Electrochemistry Questions

1. Using the Activity Series on the Useful Information pages of the exam write the chemical formula(s) of the product(s) and balance the following reactions. Identify all products phases as either (g)as, (l)iquid, (s)olid or (aq)ueous. Soluble ionic compounds should be written in the form of their component ions. If no reaction occurs write NR.

a) $Mg_{(s)}$ + $Pb(NO_3)_{2(aq)}$ →

b) $Fe^{2+}_{(aq)}$ + $Cr_{(s)}$ →

2. Write the chemical formula(s) of the product(s) and balance the following spontaneous reactions.
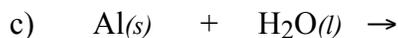
a) $Sn^{2+}_{(aq)}$ + $Ag^+_{(aq)}$ →

b) $Hg_{(l)}$ + $MnO_4^-_{(aq)}$ + $H^+_{(aq)}$ →

c) $Al_{(s)}$ + $H_2O_{(l)}$ →

3. Write the chemical formula(s) of the product(s) and balance the following reactions. Identify all products phases as either (g)as, (l)iquid, (s)olid or (aq)ueous. Soluble ionic compounds should be written in the form of their component ions. If no reaction occurs write NR.
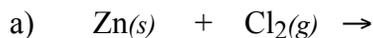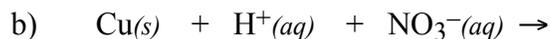
a) $Zn_{(s)}$ + $Ag^+_{(aq)}$ →

b) $HC_3H_5O_{2(aq)}$ + $NaOH_{(aq)}$ →
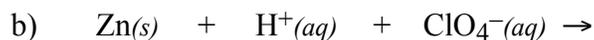
c) $Fe_{(s)}$ + $HCl_{(aq)}$ →

4. Complete and balance the following reactions. Identify all product's phases as either (g)as, (l)iquid, (s)olid or (aq)ueous. If no reaction occurs, write NR.

a) $Zn_{(s)}$ + $Cl_{2(g)}$ →

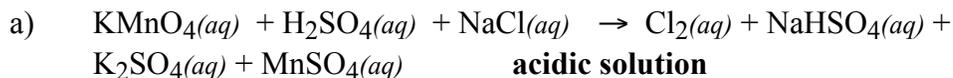b)　　Cu$_{(s)}$　+　H$^+$$_{(aq)}$　+　NO$_3^-$$_{(aq)}$　→

5. Complete and balance the following reactions. Identify all product's phases as either (g)as, (l)iquid, (s)olid or (aq)ueous. If no reaction occurs, write NR.
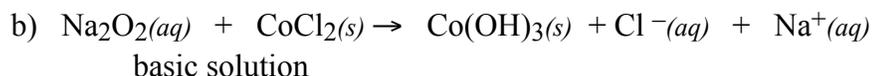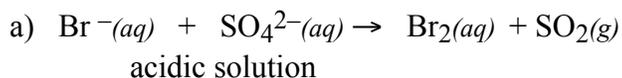
a)　　Mg$_{(s)}$　+　Br$_2$$_{(l)}$　→

b)　　Zn$_{(s)}$　+　H$^+$$_{(aq)}$　+　ClO$_4^-$$_{(aq)}$　→

c)　　Sn$_{(s)}$　+　Fe^{2+}$_{(aq)}$　→

6. Balance the following oxidation-reduction reaction using the half-reaction method.

a)　　KMnO$_4$$_{(aq)}$　+　H$_2$SO$_4$$_{(aq)}$　+　NaCl$_{(aq)}$　→　Cl$_2$$_{(aq)}$　+　NaHSO$_4$$_{(aq)}$　+　K$_2$SO$_4$$_{(aq)}$　+　MnSO$_4$$_{(aq)}$　　　**acidic solution**

　　identify the oxidizing agent _____　　　　　　　identify the reducing agent

_____

7. Balance the following oxidation-reduction reactions using the half-reaction method.

a)　Br$^-$$_{(aq)}$　+　SO$_4^{2-}$$_{(aq)}$　→　Br$_2$$_{(aq)}$　+　SO$_2$$_{(g)}$

　　　acidic solution

b)　Na$_2$O$_2$$_{(aq)}$　+　CoCl$_2$$_{(s)}$　→　Co(OH)$_3$$_{(s)}$　+　Cl$^-$$_{(aq)}$　+　Na$^+$$_{(aq)}$
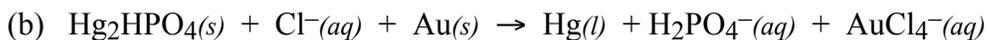
　　　basic solution

Identify the oxidizing agent _____　　　　　　　identify the reducing agent

_____

8. Balance the following oxidation-reduction reactions using the half-reaction method.

(a) $N_2H_4(aq) + CO_3^{2-}(aq) \rightarrow N_2(g) + CO(g)$ **basic solution**

     identify the oxidizing agent _____   identify the reducing agent
_____

(b) $Hg_2HPO_4(s) + Cl^-(aq) + Au(s) \rightarrow Hg(l) + H_2PO_4^-(aq) + AuCl_4^-(aq)$
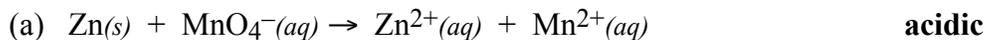   **acidic**
     solution

     identify the oxidizing agent _____   identify the reducing agent
_____

9. Balance the following oxidation-reduction reactions using the half-reaction method.

(a) $Zn(s) + MnO_4^-(aq) \rightarrow Zn^{2+}(aq) + Mn^{2+}(aq)$ **acidic**
     solution

     identify the oxidizing agent _____   identify the reducing agent
_____

(b) $P_4(s) \rightarrow PH_3(g) + H_2PO_2^-(aq)$ **basic solution**
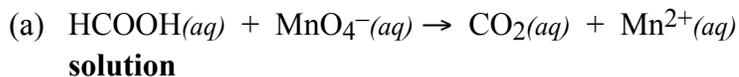
     identify the oxidizing agent _____   identify the reducing agent
_____

10. Balance the following oxidation-reduction reactions using the half-reaction method.

(a) $HCOOH_{(aq)} + MnO_4^-{}_{(aq)} \rightarrow CO_{2(aq)} + Mn^{2+}{}_{(aq)}$ **acidic solution**

    identify the oxidizing agent _____   identify the reducing agent
_____

(b) $HXeO_4^-{}_{(s)} \rightarrow XeO_6^{4-}{}_{(aq)} + Xe_{(g)}$ **basic solution**
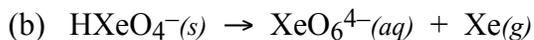
    identify the oxidizing agent _____   identify the reducing agent
_____

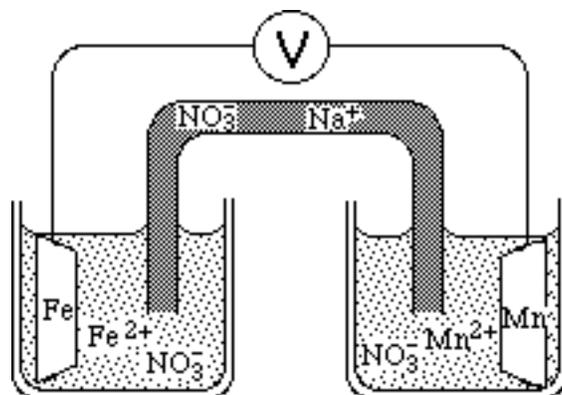11. Given the electrochemical cell shown in Figure I.

Figure I

a) Write the balanced net ionic equation for the spontaneous reaction that occurs as the cell operates.

b) Determine the standard cell voltage, E°.

c) Describe (use a diagram if you like) what is happening at the submicroscopic/atomic level on the surface of the anode in the cell.

d) Indicate the direction of flow of the ions in the salt bridge. (Be sure the anode and cathode compartments are identified.)

12. Short answer.

a) The standard reduction potential for the reduction of RuO_4^- to RuO_4^{2-} is +0.59 volts. Which of the following substances can oxidize RuO_4^{2-} to RuO_4^- under standard condition?

$$H_2SO_4(aq), Pb^{2+}(aq), Cr_2O_7^{2-} \text{ (in acidic solution)}$$

Write the balanced chemical equation to support your choice.

13. A standard chlorine half-cell $Cl_2(g) \mid Cl^-(aq)$ is coupled with a metal(s)|metal ion(aq) half-cell. The voltage of the cell reads 3.73 volts.

a) Identify the metal(s)|metal ion(aq) combination that is involved.

b) Write the balanced chemical equation for the overall oxidation-reduction reaction.

c) Identify the oxidizing agent and the reducing agent in the reaction in part (b).

oxidizing agent _____   reducing agent _____

d) Predict what will happen to the cell potential (more positive, less positive, remain constant) if the [Cl-] is raised and all other species amounts remain at standard state conditions. Without doing any calculations use the Nernst equation to explain your answer. You must explain, do NOT do any calculations.

14. The following materials are available for the construction of an electrochemical cell.
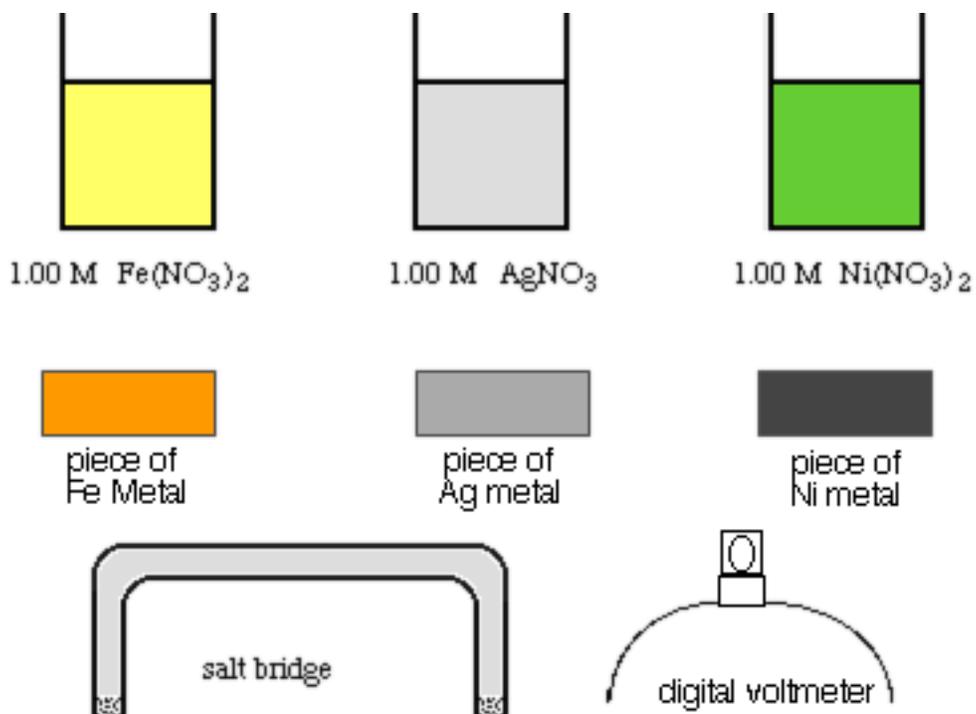
1.00 M Fe(NO$_3$)$_2$ 1.00 M AgNO$_3$ 1.00 M Ni(NO$_3$)$_2$

piece of
Fe Metal   piece of
Ag metal   piece of
Ni metal

salt bridge   digital voltmeter

Figure I

a) Three different electrochemical cells can be constructed from the solutions, metals and materials illustrated above. Write a balanced net-ionic equation for the reaction that occurs in the cell that has the largest positive value of E° (electrochemical potential). (4)

b) Identify the reducing agent in your reaction. (2)

c) Determine the standard cell voltage, E° for the equation written part a. (4)

d) A cell is constructed based on the reaction in part a). Identify the metal used for the anode and indicate the ion(s) in the cathode compartment. Clear indicate your answers. (6)

e) Which of the following compounds is most appropriate to use in a salt bridge, NaOH, CuS, or NaNO₃? Justify your answer with a brief explanation why you selected the particular compound, and why you did not select the other compounds. (6)

f) If the concentration of the solution in the cathode is reduced from 1.0 M to 0.25 M, calculate the new cell potential, E_{cell}. (10)

15. Short Answer

a) In an electrochemical cell Zn is oxidized at the anode.
    i.   Identify the ions in the anode compartment; (4)

    ii.   The salt bridge is a saturated solution of NH₄NO₃. Which ion migrates towards the anode? Explain why that ion migrates towards the anode. (4)

    iii.   An unknown metal ion is reduced at the cathode. The E° for the reaction between Zn and the unknown metal ion is +1.56 v. What is the unknown metal ion? (3)

    iv.   Write the half-reaction that occurs at the cathode.(2)

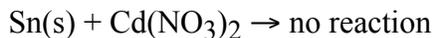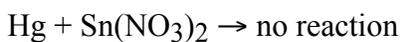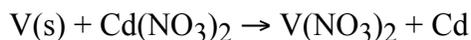   v.   What is the electrode metal for the cathode? (2)

b) In the reaction

$5Fe_{(s)}$ + $2MnO_4^-{}_{(aq)}$ +   16 $H^+{}_{(aq)}$ → $5Fe^{2+}{}_{(aq)}$ + $2Mn^{2+}{}_{(aq)}$ + $8H_2O_{(l)}$
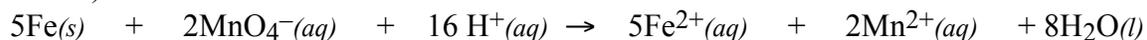
    i.   Calculate E° for the reaction. (4)

ii.   If the pH of the solution is rasied to 3, and assuming all other ions in the solution remain 1 M, will the E_{cell} increase, decrease or remain the same? Briefly, explain your answer. (NOTE: No calculation is required.) (4)

16. Given the following reactions and information about the spontaneous direction, prepare the corresponding activity series showing the four reduction half-reactions ($V^{2+} + 2e^- \rightarrow V$: $Hg^{2+} + 2e^- \rightarrow Hg$: $Sn^{2+} + 2e^- \rightarrow Sn$: $Cd^{2+} + 2e^- \rightarrow Cd$) in the order of most reactive metal ion (least reactive metal) at the top of the list to least reactive metal ion (most reactive metal) at the bottom of the list.

$$V(s) + Cd(NO_3)_2 \rightarrow V(NO_3)_2 + Cd$$

$$Hg + Sn(NO_3)_2 \rightarrow \text{no reaction}$$

$$Sn(s) + Cd(NO_3)_2 \rightarrow \text{no reaction}$$

17. A galvanic cell is constructed using a chromium electrode in a 1.00-molar solution of $Cr(NO_3)_3$ and a copper electrode in a 1.00-molar solution of $Cu(NO_3)_2$. Both solutions are at 25°C.

(a) Write a balanced net ionic equation for the spontaneous reaction that occurs as the cell operates. Identify the oxidizing agent and the reducing agent. (6)

(b) A partial diagram of the cell is shown below.

1.0 M $Cr(NO_3)_3$ 1.0 M $Cu(NO_3)_3$
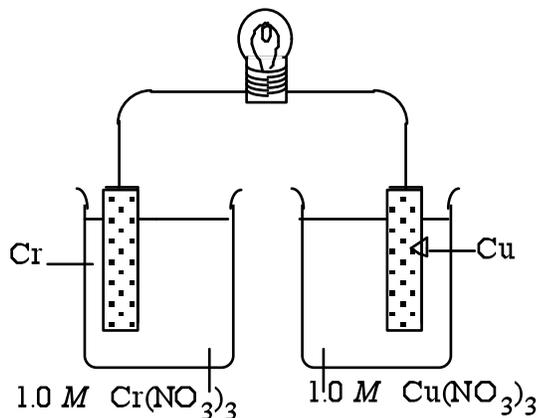
(i) Which metal is the cathode? (2)

(ii) What additional component is necessary to make the cell operate? (3)

(iii) What function does the component in (ii) serve? (4)

(c) How does the potential of this cell change if the concentration of $Cr(NO_3)_3$ is changed to 3.00-molar at 25°C? Explain. (4)

(d) Calculate K and $\Delta G°$ for the reaction you wrote in part a). (14)

e) In another Galvanic cell Cr^{3+} is reduced at the cathode by the metal, X. The anode electrode, composed of the metal X, is immersed in a solution of 1.0 M $X(NO_3)_2$. The measured standard cell potential for the reaction between X and $Cr(NO_3)_3$ is +0.44 v. Calculate the standard reduction potential for X^{2+}. (4)

17.   Which of the following is an oxidation reduction reaction?

A) $Na_2CO_{3(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + 2H_2O_{(l)} + CO_{2(g)}$

B) $H_2SO_{4(aq)} + Ba(OH)_{2(aq)} \rightarrow BaSO_{4(s)} + 2H_2O_{(l)}$

C) $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$

D) $Pb^{2+}_{(aq)} + Ni_{(s)} \rightarrow Pb_{(s)} + Ni^{2+}_{(aq)}$

E) $NaCl_{(s)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$

18.   Mercury batteries, like those used in electric watches, furnish a voltage of 1.35 v.
       If the overall oxidation-reduction reaction taking place in basic solution is

$$Zn_{(s)} + HgO_{(s)} + 2H_2O_{(l)} \rightarrow Zn(OH)_{2(s)} + Hg_{(l)}$$

the anode reaction must be

A) $HgO_{(s)} + 2H_2O_{(l)} + 2e^- \rightarrow Hg_{(l)} + 2OH^-_{(aq)}$

B) $Zn_{(s)} + 2OH^-_{(aq)} \rightarrow Zn(OH)_{2(s)} + 2e^-$

C) $Hg_{(l)} + 2OH^-_{(aq)} \rightarrow HgO_{(s)} + 2H_2O_{(l)} + 2e^-$

D) $Zn(OH)_{2(s)} + 2e^- \rightarrow Zn_{(s)} + 2OH^-_{(aq)}$

19.   Aluminum oxide may be electrolyzed at 1000 ˚C to furnish aluminum metal.
       The cathode reaction is

$$Al^{3+} + 3e^- \rightarrow Al$$
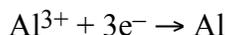
To prepare 5.12 kg of aluminum metal by this method would require

A) 5.49×10^7 C of electricity

B) 1.83×10^7 C of electricity

C) 5.49×10^4 C of electricity

D) 5.49×10^1 C of electricity

20.   The cell potential, E_{cell}, for the reaction

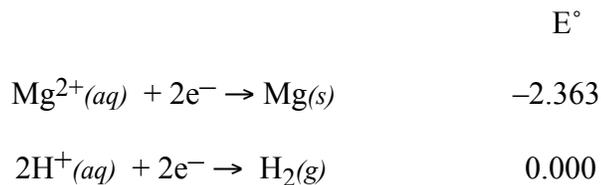$$Zn^{2+}(0.500 \text{ M}) \rightarrow Zn^{2+}(0.0100 \text{ M})$$
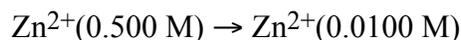
is

A)   +0.050 v

B)   0.00 v

C)   -0.050 v

D)   can not be determine without additional information

21.   Given the following half-reactions

$$E°$$

$$Mg^{2+}\text{(aq)} + 2e^- \rightarrow Mg\text{(s)} \qquad -2.363$$

$$2H^+\text{(aq)} + 2e^- \rightarrow H_2\text{(g)} \qquad 0.000$$

The measured cell potential for the spontaneous reaction

$$Mg\text{(s)} + 2H^+\text{(aq)} \rightarrow Mg^{2+}\text{(aq)} + H_2\text{(g)}$$

under standard conditions is +2.363 v. The pH of the solution in the cathode
compartment must be

A)   0

B)   1

C)   7

D)   10

22. Draw a diagram of the cell in which the following reaction occurs.

$$Cr\text{(s)} + Sn^{2+}\text{(aq)} \rightarrow Cr^{3+}\text{(aq)} + Sn\text{(s)}$$

Label the anode and cathode, the anode and cathode electrode material, the half-
reaction at each electrode, the ions in the anode and cathode compartments and
salt bridge, the direction of electron flow and the direction of ion movement, and
calculate $E°$.

23. A voltaic cell is constructed to carry out the reaction,

$$2Cr^{2+}{}_{(aq)} + Pb^{2+}{}_{(aq)} \rightarrow 2Cr^{3+}{}_{(aq)} + Pb_{(s)}$$
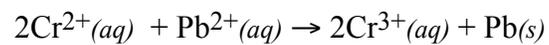
If the $[Cr^{3+}] = 0.0300$ M, $[Pb^{2+}] = 0.150$ M and the $[Cr^{2+}] = 0.250$ M, calculate E_{cell} at 25 ˚C.

24. Calculate K and ΔG° for the reaction in Problem #5.

25. A voltaic cell consists of an anode compartment with a nickel electrode immersed in a $NiSO_4$ solution and a cathode compartment with a copper electrode immersed in a $CuSO_4$ solution. A salt bridge connects the two cells.

a. Write a balanced equation for the cell reaction.

b. A current of 1.50 amps is observed to flow for a period of 2.00 hours. How much charge passes through the circuit during this time? How many moles of electrons is this charge equivalent to?

c. Calculate the change in mass at the copper electrode.

d. Calculate the change in mass at the nickel electrode.

e. At which of the two electrodes does the mass increase?

26. Write the half-reactions which occur at each electrode when the following solutions are electrolyzed between inert electrodes. If more than one half-reaction can occur at the anode or cathode, write all possible half-reactions and briefly explain which half-reaction will occur.
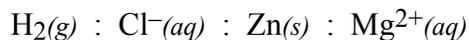
a. $AlCl_3(l)$

   ANODE:

   CATHODE:

b. $CuBr_2(aq)$

   ANODE:

   CATHODE:
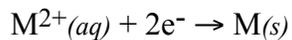
27. Which of the following species is the strongest reducing agent? (Circle your answer.)

$$H_2(g) \quad : \quad Cl^-(aq) \quad : \quad Zn(s) \quad : \quad Mg^{2+}(aq)$$

28. Will $Sn(s)$ displace $Ag^+(aq)$ from solution?

29. Will gold dissolve in $HCl(aq)$?

30. From the following information estimate the $E°$ for

$$M^{2+}(aq) + 2e^- \rightarrow M(s)$$

knowing the metal, M, dissolves in 1 M HCl. It will displace $Fe^{2+}(aq)$, but not $Al^{3+}(aq)$.

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31. Write the half-reactions and determine E° for the electrochemical cells as described below, in part b include the overall chemical equation.

a. $2Al_{(s)} + 3Ni^{2+}_{(aq)} \rightarrow 2Al^{3+}_{(aq)} + 3Ni_{(s)}$
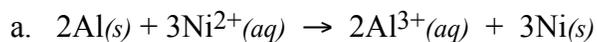
32. A voltaic cell is constructed inwhich the reaction is,

$$2Cr_{(s)} + 3Sn^{4+}_{(aq)} \rightarrow 2Cr^{3+}_{(aq)} + 3Sn^{2+}_{(aq)}$$
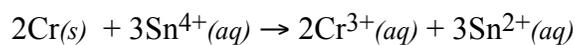
If the $[Cr^{3+}] = 1.00$ M, $[Sn^{4+}] = 0.500$ M and the $[Sn^{2+}] = 0.025$, calculate E_{cell} at 25 °C.

33. Calculate $[Cr^{3+}]$ for the reaction in Problem #5 if the $[Sn^{2+}]$ is 1.00 M and $[Sn^{4+}]$ is 1.00 M and the measured cell potential is +0.994 volts.

34. Write the half-reactions which occur at each electrode when the following solutions are electrolyzed between inert electrodes. If more than one half-reaction can occur at the anode or cathode, write all possible half-reactions and briefly explain which half-reaction will occur.

a. $KBr_{(l)}$

   ANODE:

   CATHODE:

b. $NiCl_{2(aq)}$

   ANODE:

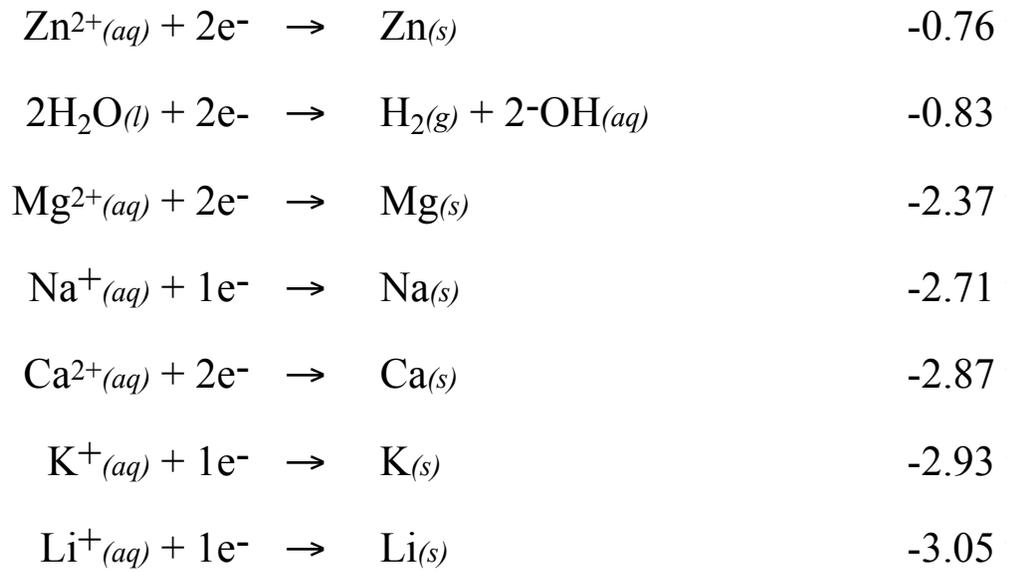   CATHODE:

35. When a dilute solution of H_2SO_4 is electrolyzed, $O_{2(g)}$ is produced at the anode and $H_{2(g)}$ is produced at the cathode.

a. Calculate the coulombs of charge produced when an average current of 8.50 amps passes through the cell for 15.0 minutes.
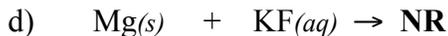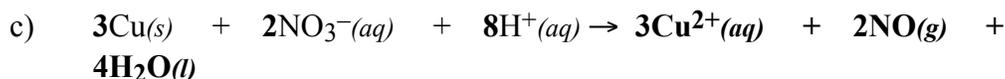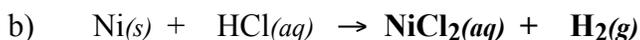
b. Calculate the mass of O_2 and H_2 produced when the cell operates for 15.0 minutes at an average current of 8.50 amps.

c. The heat of formation of $H_2O_{(g)}$ at 25 °C is -242 $\frac{kJ}{mol}$. Calculate the heat liberated when the mixture of H_2 and O_2 in part b. is ignited.

36. Calculate E_{cell} for the reaction in Problem #9a if the $[Zn^{2+}]$ is 1.20 M and $[Cr^{3+}]$ is 1.00 x 10^{-3} M.

37. Using the table of standard reduction potentials, calculate K_{sp} for $PbSO_4$.

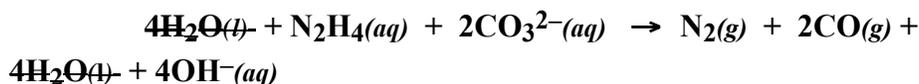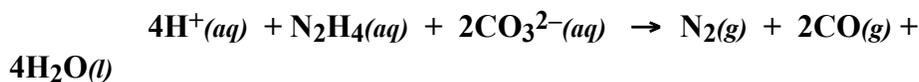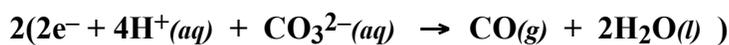Standard Reduction Potentials at 25 ˚C

<u>Half-Reaction</u>
E˚

Half-Reaction		E˚
$F_{2(g)} + 2e^-$ →	$2F^-_{(aq)}$	+2.87
$H_2O_{2(aq)} + 2H^+_{(aq)} + 2e^-$ →	$2H_2O_{(l)}$	+1.77
$MnO_4^-_{(aq)} + 8H^+_{(aq)} + 5e^-$ →	$Mn^{2+}_{(aq)} + 4H_2O_{(l)}$	+1.52
$Cl_{2(g)} + 2e^-$ →	$2Cl^-_{(aq)}$	+1.36
$Cr_2O_7^{2-}_{(aq)} + 14H^+_{(aq)} + 6e^-$ →	$Cr^{3+}_{(aq)} + 7H_2O_{(l)}$	+1.33 v
$O_{2(g)} + 4H^+_{(aq)} + 4e^-$ →	$2H_2O_{(l)}$	+1.23
$Br_{2(l)} + 2e^-$ →	$2Br^-_{(aq)}$	+1.07
$NO_3^-_{(aq)} + 4H^+_{(aq)} + 3e^-$ →	$NO_{(g)} + 2H_2O_{(l)}$	+0.96
$Ag^+_{(aq)} + 1e^-$ →	$Ag_{(s)}$	+0.80
$I_{2(s)} + 2e^-$ →	$2I^-_{(aq)}$	+0.53
$Cu^{2+}_{(aq)} + 2e^-$ →	$Cu_{(s)}$	+0.34
$AgCl_{(s)} + 1e^-$ →	$Ag_{(s)} + Cl^-_{(aq)}$	+0.22
$2H^+_{(aq)} + 2e^-$ →	$H_{2(s)}$	0.00
$Pb^{2+}_{(aq)} + 2e^-$ →	$Pb_{(s)}$	-0.13
$Sn^{2+}_{(aq)} + 2e^-$ →	$Sn_{(s)}$	-0.136
$Ni^{2+}_{(aq)} + 2e^-$ →	$Ni_{(s)}$	-0.25
$Cr^{3+}_{(aq)} + 3e^-$ →	$Cr_{(s)}$	-0.74

$Zn^{2+}{}_{(aq)} + 2e^-$	$\rightarrow$	$Zn_{(s)}$	-0.76
$2H_2O_{(l)} + 2e^-$	$\rightarrow$	$H_{2(g)} + 2{}^-OH_{(aq)}$	-0.83
$Mg^{2+}{}_{(aq)} + 2e^-$	$\rightarrow$	$Mg_{(s)}$	-2.37
$Na^+{}_{(aq)} + 1e^-$	$\rightarrow$	$Na_{(s)}$	-2.71
$Ca^{2+}{}_{(aq)} + 2e^-$	$\rightarrow$	$Ca_{(s)}$	-2.87
$K^+{}_{(aq)} + 1e^-$	$\rightarrow$	$K_{(s)}$	-2.93
$Li^+{}_{(aq)} + 1e^-$	$\rightarrow$	$Li_{(s)}$	-3.05

1. Complete and balance the following reactions. Identify all product's phases as either (g)as, (l)iquid, (s)olid or (aq)ueous. If no reaction occurs, write NR.

a)   Zn$_{(s)}$ +   Cu(NO$_3$)$_{2(aq)}$ →   **Zn(NO$_3$)$_{2(aq)}$**   +   **Cu$_{(s)}$**

b)   Ni$_{(s)}$ +   HCl$_{(aq)}$ →   **NiCl$_{2(aq)}$** +   **H$_2$$_{(g)}$**

c)   **3Cu$_{(s)}$**   +   **2NO$_3$$^-$$_{(aq)}$**   +   **8H$^+$$_{(aq)}$** → **3Cu^{2+}$_{(aq)}$**   +   **2NO$_{(g)}$**   +   **4H$_2O_{(l)}$**

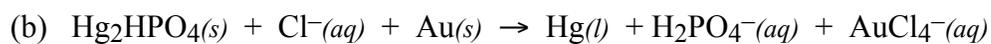d)   Mg$_{(s)}$ +   KF$_{(aq)}$ → **NR**

2. Balance the following oxidation-reduction reactions using the half-reaction method.

(a) N$_2$H$_4$$_{(aq)}$ + CO$_3$$^{2-}$$_{(aq)}$ → N$_2$$_{(g)}$ + CO$_{(g)}$ **basic solution**

$$N_2H_{4(aq)} \ \rightarrow \ N_{2(g)} + 4H^+{}_{(aq)} + 4e^-$$

$$2(2e^- + 4H^+{}_{(aq)} + CO_3{}^{2-}{}_{(aq)} \ \rightarrow \ CO_{(g)} + 2H_2O_{(l)} \ )$$

_____

4H$^+$$_{(aq)}$ + N$_2$H$_4$$_{(aq)}$ + 2CO$_3$$^{2-}$$_{(aq)}$ → N$_2$$_{(g)}$ + 2CO$_{(g)}$ + 4H$_2$O$_{(l)}$

   + 4⁻OH   **+ 4⁻OH**

_____

~~4H$_2$O$_{(l)}$~~ + N$_2$H$_4$$_{(aq)}$ + 2CO$_3$$^{2-}$$_{(aq)}$ → N$_2$$_{(g)}$ + 2CO$_{(g)}$ + ~~4H$_2$O$_{(l)}$~~ + 4OH$^-$$_{(aq)}$
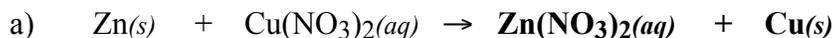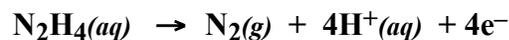
identify the oxidizing agent $CO_3{}^{2-}$ identify the reducing agent $N_2H_4(aq)$

(b) $Hg_2HPO_4(s) + Cl^-(aq) + Au(s) \rightarrow Hg(l) + H_2PO_4{}^-(aq) + AuCl_4{}^-(aq)$

acidic solution

$$2(4Cl^-(aq) + Au(s) \rightarrow AuCl_4{}^-(aq) + 3e^-)$$

$$3(2e^- + H^+ + Hg_2HPO_4(s) \rightarrow 2Hg(l) + H_2PO_4{}^-(aq))$$

_____

$3H^+ + 3Hg_2HPO_4(s) + 8Cl^-(aq) + 2Au(s) \rightarrow 2AuCl_4{}^-(aq) + 6Hg(l) + 3H_2PO_4{}^-(aq)$
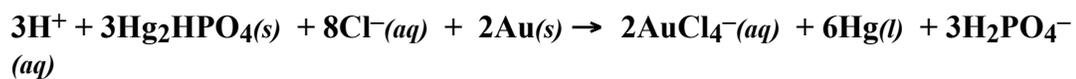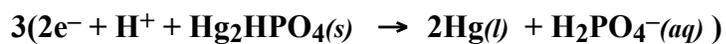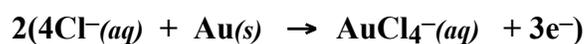
identify the oxidizing agent $Hg_2HPO_4(s)$ identify the reducing agent $Au(s)$

3. Write the half-reactions and determine E° for the electrochemical cells as described below, in part b include the overall chemical equation.

a. $3Pb_{(s)} + SO_4^{2-}_{(aq)} + 8H^+_{(aq)} \rightarrow 3Pb^{2+}_{(aq)} + S_{(s)} + 4H_2O_{(l)}$

$3(Pb \rightarrow Pb^{2+} + 2e^-)$ **+0.126 v**

$6e^- + SO_4^{2-}_{(aq)} + 8H^+_{(aq)} \rightarrow S_{(s)} + 4H_2O_{(l)}$ **+0.357 v**

$3Pb_{(s)} + SO_4^{2-}_{(aq)} + 8H^+_{(aq)} \rightarrow 3Pb^{2+}_{(aq)} + S_{(s)} + 4H_2O_{(l)}$ **+.483 v = E°**

b. $Ni_{(s)} | Ni^{2+}_{(aq)} || Cl_2_{(g)}, Cl^-_{(aq)} | Pt_{(s)}$

$Ni_{(s)} \rightarrow Ni^{2+}_{(aq)} + 2e^-$ **+0.250 v**

$2e^- + Cl_2_{(g)} \rightarrow 2Cl^-_{(aq)}$ **+1.36 v**

$Ni_{(s)} + Cl_2_{(g)} \rightarrow Ni^{2+}_{(aq)} + 2Cl^-_{(aq)}$ **+1.61 v = E°**
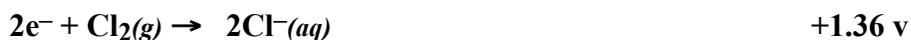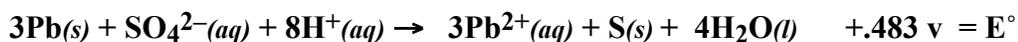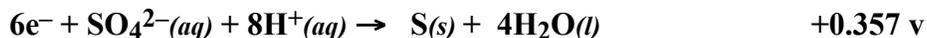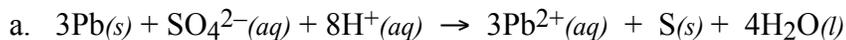
4. Given the following data

Substance	$\Delta H° \left(\dfrac{kJ}{mol}\right)$	$S° \left(\dfrac{J}{mol \cdot K}\right)$	$\Delta G° \left(\dfrac{kJ}{mol}\right)$
$FeO_{(s)}$	-266.3	57.49	-245.12
$C_{(graphite)}$	0	5.74	0
$Fe_{(s)}$	0	27.28	0
$CO_{(g)}$	-110.5	197.6	-137.15

a. Using the above data, calculate $\Delta G°$, $\Delta S°$, and $\Delta H°$ for the reaction,

$$FeO_{(s)} + C_{(graphite)} \rightarrow Fe_{(s)} + CO_{(g)}$$

$$\Delta H°_{rxn} = \Sigma(\Delta H_f°(CO)) - \Sigma(\Delta H_f°(FeO)) = -110.5 \text{ kJ} - (-266.3 \text{ kJ}) = +155.8 \text{ kJ}$$

$$\Delta S°_{rxn} = \Sigma(S°(Fe) + S°(CO)) - \Sigma(S°(FeO) + S°(C))$$

$$= \Sigma(27.28\frac{J}{K} + 197.6\frac{J}{K}) - \Sigma(57.49\frac{J}{K} + 5.74\frac{J}{K}) = +162\frac{J}{K}$$

$$\Delta G°_{rxn} = \Sigma(\Delta G_f°(CO)) - \Sigma(\Delta G_f°(FeO)) = -137.15 \text{ kJ} - (-245.12 \text{ kJ}) = +108 \text{ kJ}$$
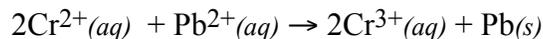
b. Determine at what temperature is the reaction spontaneous.
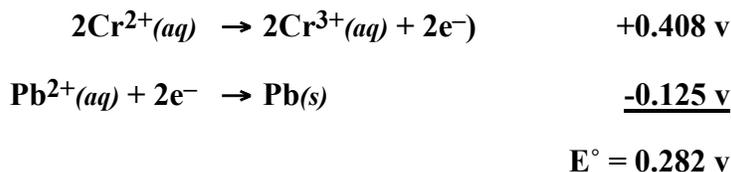
$$\Delta G° = \Delta H° - T\Delta S°$$

$$\Delta G° = 0$$

$$\frac{\Delta H°}{\Delta S°} = T = \frac{155800 \text{ J}}{161.6 \frac{J}{K}} = 964 \text{ K}$$

5. A voltaic cell is constructed to carry out the reaction,

$$2Cr^{2+}(aq) + Pb^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + Pb(s)$$

If the $[Cr^{3+}] = 0.0300$ M, $[Pb^{2+}] = 0.150$ M and the $[Cr^{2+}] = 0.250$ M, calculate E_{cell} at 25 °C.

$$2Cr^{2+}(aq) \rightarrow 2Cr^{3+}(aq) + 2e^-) \qquad +0.408 \text{ v}$$

$$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s) \qquad \underline{-0.125 \text{ v}}$$

$$E° = 0.282 \text{ v}$$

$$E_{cell} = E° - \frac{.059}{n} \log Q$$

$$= E° - \frac{.059}{2} \log \frac{[Cr^{3+}]^2}{[Cr^{2+}]^2[Pb^{2+}]}$$

$$= +0.282 - \frac{.059}{2} \log \frac{[0.03]^2}{[0.25]^2[0.15]}$$

$$= +0.282 - \frac{.059}{2} \log (0.096)$$

$$= +0.282 - \frac{.059}{2} (-1.018)$$

$$= +0.282 + 0.0301$$

$$= +0.312 \text{ v}$$

6. Calculate K and $\Delta G°$ for the reaction in Problem #5.

$$\Delta G° = -nFE° = -2 \cdot 96{,}500 \, \frac{J}{mol \cdot v} \, (+0.282 \, v) = -54.4 \, kJ$$
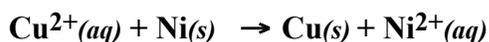
$$E° = \frac{.059}{n} \, \log K$$

$$+0.282 \, v = \frac{.059}{2} \, \log K$$

$$9.543 = \log K$$

$$K = 3.49 \times 10^9$$

7. A voltaic cell consists of an anode compartment with a nickle electrode immersed in a $NiSO_4$ solution and a cathode compartment with a copper electrode immersed in a $CuSO_4$ solution. A salt bridge connects the two cells.

a. Write a balanced equation for the cell reaction.

$$Ni_{(s)} \rightarrow Ni^{2+}_{(aq)} + 2e^-$$

$$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$$

$$Cu^{2+}_{(aq)} + Ni_{(s)} \rightarrow Cu_{(s)} + Ni^{2+}_{(aq)}$$

b. A current of 1.50 amps is observed to flow for a period of 2.00 hours. How much charge passes through the circuit during this time? How many moles of electrons is this charge equivalent to?

$$1.50 \text{ amps} \cdot 2.00 \text{ hours} \left(\frac{60 \text{ min}}{1 \text{ hr}}\right)\left(\frac{60 \text{ s}}{1 \text{ min}}\right)\left(\frac{1 \text{ mol}}{96500 \text{ C}}\right) = 0.112 \text{ mol e}^-$$

c. Calculate the change in mass at the copper electron.

$$0.112 \text{ mol e}^- \left(\frac{1 \text{ mol Cu}}{2 \text{ mol e}^-}\right)\left(\frac{63.5 \text{ g}}{1 \text{ mol}}\right) = 3.56 \text{ g Cu}$$

d. Calculate the change in mass at the nickle electrode.

$$0.112 \text{ mol e}^- \left(\frac{1 \text{ mol Ni}}{2 \text{ mol e}^-}\right)\left(\frac{58.7 \text{ g}}{1 \text{ mol}}\right) = 3.29 \text{ g Cu}$$

e. At which of the two electrodes does the mass increase?

mass increases at the cathode

8. Write the half-reactions which occur at each electrode when the following solutions are electrolyzed between inert electrodes. If more than one half-reaction can occur at the anode or cathode, write all possible half-reactions and briefly explain which half-reaction will occur.
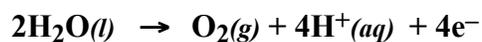
a. $AlCl_{3(l)}$

ANODE:

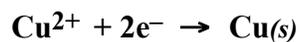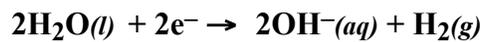$$2Cl^- \rightarrow Cl_2 + 2e^-$$

CATHODE:

$$Al^{3+} + 3e^- \rightarrow Al$$

b. $CuBr_{2(aq)}$

ANODE:

$$2H_2O_{(l)} \rightarrow O_{2(g)} + 4H^+_{(aq)} + 4e^-$$

$$2Br^- \rightarrow Br_{2(l)} + 2e^-$$

CATHODE:

$$Cu^{2+} + 2e^- \rightarrow Cu_{(s)}$$

$$2H_2O(l) + 2e^- \rightarrow 2OH^-(aq) + H_2(g)$$

Short Answer:

9. Which of the following species is the strongest reducing agent? (Circle your answer.)

$$H_2(g) \ : \ Cl^-(aq) \ : \ \boxed{Zn(s)} \ : \ Mg^{2+}(aq)$$
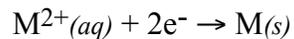
10. Will $Sn(s)$ displace $Ag^+(aq)$ from solution?

YES

11. Will gold dissolve in $HCl(aq)$?

NO

12. From the following information estimate the E° for

$$M^{2+}(aq) + 2e^- \rightarrow M(s)$$

knowing the metal, M, dissolves in 1 M HCl. It will displace $Fe^{2+}(aq)$, but not $Al^{3+}(aq)$.

Mn or Zn