liberate hydrogen gas from an acid. This information enables one to predict that which of the following reactions has a positive emf ?

a. $H_{2(g)} + Zn^{2+} \rightarrow 2H^+ + Zn$

b. $2\,Ag + Zn^{2+} \rightarrow 2Ag^+ + Zn$

c. $2\,Ag^+ + Zn \rightarrow 2Ag + Zn^{2+}$

d. $2\,Ag + 2H^+ \rightarrow H_{2(g)} + 2\,Ag^+$

e. $2\,Ag^+ + Zn^{2+} \rightarrow 2Ag + Zn$

25. If solid nickel metal were added to separate aqueous solutions each containing 1M concentrations of Ag^+, Cd^{2+}, Sn^{2+}, and Ni^{2+} ions, which metals would **plate out**, based on the given standard reaction potentials?

Standard Reduction Potentials	
Ag^+/Ag	0.799 V
Sn^{2+}/Sn	−0.141 V
Ni^{2+}/Ni	−0.236 V
Cd^{2+}/Cd	−0.400 V

I. Ag

II. Sn

III. Ni

IV. Cd

a. I only

b. IV only

c. I and II only

d. III and IV only

e. I, II, and III

"Plate out" means metal forming on an electrode when metal cations in the solution are reduced on to the electrode.

26. Which of these ions is most able to be oxidized?

Standard Reduction Potentials, E°	
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.77 V
$Cu^{2+} + e^- \rightarrow Cu^+$	+0.15 V

a. Fe^{3+}

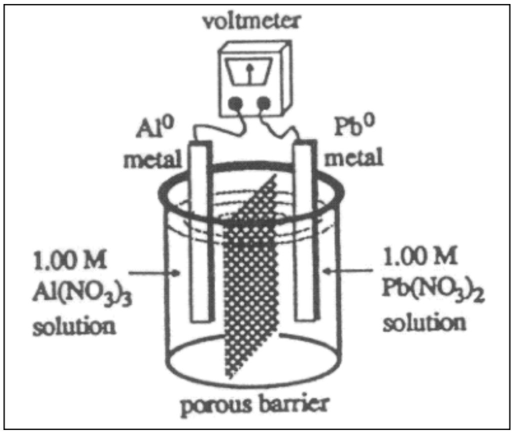
b. Fe^{2+}

c. Cu^{2+}

d. Cu^+

e. None of them can act as reducing agents

27. What voltage will be produced by the electrochemical cell?



Reduction Potentials

$Pb^{2+} + 2e^- \rightarrow Pb$	−0.13 V
$Al^{3+} + 3e^- \rightarrow Al$	−1.68 V

a. +4.78 V

b. +1.55 V

c. +1.81 V

d. +2.97 V

e. +1.29 V

28. Which statements below are true about the voltaic cell in the previous problem.

I. electrons move right to left through the external circuit

II. the lead electrode will gain mass

III. the porous barrier maintains the charge balance in both half-cells

a. I only

b. II only

c. I and II only

d. II and III only

e. I, II, and III

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29. Use these reduction potentials to determine which one of the reactions below is spontaneous.

Reaction	Reduction Potentials, E°
$Ag^+ + e^- \rightarrow Ag$	0.800 V
$Pb^{2+} + 2e^- \rightarrow Pb$	- 0.126 V
$V^{2+} + 2e^- \rightarrow V$	- 1.18 V

I. $V^{2+} + 2\,Ag \rightarrow V + 2\,Ag^+$

II. $V + Pb^{2+} \rightarrow V^{2+} + Pb$

III. $2\,Ag^+ + Pb^{2+} \rightarrow 2\,Ag + Pb$

a. II only

b. III only

c. I and II only

d. II and III only

e. I, II, and III

30. For this reaction, $E°_{cell} = 0.79$ V.

$$6I^- + Cr_2O_7^{2-} + 14H^+ \rightarrow 3I_2 + 2Cr^{3+} + 7H_2O$$

Given that the standard reduction potential for $Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$ is 1.33 V, what is $E°_{red}$ for $I_{2(aq)}$?

a. +0.54 V

b. −0.54 V

c. +0.18 V

d. −0.18 V

e. +2.12 V

31. For this reaction,

$$KClO_4 + H_2O_2 \rightarrow KClO_3 + O_2 + H_2O$$

Choose the true statement from the following list.

a. The Cl oxidation state is reduced form +8 to +6

b. This is not an oxidation-reduction reaction.

c. H_2O_2 caused reduction.

d. Hydrogen is reduced from +2 to +1

e. Potassium is reduced during the reaction

Practice Test Unit B (Parts 1 & 2) (pg 6 of 7) ***ANSWERS***

1. true or false?
   - a. F Gold is oxidized from 0 to 2+.
   - b. F Nitrogen is reduced from +5 to +2.
   - c. F Hydrogen's oxidation state does not change, thus it is neither oxidized nor reduced.
   - d. F Actually it is 6 electrons transferred during this reaction.
2. a LEO says GER or OIL RIG The element that is reduced in this reaction is chromium, and the dichromate ion contains this element.
3. e $(+1) + (Mn) + (4 \times -2) = 0$, thus the Mn must be +7
4. e You can determine this by oxidation numbers, Cr is +6 on left and changes to +3 on the right, but twice thus 6e− tranferred OR you can do it by balancing the charge with the correct number of electrons. As written, the charge on left is $(2-) + (14+) \neq (6+)$, thus you need 6− more on the left, you can get that with 6 electrons.
5. b The easiest way to balance these electrons is by balancing the charge. $2- \neq (2 \times 2-)$, thus adding 2e− to the left side will balance the charge. It is a bit trickier to try to use oxidation numbers since the S on the left has an oxidation number of +2.5, and the oxidation number of +2. This is a change of oxidation state of +½, four times, thus 2 electrons !
6. c Remember "An Ox" and thus we are trying to pull out the oxidation half reaction. Looking at the oxidation numbers, you can see that iron moves from an oxidation state of 0 to +3, which is oxidation.
7. a The standard hydrogen electrode has either oxidation or reduction occurring, depending on what other half reaction it is paired with. $H_2 \rightarrow 2H^+ + 2e^-$ OR $2H^+ + 2e^- \rightarrow H_2$, and the first one listed is the oxidation.
8. b Change to millimoles: 25 ml × 0.1 M = 2.5 mmol, then "stoich" over to permanganate. $2.5mmol \times \frac{2MnO_4^-}{5H_2O_2} = 1mmol$ and of course, 1 mmol is 0.001 mol
9. c This is one of those problems that you need to combine the Cl− ions together since the precipitate reaction is $Ag^+ + Cl^- \rightarrow AgCl$ Calculate mole of NaCl: 1 L × 0.1 M = 0.1 mol of NaCl and NaCl gives us only one Cl^- per NaCl. But for the $CaCl_2$, a "buy 1 get 2" substance, for 0.1 mol of $CaCl_2$ we get 0.2 mole of Cl^-. Thus we have a total of 0.3 mol of Cl^- which will require 0.3 mol of Ag^+ to precipitate out all of the Cl^- from solution.
10. c This problem is very similar to the last, we need to combine the total moles of chloride ions, and then consider how many lead ions would be needed to precipitate out the chloride ions.
    Write a balanced net ionic equation. $Pb^{2+} + 2\ Cl^- \rightarrow PbCl_2$
    Then calculate the total millimol of Cl^-. For NaCl: 400 ml × 0.1 M = 40 mmol NaCl which is 40 mmol of Cl^-
    For $AlCl_3$: 200 ml × 0.1 M = 20 mmol $AlCl_3$ which is 60 mmol of Cl^-
    Add the total Cl^- together to get 100 mmol of Cl^- which when your "stoich" over to the lead ions required, you see that you only need 50 mmol of Pb^{2+} which is of course 0.05 mol of Pb^{2+}.
11. d For this problem, you must realize that barium ions and carbonate ions combine in a 1 to 1 ratio as shown by this net ionic precipitation reaction: $Ba^{2+} + CO_3^{2-} \rightarrow BaCO_3$ The question is indirectly telling you that the carbonate ion limits, because they tell you that barium ions are left in solution. So lets calculate the amount of moles of each ion. For carbonate: 20 ml × 0.2 M = 4 mmol CO_3^{2-} and for barium ion: 30 ml × 0.3 M = 9 mmol Ba^{2+}. Since 4 mmol of CO_3^{2-} requires only 4 mmol of Ba^{2+} to go with it, thus there must be 5 mmol of Ba^{2+} left over, swimming in a TOTAL volume of 50 ml, thus 0.1 M for left over Ba^{2+}
12. a This may be the trickiest problem on the practice test, so buckle your seat belt. This problem is similar to #'s 10 and 11 because of the need to add together the nitrate ions. We want to make a solution that will be 1 M for nitrate ions. Since it is 300 ml of solution, we will need 300 mmole of NO_3^- ions to achive a 1 M solution.
    Calculate the NO_3^- ions from the $Fe(NO_3)_3$ solution: 300 ml × 0.2 M = 60 mmol $Fe(NO_3)_3$, but this is a "buy 1 get 3" substance, thus we actually have 180 mmol of NO_3^- ions. To make up our 300 mmol of NO_3^- ions, we will need 120 mmol more of NO_3^- ions. But wait, before you choose (b) there is one more twist. The source of the nitrate ions you are dumping into this solution is $Ba(NO_3)_2$ and you can see that $Ba(NO_3)_2$ is a "buy 1 get 2" substance, so you only need 60 mmol of $Ba(NO_3)_2$ to provide the 120 mmol of NO_3^- ions required, and of course 60 mmol is 0.06 mol of $Ba(NO_3)_2$. Whew!
13. b Flip the reduction reaction and the add $E^o_{oxidation} + E^o_{reduction} = E^o_{cell}$. (−0.15 V) + (+0.77 V) = +0.62 V
14. c You can balance this by inspection by balancing the charge.
15. a The first reaction listed is an overall precipitation reaction, the second reaction listed is net ionic precipitation reaction, and precipitation reactions are never redox reactions.
16. e (a) and (d) are acid base reactions and (b) and (c) are precipitation. None of these are redox reactions.

17. d Since we are looking for the reducing agent, remember that this means we are looking for the substance that contains the element that is being oxidized. The sulfur changes from −2 on the left to 0 on the right, thus the sulfur is oxidized and the Fe_2S_3 is the reducing agent.

18. e Mn changes from +7 to +2, and Br changes from −1 to 0, twice for 2e–, 2 $Br^- \rightarrow Br_2$ thus we need to multiply the bromine half reaction by 5 (and the manganese half reaction by 2 to balance the electrons transferred.
and 2×(2 $Br^- \rightarrow Br_2$) will give us a coefficient of 10 on the bromide ion.

19. d Since we want to choose a spontaneous reaction, we must produce a + E^o_{cell} . In (a) two reductions are occuring − not possible, and in (c) two oxidations are occuring − not possible. In (b) when you flip the sign of the $E^o_{reduction}$ for Fe^{2+}, you will end up with a negative E^o_{cell}. In (d) when you reverse the Cr^{3+} to the oxidation the
voltage is (+0.15) + (+0.74) = +E^o_{cell}

20. a LEO says GER or OIL RIG

21. c Reverse the Al^{3+} reduction reaction and change the sign. $E^o_{oxidation} + E^o_{reduction} = E^o_{cell}$. (+1.66 V) + (+0.34 V) = +2 V

22. d If we are making a cell, we would want the voltage to be positive. Remember you need a reduction half reaction AND an oxidation half reaction. Flip the half reaction that will allow you to end up with a positive voltage. Thus flip the Cadmium reduction half-reaction, and $E^o_{oxidation} + E^o_{reduction} = E^o_{cell}$. (+0.40 V) + (+0.80 V) = +1.2 V

23. e Sulfur is +4 in SO_2 and −2 in H_2S and it gets both reduced and oxidized to the oxidation state of 0 in S on the product side.

24. c The information in the question tell us that Zn can get oxidized by acid, and is "active" but the Ag is not oxidized by acid and thus is "less active." We are then looking for a reaction which would be spontaneous. This occurs when we allow the more active Zn atoms oxidize to ions, and the less active silver atoms get reduced from Ag^+ to Ag

25. c "Plate-out" is the process of metal sticking to the cathode, meaning which elements listed can reduce and "plate out" while the nickel is getting oxidized. Since the oxidation of nickel has a voltage of +0.236 V, (opposite sign of the reduction listed) only reductions that have a voltage of greater that −0.236 will result in a positve voltage, and thus occur.

26. d Remember, a reducing agent is a substance that gets oxidized, thus in this problem, on ly the Cu^+ and Fe^{2+} can be oxidized. While both oxidations will be negative, the Cu^+ is less negative and therefore better at being oxidized that the Fe^{2+}

27. b Reversing the aluminum half reaction will result in a positve voltage for the cell.
$E^o_{oxidation} + E^o_{reduction} = E^o_{cell}$. (+1.68 V) + (−0.13 V) = +1.55 V

28. d Since the lead is being reduced in the previous problem, the electrons must be flowing toward the lead electrode, thsu left to right. Since the lead is being reduced, lead will "plate" onto the lead electrode, and it will increase in mass. A pourous barrier is the same as a salt bridge.

29. a The first one listed results in a negative voltage $E^o_{oxidation} + E^o_{reduction} = E^o_{cell}$. (−0.8 V) + (−1.18 V) = negative
The second one listed results in a positive voltage $E^o_{oxidation} + E^o_{reduction} = E^o_{cell}$. (+1.18 V) + (−0.126 V) = positive
and the third one listed is two reductions − not possible

30. a $E^o_{oxidation} + E^o_{reduction} = E^o_{cell}$. (x) + (1.33 V) = 0.79 V Then solve for the x which is $E^o_{oxidation}$, but there is a twist, the question asks for the $E^o_{reduction}$ of I_2, so we need to flip the sign of the $E^o_{oxidation}$

31. c The oxygen in H_2O_2 has an oxidation state of −1 and is oxidized to 0 and is thus a reducing agent. We know the oxygen must be oxidized, not reduced to −2 in water because the Cl is reduced from +7 to +5. Some of the oxygens are −2 on the left in the potassium perchlorate, and they remain −2 on the right in water.