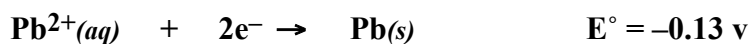
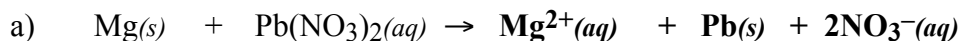
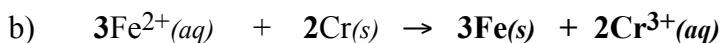


1. Using the Activity Series on the Useful Information pages of the exam write the chemical formula(s) of the product(s) and balance the following reactions. Identify all products phases as either (g)as, (l)iquid, (s)olid or (aq)ueous. Soluble ionic compounds should be written in the form of their component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

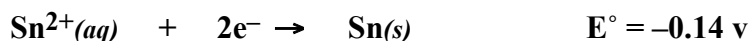
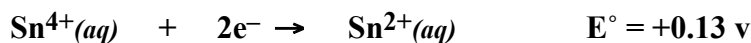
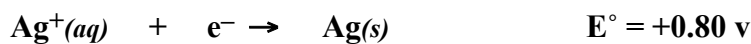
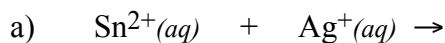


$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = -0.13 \text{ v} - (-2.37 \text{ v}) = +2.24 \text{ v}$$



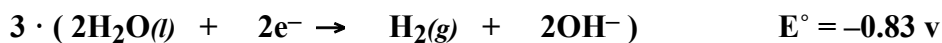
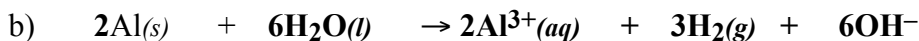
$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = -0.44 \text{ v} - (-0.74 \text{ v}) = +0.30 \text{ v}$$

2. Write the chemical formula(s) of the product(s) and balance the following spontaneous reactions.



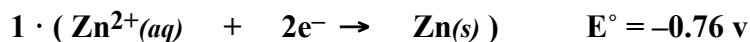
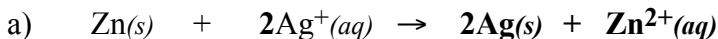
$\text{Ag}^+(aq)$ must be reduced so $\text{Sn}^{2+}(aq)$ must be oxidized, so the reaction of $\text{Sn}^{2+}(aq)$ being reduced can not be used.

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = +0.80 \text{ v} - (+0.13 \text{ v}) = +0.67 \text{ v}$$

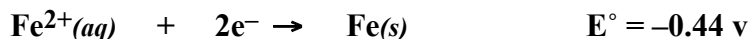
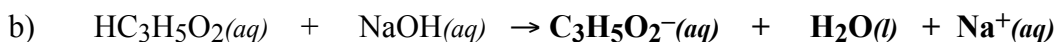


$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = -0.83 \text{ v} - (-1.66 \text{ v}) = +0.83 \text{ v}$$

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.

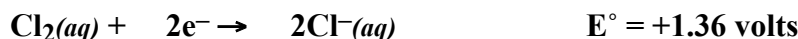
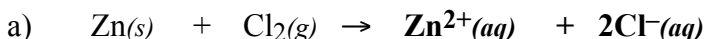


$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = +0.80 \text{ v} - (-0.76 \text{ v}) = +1.56 \text{ v}$$

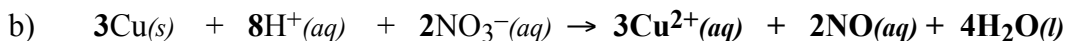


$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = +0.00 \text{ v} - (-0.44 \text{ v}) = +0.44 \text{ v}$$

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.



$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = +1.36 \text{ v} - (-0.76 \text{ v}) = +2.12 \text{ v}$$

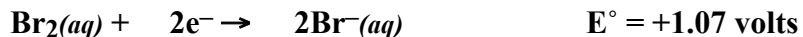
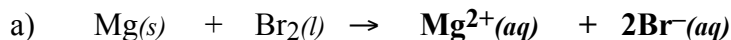


$$2 \cdot (\text{NO}_3^-(aq) + 4\text{H}^+(aq) + 3e^- \rightarrow \text{NO}(aq) + 2\text{H}_2\text{O}(l)) \quad E^\circ = +0.96 \text{ volts}$$

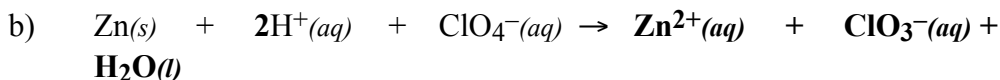
$$3 \cdot (\text{Cu}^{2+}(aq) + 2e^- \rightarrow \text{Cu}(s)) \quad E^\circ = +0.34 \text{ v}$$

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = +0.96 \text{ v} - (+0.34 \text{ v}) = +0.62 \text{ v}$$

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.



$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} = +1.07 \text{ v} - (-2.37 \text{ v}) = +3.44 \text{ v}$$

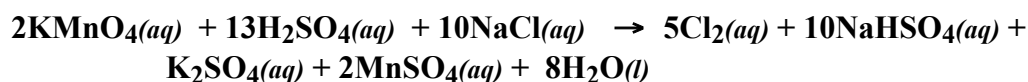
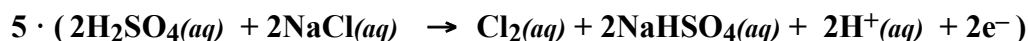
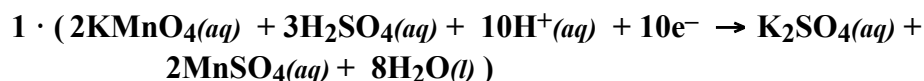
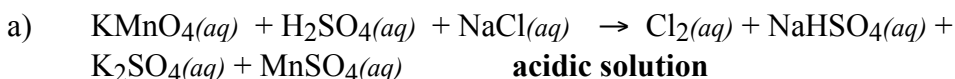


volts



$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}} = +1.19 \text{ v} - (-0.76 \text{ v}) = +1.95 \text{ v}$$

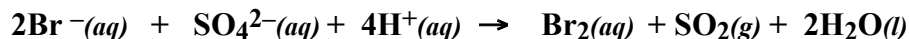
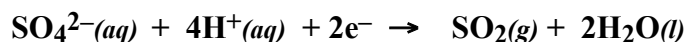
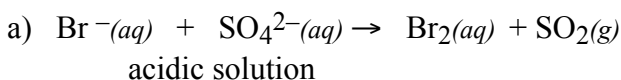
6. Balance the following oxidation-reduction reaction using the half-reaction method.

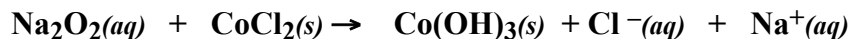
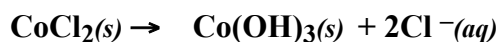
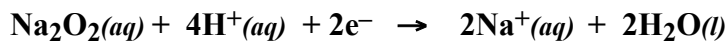
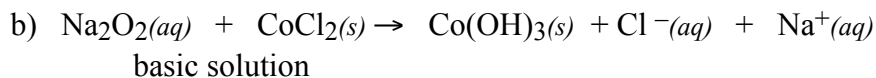


identify the oxidizing agent **NaCl**

identify the reducing agent **KMnO₄**

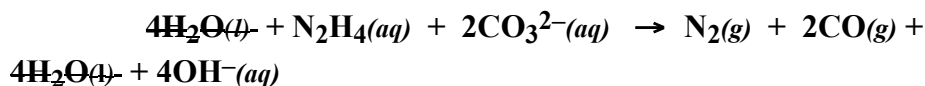
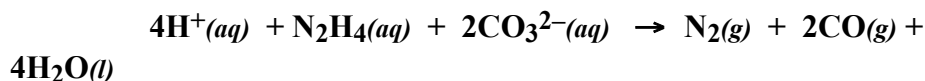
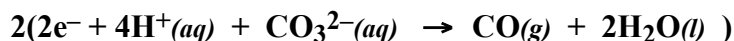
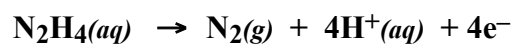
7. Balance the following oxidation-reduction reactions using the half-reaction method.



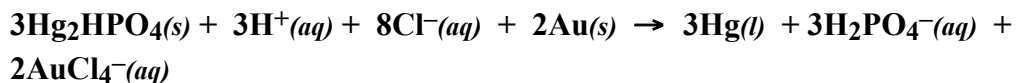
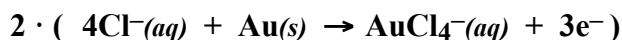
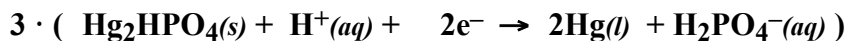
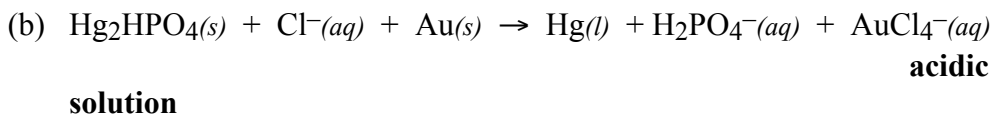


Identify the oxidizing agent _____ identify the reducing agent _____

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

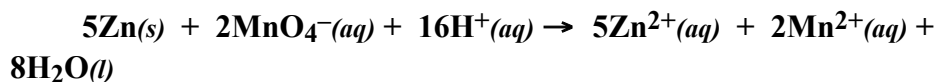
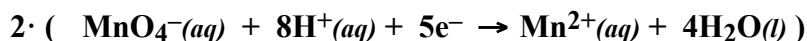
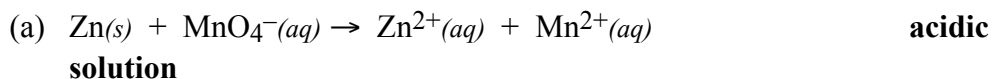


identify the oxidizing agent CO_3^{2-} identify the reducing agent $\text{N}_2\text{H}_4(aq)$

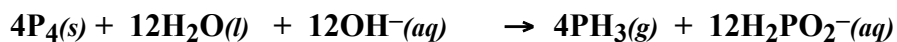
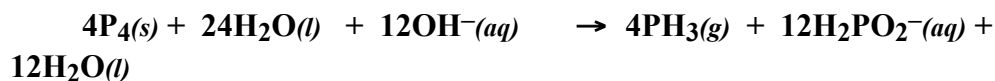
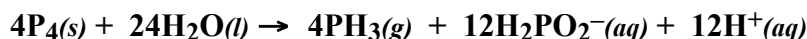
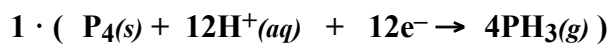


identify the oxidizing agent $\text{Hg}_2\text{HPO}_4(s)$ identify the reducing agent $\text{Au}(s)$

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

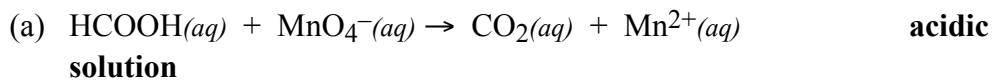


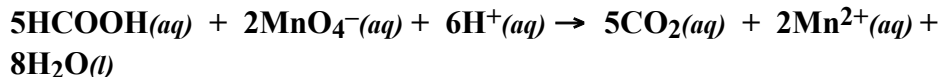
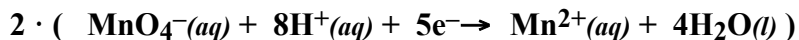
identify the oxidizing agent MnO_4^- identify the reducing agent $\text{Zn}(s)$



identify the oxidizing agent $\text{P}_4(s)$ identify the reducing agent $\text{P}_4(s)$

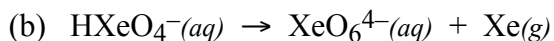
10. Balance the following oxidation-reduction reactions using the half-reaction method.



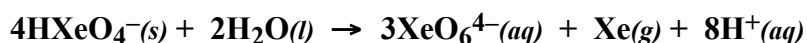
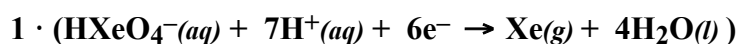
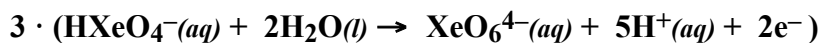


identify the oxidizing agent $\text{MnO}_4^-(aq)$
 $\text{HCOOH}(aq)$

identify the reducing agent



basic solution



identify the oxidizing agent $\text{HXeO}_4^-(aq)$
 $\text{HXeO}_4^-(aq)$

identify the reducing agent

11. Given the electrochemical cell shown in Figure I.

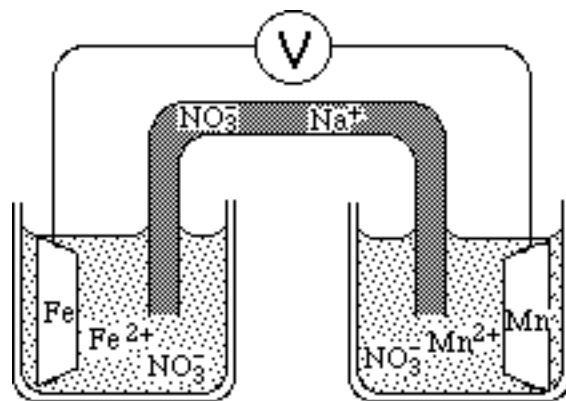
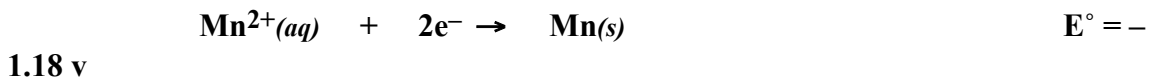
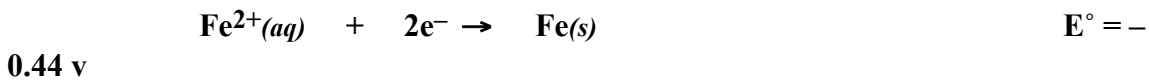


Figure I

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



- b) Determine the standard cell voltage, E° .

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$

$$E^{\circ}_{\text{cell}} = -0.44 \text{ v} - (-1.18 \text{ v}) = +0.74 \text{ v}$$

- 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.

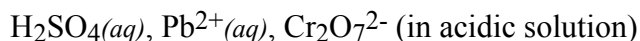
At the anode electrode surface manganese atoms lose two electrons to produce manganese(II) ions in solution.

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

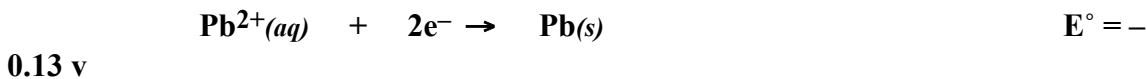
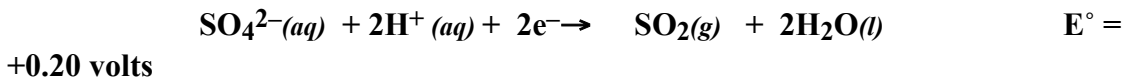
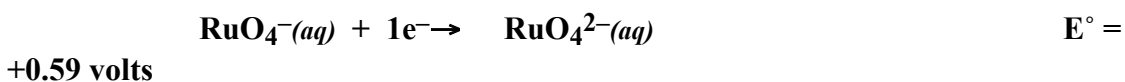
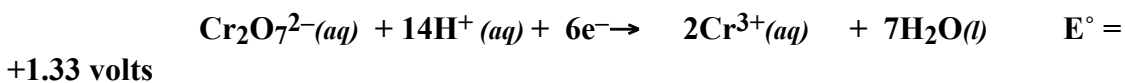
Anions flow towards the anode compartment to balance the positive charged manganese cations that are being formed from the oxidation of the manganese atoms, and cations flow towards the cathode compartment to replace the cations that are removed from solution when iron cations are reduced at the cathode surface to produce iron atoms.

12. Short answer.

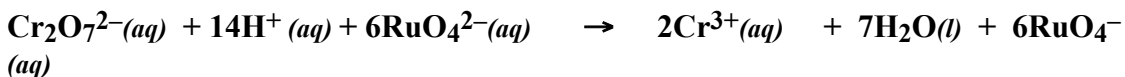
- a) The standard reduction potential for the reduction of RuO_4^- to RuO_4^{2-} is +0.59 volts. Which of the following substances can oxidize RuO_4^{2-} to RuO_4^- under standard condition?



Write the balanced chemical equation to support your choice.



The only substance that can oxidize RuO_4^{2-} to RuO_4^- is $\text{Cr}_2\text{O}_7^{2-}(aq)$ in acidic solution. The balanced reaction is,



$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$E^\circ_{\text{cell}} = +1.33 \text{ v} - (+0.59 \text{ v}) = +0.72 \text{ v}$$

13. A standard chlorine half-cell $\text{Cl}_2(g) | \text{Cl}^-(aq)$ 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.

The standard reduction potential for chlorine is;



To calculate the reduction potential for the metal;

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$+3.73 \text{ v} = +1.36 \text{ v} - E^\circ_{\text{anode}}$$

$$E^{\circ}_{\text{anode}} = +1.36 \text{ v} - (+3.73 \text{ v}) = -2.37 \text{ v}$$

This reduction cell potential corresponds to;



So the metal is $\text{Mg}(\text{s})$ and the ion is $\text{Mg}^{2+}(\text{aq})$.

- 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 $\text{Cl}_2(\text{aq})$ reducing agent $\text{Mg}(\text{s})$

- d) Predict what will happen to the cell potential (more positive, less positive, remain constant) if the $[\text{Cl}^{-}]$ 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.

$$E_{\text{cell}} = E^{\circ} - \frac{0.0592}{n} \log \frac{[\text{Mg}^{2+}][\text{Cl}^{-}]^2}{P_{\text{Cl}_2}} \quad E_{\text{cell}} = +3.73 - \frac{0.0592}{2} \log \frac{[\text{Mg}^{2+}][\text{Cl}^{-}]^2}{P_{\text{Cl}_2}}$$

If the $\text{Cl}^{-}(\text{aq})$ ions concentration is raised, while the concentration of $\text{Mg}^{2+}(\text{aq})$ and the partial pressure of Cl_2 remain constant, then the numerator of the Nernst equation will get larger, and the quotient, $\frac{[\text{Mg}^{2+}][\text{Cl}^{-}]^2}{P_{\text{Cl}_2}}$, will be larger than one. The log of a number greater than

one is positive so the term $\frac{0.0592}{2} \log \frac{[\text{Mg}^{2+}][\text{Cl}^{-}]^2}{P_{\text{Cl}_2}}$ is positive and is

subtracted from the standard cell potential. This make E_{cell} smaller (less positive).

14. The following materials are available for the construction of an electrochemical cell.

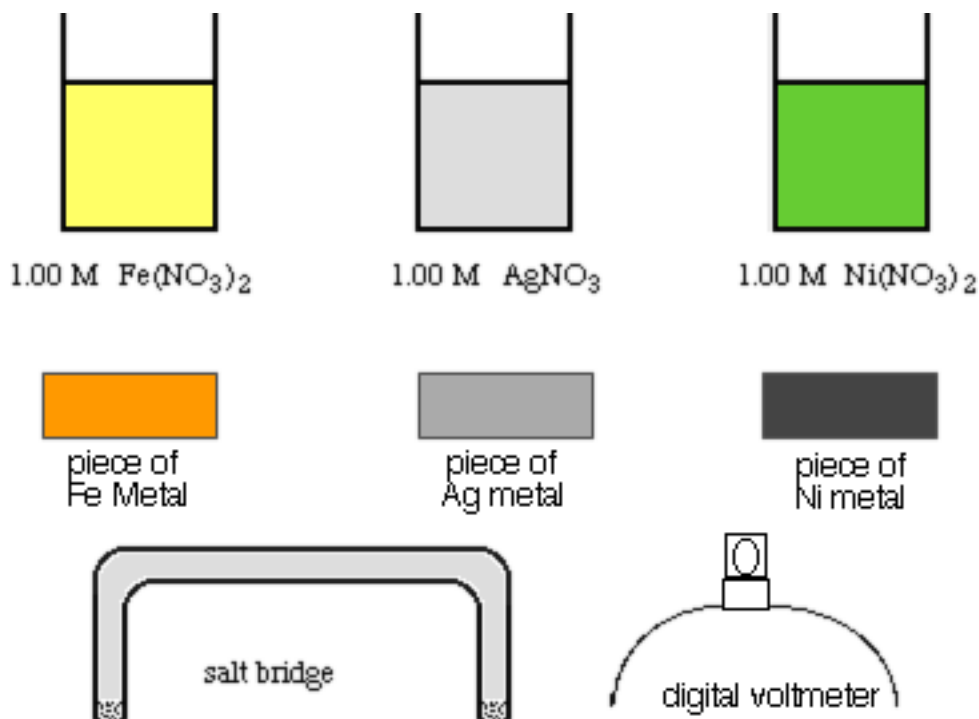
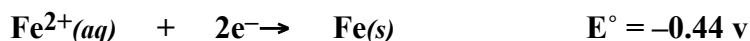
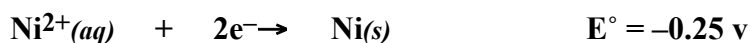


Figure I

- 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)

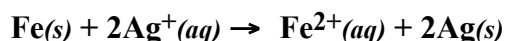
The three reduction reactions from the table of standard reduction potentials are;



To calculate the most positive cell potential, we use the equation;

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

To get the most positive cell potential $\text{Ag}^+(\text{aq})$ must be reduced and $\text{Fe}(\text{s})$ must be oxidized. So the balanced net ionic reaction is;



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

The reducing agent is the species that is oxidized. Fe(s) is oxidized in the reaction, so Fe(s) is the reducing agent. Remember metals like to lose electrons, so they are excellent reducing agents.

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

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$E^\circ_{\text{cell}} = +0.80 \text{ v} - (-0.44 \text{ v}) = +1.24 \text{ v}$$

- 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)

Since Fe(s) is oxidized in the reaction the anode material is Fe(s). The cathode compartment must contain $\text{Ag}^+(\text{aq})$ and an anion that forms a soluble salt. The best anion would be $\text{NO}_3^-(\text{aq})$.

- e) Which of the following compounds is most appropriate to use in a salt bridge, NaOH, CuS, or NaNO_3 ? Justify your answer with a brief explanation why you selected the particular compound, and why you did not select the other compounds. (6)

The best choice is NaNO_3 . NaNO_3 is soluble in water and neither cation ion or anion will form insoluble ionic solids when migrating into the anode or cathode compartments.

NaOH is not a good choice because the anion hydroxide, OH^- will form an insoluble solid in the presence of $\text{Fe}^{2+}(\text{aq})$. Since $\text{Fe}^{2+}(\text{aq})$ is formed in the anode compartment and OH^- ions will migrate towards the anode compartment a precipitate of $\text{Fe}(\text{OH})_2$ will form, reducing the concentration of $\text{Fe}^{2+}(\text{aq})$ in the solution and changing the cell potential.

CuS is not a good choice because it is insoluble in water, so a salt bridge solution prepared by adding CuS to water contains so few ions the cell will not discharge and no reaction will occur.

- 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)

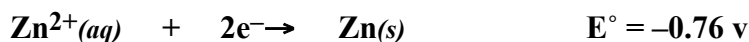
$$E_{\text{cell}} = E^\circ - \frac{0.0592}{n} \log \frac{[\text{Fe}^{2+}]}{[\text{Ag}^+]^2}$$

The cathode compartment contains $\text{Ag}^+(\text{aq})$ ions, that appear in the denominator of the Nernst equation. Reducing the concentration to 0.25 M means the denominator is a small number, and the quotient will be larger than one. The log of a number greater than one is positive so the term $\frac{0.0592}{n} \log \frac{[\text{Fe}^{2+}]}{[\text{Ag}^+]^2}$ is positive and is subtracted from the standard cell potential. This make E_{cell} smaller (less positive).

15. Short Answer

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

Since the only half reaction that contains Zn is;



The ions in the anode compartment (where oxidization occurs) are $\text{Zn}^{2+}(\text{aq})$ ions.

- ii. The salt bridge is a saturated solution of NH_4NO_3 . Which ion migrates towards the anode? Explain why that ion migrates towards the anode. (4)

The $\text{NO}_3^-(\text{aq})$ ions migrate towards the anode compartment. Since oxidation occurs in the anode compartment, Zn metal is oxidized to form $\text{Zn}^{2+}(\text{aq})$ ions, increasing the concentration of cations in the anode compartment. To maintain a charge balance $\text{NO}_3^-(\text{aq})$ ions in the salt bridge must migrate towards the anode compartment.

- 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)

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$

$$+1.56 \text{ v} = E^\circ_{\text{cathode}} - (-0.76 \text{ v})$$

$$E^\circ_{\text{cathode}} = +1.56 \text{ v} + (-0.76 \text{ v}) = +0.80 \text{ v}$$

Look up a standard reduction potential of +0.80 volts and the reduction reaction is,



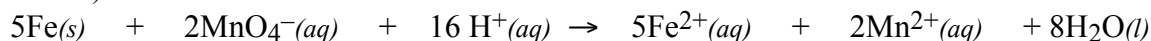
v. Write the half-reaction that occurs at the cathode.(2)



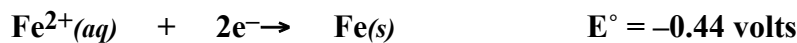
vi. What is the electrode metal for the cathode? (2)

The cathode is made of silver solid.

b) In the reaction



i. Calculate E° for the reaction. (4)



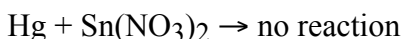
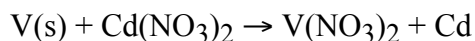
$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = +1.51 \text{ v} - (-0.44 \text{ v}) = +1.95 \text{ v}$$

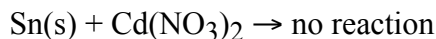
ii. If the pH of the solution is raised to 3, and assuming all other ions in the solution remain 1 M, will the E_{cell} increase, decrease or remain the same? Briefly, explain your answer. (NOTE: No calculation is required.) (4)

$$E_{\text{cell}} = E^\circ - \frac{0.0592}{n} \log \frac{[\text{Fe}^{2+}]^5 [\text{Mn}^{2+}]^2}{[\text{MnO}_4^-]^2 [\text{H}^+]^{16}}$$

Raising the pH to 3 means the $[\text{H}^+]$ is reduced to 1×10^{-3} M. So the denominator now has a value of 1×10^{-48} . The ratio is now larger than 1 so the log of a number larger than one is positive, so the $\frac{0.0592}{n} \log \frac{[\text{Fe}^{2+}]^5 [\text{Mn}^{2+}]^2}{[\text{MnO}_4^-]^2 [\text{H}^+]^{16}}$ term is positive and will be subtracted from the positive E° , so E_{cell} will decrease.

16. Given the following reactions and information about the spontaneous direction, prepare the corresponding activity series showing the four reduction half-reactions ($\text{V}^{2+} + 2e^- \rightarrow \text{V}$; $\text{Hg}^{2+} + 2e^- \rightarrow \text{Hg}$; $\text{Sn}^{2+} + 2e^- \rightarrow \text{Sn}$; $\text{Cd}^{2+} + 2e^- \rightarrow \text{Cd}$) 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.

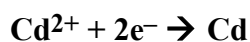




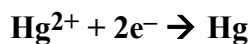
From reaction 1) we know that Cd^{2+} is more likely to be reduced compared to V^{2+} ;



From reaction 3) Sn can not reduce Cd^{2+} ;

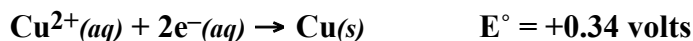
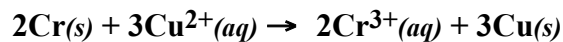


From reaction 2) Hg cannot reduce Sn^{2+} ;



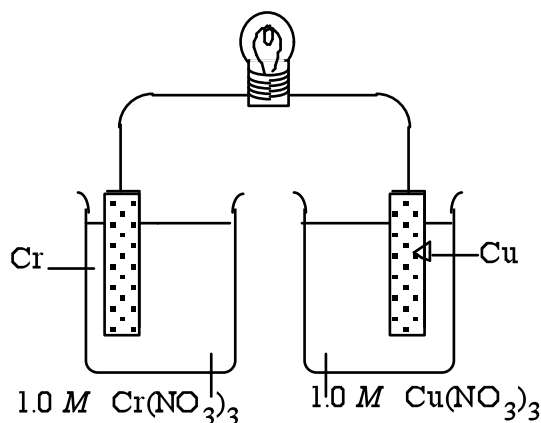
17. A galvanic cell is constructed using a chromium electrode in a 1.00-molar solution of $\text{Cr(NO}_3)_3$ and a copper electrode in a 1.00-molar solution of $\text{Cu(NO}_3)_2$. Both solutions are at 25°C .

- (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)



To get a positive cell potential $E^\circ = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = 0.34 \text{ v} - (-0.74 \text{ v}) = +1.08 \text{ v}$

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



- (i) Which metal is the cathode? (2)

Cu metal is the cathode

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

A salt bridge is needed.

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

The ions in the salt bridge migrate towards the electrode compartments to balance the charge. As oxidation occurs in the anode compartment, cations are produced, so the anions in the salt bridge must migrate towards the anode compartment. In the cathode compartment, reduction occurs and cations are removed, so cations in the salt bridge migrate towards the cathode compartment.

- (c) How does the potential of this cell change if the concentration of $\text{Cr}(\text{NO}_3)_3$ is changed to 3.00-molar at 25°C ? Explain. (4)

$$E_{\text{cell}} = E^\circ - \frac{0.0592}{n} \log \frac{[\text{Cr}^{3+}]^2}{[\text{Cu}^{2+}]^3}$$

Since $[\text{Cr}^{3+}]$ is in the numerator in the ratio, a higher concentration will make the ratio greater than one. The log of a number larger than one is positive so the $\frac{0.0592}{n} \log \frac{[\text{Cr}^{3+}]^2}{[\text{Cu}^{2+}]^3}$ term must be subtracted from the standard cell potential, and the cell potential decreases.

- (d) Calculate K and ΔG° for the reaction you wrote in part a). (14)

$$E^\circ = \frac{0.0592}{n} \log K$$

$$+1.08 = \frac{0.0592}{6} \log K$$

$$+109.46 = \log K$$

$$10^{+109.46} = 10^{\log K}$$

$$K = 2.88 \times 10^{109}$$

$$\Delta G^\circ = -nFE^\circ$$

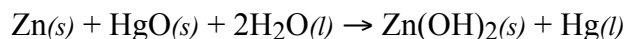
$$\Delta G^\circ = -6 \cdot 96,500 \frac{\text{J}}{\text{mol} \cdot \text{volt}} + 1.08 \text{ v} = -625,000 \frac{\text{J}}{\text{mol}} \left(\frac{1 \text{ kJ}}{1000 \text{ J}} \right) = -625 \frac{\text{kJ}}{\text{mol}}$$

- 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 $\text{X}(\text{NO}_3)_2$. The measured standard cell potential for the reaction between X and $\text{Cr}(\text{NO}_3)_3$ is +0.44 v. Calculate the standard reduction potential for X^{2+} . (4)

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

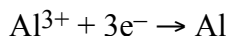
- A) $\text{Na}_2\text{CO}_3(s) + 2\text{HCl}(aq) \rightarrow 2\text{NaCl}(aq) + 2\text{H}_2\text{O}(l) + \text{CO}_2(g)$
- B) $\text{H}_2\text{SO}_4(aq) + \text{Ba}(\text{OH})_2(aq) \rightarrow \text{BaSO}_4(s) + 2\text{H}_2\text{O}(l)$
- C) $\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l)$
- D) $\text{Pb}^{2+}(aq) + \text{Ni}(s) \rightarrow \text{Pb}(s) + \text{Ni}^{2+}(aq)$**
- E) $\text{NaCl}(s) \rightarrow \text{Na}^+(aq) + \text{Cl}^-(aq)$

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



the anode reaction must be

- A) $\text{HgO}(s) + 2\text{H}_2\text{O}(l) + 2e^- \rightarrow \text{Hg}(l) + 2\text{OH}^-(aq)$
- B) $\text{Zn}(s) + 2\text{OH}^-(aq) \rightarrow \text{Zn}(\text{OH})_2(s) + 2e^-$**
- C) $\text{Hg}(l) + 2\text{OH}^-(aq) \rightarrow \text{HgO}(s) + 2\text{H}_2\text{O}(l) + 2e^-$
- D) $\text{Zn}(\text{OH})_2(s) + 2e^- \rightarrow \text{Zn}(s) + 2\text{OH}^-(aq)$
19. Aluminum oxide may be electrolyzed at 1000 °C to furnish aluminum metal. The cathode reaction is

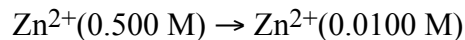


To prepare 5.12 kg of aluminum metal by this method would require

- 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

$$5.12 \text{ kg} \left(\frac{1000 \text{ g}}{1 \text{ kg}} \right) \left(\frac{1 \text{ mol Al}}{27.0 \text{ g}} \right) \left(\frac{3 \text{ mol } e^-}{1 \text{ mol Al}} \right) \left(\frac{96500 \text{ coulomb}}{1 \text{ mol } e^-} \right) \\ = 5.49 \times 10^7 \text{ C}$$

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

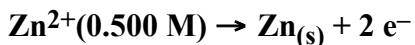


is

- A) +0.050 v
- B) 0.00 v
- C) -0.050 v
- D) can not be determine without additional information

$$E_{\text{cell}} = E^{\circ} - \frac{0.0592}{n} \log \frac{[\text{Zn}^{2+}]_{\text{(dilute)}}}{[\text{Zn}^{2+}]_{\text{(concentrated)}}}$$

Note: Since the half-reactions must be;



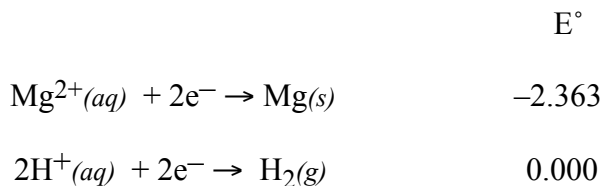
2 electrons must be transferred.

$$E_{\text{cell}} = E^{\circ} - \frac{0.0592}{2} \log \frac{[\text{Zn}^{2+}]_{\text{(dilute)}}}{[\text{Zn}^{2+}]_{\text{(concentrated)}}}$$

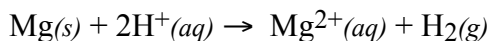
$$E_{\text{cell}} = 0.00 \text{ volts} - \frac{0.0592}{2} \log \frac{0.0100}{0.500}$$

$$E_{\text{cell}} = 0.00 \text{ volts} - (-0.0503 \text{ volts}) = +0.0503 \text{ volts}$$

21. Given the following half-reactions



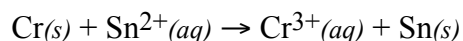
The measured cell potential for the spontaneous reaction



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

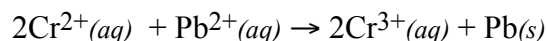
- A) 0
- B) 1
- C) 7
- D) 10

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



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,



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

$$E_{\text{cell}} = E^\circ - \frac{0.0592}{n} \log \frac{[\text{Cr}^{3+}]^2}{[\text{Cr}^{2+}]^2[\text{Pb}^{2+}]}$$

$$E_{\text{cell}} = E^\circ - \frac{0.0592}{2} \log \frac{[0.0300]^2}{[0.250]^2[0.150]}$$

$$E_{\text{cell}} = E^\circ - \frac{0.0592}{2} \log \frac{[0.0300]^2}{[0.250]^2[0.150]}$$

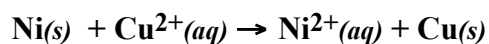
24. Calculate K and ΔG° for the reaction in Problem #5.

$$E^\circ = \frac{0.0592}{n} \log K$$

$$\Delta G^\circ = -nFE^\circ$$

25. A voltaic cell consists of an anode compartment with a nickel electrode immersed in a NiSO₄ solution and a cathode compartment with a copper electrode immersed in a CuSO₄ 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?

$$1.50 \text{ amps} \cdot 2 \text{ hrs} \left(\frac{3600 \text{ sec}}{1 \text{ hr}} \right) \left(\frac{1 \text{ coulomb}}{1 \text{ amp-sec}} \right)$$

$$1.50 \text{ amps} \cdot 2 \text{ hrs} \left(\frac{3600 \text{ sec}}{1 \text{ hr}} \right) \left(\frac{1 \text{ coulomb}}{1 \text{ amp-sec}} \right) \left(\frac{1 \text{ mol e}^-}{96500 \text{ coulomb}} \right)$$

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. $\text{AlCl}_3(l)$

ANODE:

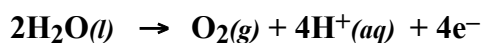


CATHODE:

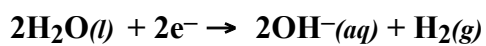
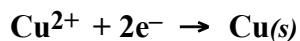


b. $\text{CuBr}_2(aq)$

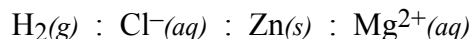
ANODE:



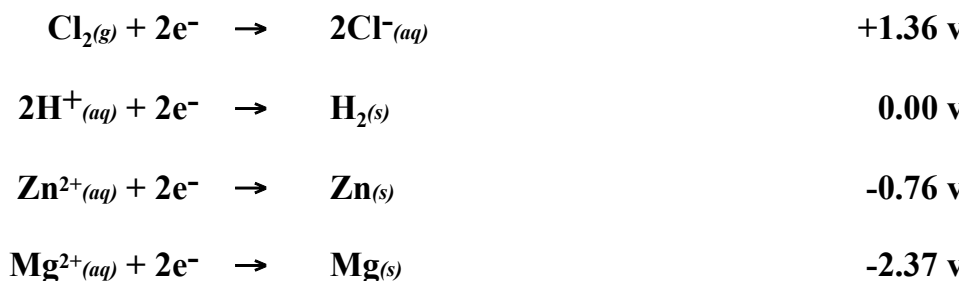
CATHODE:



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



The strongest reducing agent (the species that will be oxidized) is the species with the most negative reduction potential. So must look for the species on the right side of the table of standard reduction potentials.



We can eliminate $\text{Mg}^{2+}(aq)$ immediately because it cannot be oxidized. Of the remaining three species Zn has the most negative

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



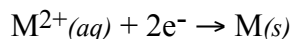
Yes, the cell potential for the reaction ($\text{Sn}(s) + \text{Ag}^+(aq) \rightarrow \text{Sn}^{2+}(aq) + \text{Ag}(s)$) is positive.

29. Will gold dissolve in $\text{HCl}(aq)$?



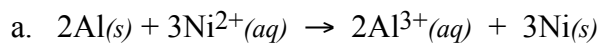
No, the cell potential for the reaction ($\text{Au} + 2\text{H}^+(aq) \rightarrow \text{H}_2(g) + \text{Au}(s)$) is negative.

30. From the following information estimate the E° for

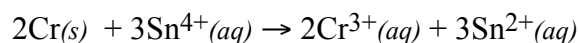


knowing the metal, M, dissolves in 1 M HCl. It will displace $\text{Fe}^{2+}(\text{aq})$, but not $\text{Al}^{3+}(\text{aq})$.

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



32. A voltaic cell is constructed that the reaction is,



If the $[\text{Cr}^{3+}] = 1.00 \text{ M}$, $[\text{Sn}^{4+}] = 0.500 \text{ M}$ and the $[\text{Sn}^{2+}] = 0.025$, calculate E_{cell} at 25°C .

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

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. $\text{KBr}(l)$

ANODE:

CATHODE:

b. $\text{NiCl}_2(aq)$

ANODE:

CATHODE:

35. When a dilute solution of H_2SO_4 is electrolyzed, $\text{O}_2(g)$ is produced at the anode and $\text{H}_2(g)$ 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.

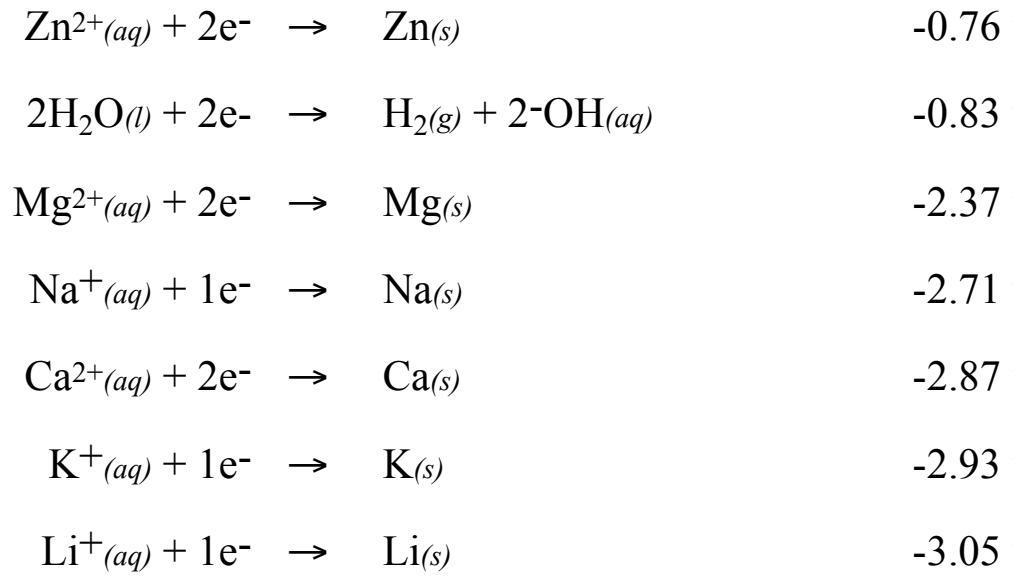
b. Calculate the mass of O_2 and H_2 produced when the cell operates for 15.0 minutes at an average current of 8.50 amps.

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

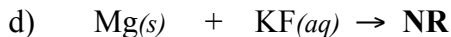
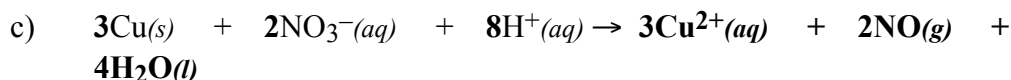
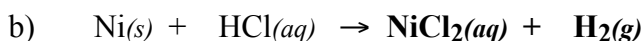
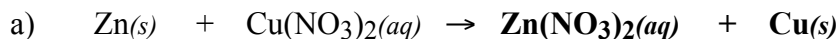
37. Using the table of standard reduction potentials, calculate K_{sp} for $PbSO_4$.

Standard Reduction Potentials at 25 °C

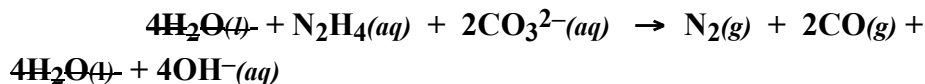
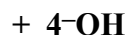
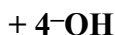
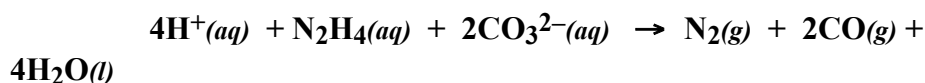
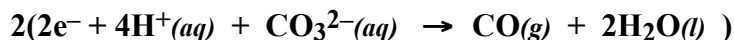
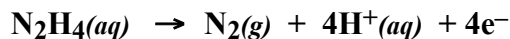
<u>Half-Reaction</u>	<u>E°</u>
$F_{2(g)} + 2e^- \rightarrow 2F^-_{(aq)}$	+2.87
$H_2O_{2(aq)} + 2H^+_{(aq)} + 2e^- \rightarrow 2H_2O_{(l)}$	+1.77
$MnO_4^-_{(aq)} + 8H^+_{(aq)} + 5e^- \rightarrow Mn^{2+}_{(aq)} + 4H_2O_{(l)}$	+1.52
$Cl_{2(g)} + 2e^- \rightarrow 2Cl^-_{(aq)}$	+1.36
$Cr_2O_7^{2-}_{(aq)}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



(12) 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.

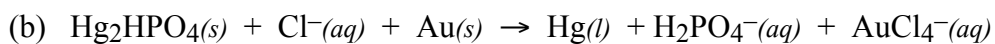


(18) 2. Balance the following oxidation-reduction reactions using the half-reaction method.



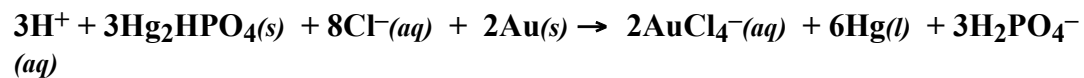
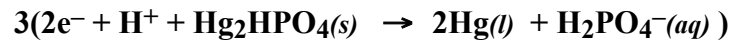
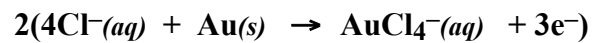
identify the oxidizing agent CO_3^{2-}

identify the reducing agent $\text{N}_2\text{H}_4(aq)$



acidic

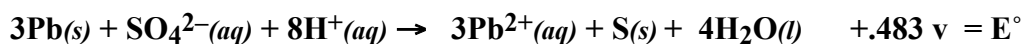
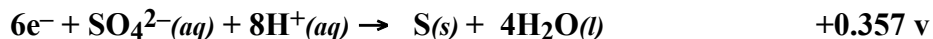
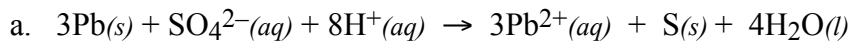
solution



identify the oxidizing agent $\text{Hg}_2\text{HPO}_4(s)$
 $\text{Au}(s)$

identify the reducing agent

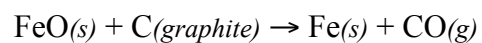
- (8) 3. Write the half-reactions and determine E° for the electrochemical cells as described below, in part b include the overall chemical equation.



- (14) 4. Given the following data

Substance	$\Delta H^\circ \left(\frac{\text{kJ}}{\text{mol}} \right)$	$S^\circ \left(\frac{\text{J}}{\text{mol} \cdot \text{K}} \right)$	$\Delta G^\circ \left(\frac{\text{kJ}}{\text{mol}} \right)$
$\text{FeO}(s)$	-266.3	57.49	-245.12
$\text{C}(\text{graphite})$	0	5.74	0
$\text{Fe}(s)$	0	27.28	0
$\text{CO}(g)$	-110.5	197.6	-137.15

- a. Using the above data, calculate ΔG° , ΔS° , and ΔH° for the reaction,



$$\Delta H^\circ_{\text{rxn}} = \sum(\Delta H_f^\circ(\text{CO})) - \sum(\Delta H_f^\circ(\text{FeO})) = -110.5 \text{ kJ} - (-266.3 \text{ kJ}) = +155.8 \text{ kJ}$$

$$\begin{aligned}\Delta S^\circ_{\text{rxn}} &= \sum(S^\circ(\text{Fe}) + S^\circ(\text{CO})) - \sum(S^\circ(\text{FeO}) + S^\circ(\text{C})) \\ &= \sum\left(27.28 \frac{\text{J}}{\text{K}} + 197.6 \frac{\text{J}}{\text{K}}\right) - \sum\left(57.49 \frac{\text{J}}{\text{K}} + 5.74 \frac{\text{J}}{\text{K}}\right) = +162 \frac{\text{J}}{\text{K}}\end{aligned}$$

$$\Delta G^\circ_{\text{rxn}} = \sum(\Delta G_f^\circ(\text{CO})) - \sum(\Delta G_f^\circ(\text{FeO})) = -137.15 \text{ kJ} - (-245.12 \text{ kJ}) = +108 \text{ kJ}$$

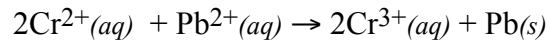
b. Determine at what temperature is the reaction spontaneous.

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

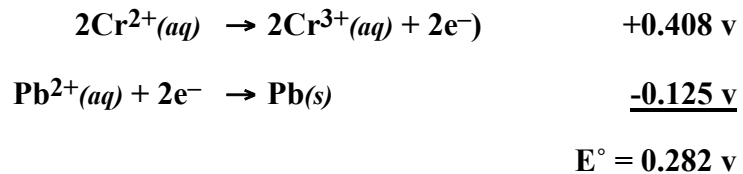
$$\Delta G^\circ = 0$$

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

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



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



$$E_{\text{cell}} = E^\circ - \frac{.059}{n} \log Q$$

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

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

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

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

$$= +0.282 + 0.0301$$

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

(8) 6. Calculate K and ΔG° for the reaction in Problem #5.

$$\Delta G^\circ = -nFE^\circ = -2 \cdot 96,500 \frac{\text{J}}{\text{mol}\cdot\text{v}} (+0.282 \text{ v}) = -54.4 \text{ kJ}$$

$$E^\circ = \frac{.059}{n} \log K$$

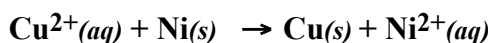
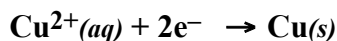
$$+0.282 \text{ v} = \frac{.059}{2} \log K$$

$$9.543 = \log K$$

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

- (12) 7. 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?

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

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

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

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

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

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

mass increases at the cathode

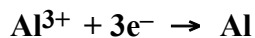
- (8) 8. Write the half-reactions which occur at each electrode when the following solutions are electrolyzed between inert electrodes. If more than one half-reaction can occur at the anode or cathode, write all possible half-reactions and briefly explain which half-reaction will occur.

a. $\text{AlCl}_3(l)$

ANODE:

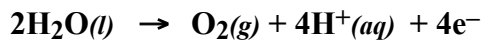


CATHODE:

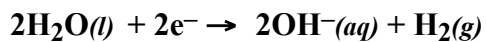
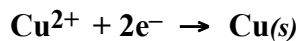


b. $\text{CuBr}_2(aq)$

ANODE:



CATHODE:



Short Answer: (12 points)

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



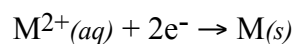
10. Will $\text{Sn}(\text{s})$ displace $\text{Ag}^+(\text{aq})$ from solution?

YES

11. Will gold dissolve in $\text{HCl}(\text{aq})$?

NO

12. From the following information estimate the E° for



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

Mn or Zn

