<u>Redox Questions: Unit Review</u>

1. Briefly define oxidation.
2. Would you call a substance that is easily oxidized an oxidizing agent or a reducing agent? Explain.
3. If a neutral atom becomes positively charged, has it been oxidized or reduced? Write a general equation using *M* for the neutral atom.
4. If an ion X^{-1} acquires a –2 charge, has it been oxidized or reduced? Write a general equation.
5. In order to make $Na_{(s)}$ and $Cl_{2\,(g)}$, an electric current is passed through $NaCl_{(l)}$. What does the energy supplied to this reaction do?
6. The four oxyacids of chlorine are: hypochlorous acid, chlorous acid, chloric acid and perchloric acid. The respective formulas for these acids are: HclO, $HClO_2$, $HClO_3$, and $HClO_4$. What is the oxidation number of the chlorine atom in each of the acids?
7. Figure ... shows elections leaving the $Cu_{(s)}$ and going to the $Ag_{(s)}$. Experimentally, both half-cells are found to be electrically neutral before current flows and to remain so as the cell operates. Explain.
8. Suppose water is added to each of the two beakers containing copper sulfate in two different electrochemical cells. What change will occur in the voltage in each cell?
9. Determine the oxidation number of carbon in the compounds carbon monoxide, carbon dioxide and in diamond.
10. Determine the oxidation number of uranium in each of the compounds: UO_3, U_3O, U_2O_5, UO_2, K_2UO_4, $Mg_2U_2O_7$, UO.
11. One method of obtaining copper metal is to let a solution containing Cu^{+2} ions trickle over scrap iron. Write equations for the two half-reactions involved. Assume the iron becomes Fe^{+2}. Indicate in which half-reaction oxidation is taking place.
12. Aluminum metal reacts with aqueous acidic solutions to liberate hydrogen Aluminum metal reacts with aqueous acidic solutions to liberate hydrogen gas. Write the two half-reactions and the net ionic reaction.
13. Nickel metal reacts with cupric ions, Cu^{+2}, but not with zinc ions, Zn^{+2}. In each case of reaction, ions of +2 charge are formed. Use these data to expand the list in your Data Book of Redox Potentials.
14. Suppose chemists had chosen to call the $2I^- \rightarrow I_2 + 2e^-$ half-cell potential zero.
    a. What would be E° for $Na \rightarrow Na^+ + e^-$?
    b. How much would the net potential for the reaction $2Na + I_2 \rightarrow 2Na^+ + 2I^-$ change?
15. If a piece of copper metal is dipped into a solution containing Cr^{+3} ions, what will happen? Explain using half-cell potentials.
16. What would happen if an aluminum spoon were used to stir an $Fe(NO_3)_2$ solution? What would happen if an iron spoon were used to stir an $AlCl_3$ solution?
17. Can 1 *M* $Fe_2(SO_4)_3$ solution be stored in a container make of nickel? Explain your answer.
18. When copper is placed in concentrated nitric acid, vigorous bubbling takes place as a brown gas is evolved. The copper disappears, and the colorless solution changes to a greenish-blue. The brown gas is nitrogen dioxide, NO_2, and the solution's color is due to the formation of cupric ion (Cu^{+2}). Using half-reactions, write the net ionic equation for this reaction.

19. In acid solution the following are true: H_2S will react with oxygen and give H_2O and sulfur. H_2S will not react in a corresponding manner with selenium or tellurium. H_2Se will react with sulfur giving H_2S and selenium, but it will not react with tellurium. Arrange the hydrides H_2O, H_2S, H_2Se, and H_2Te in order of their tendency to lose electrons to form the elements O_2, S, Se, and Te.

20. If you wish to replate a silver spoon, would you make it the anode or the cathode in a cell? Use half- reactions in your explanation. How many moles of electrons are needed to plate out a gram of Ag?

20. In the electrolysis of aqueous cupric bromide, $CuBr_2$, 0.500 gram of copper is deposited at one electrode. How many grams of bromine are formed at the other electrode? Write the anode and cathode half-reactions.

21. Complete the following equations. Determine the net potential of each cell and decide whether reaction can occur.
    a. $Zn + Ag^+ \rightarrow$
    b. $Cu + Ag^+ \rightarrow$
    c. $Sn + Fe^+ \rightarrow$
    d. $Hg + H^+ \rightarrow$

22. For each of the following, write the half-reactions; determine the net reaction; predict whether the reaction can occur, giving the basis for your prediction.
    a. $Mg_{(s)} + Sn^{+2} \rightarrow$
    b. $Mn_{(s)} + Cs^{+2} \rightarrow$
    c. $Cu_{(s)} + Cl_2 \rightarrow$
    d. $Zn_{(s)} + Fe^{+2} \rightarrow$
    e. $Fe_{(s)} + Fe^{+3} \rightarrow$

23. A half-cell consisting of a palladium rod dipping into a 1 *M* $Pd(NO_3)_2$ solution is connected with a standard hydrogen half-cell. The cell voltage is 0.99 volt, and the platinum electrode in the hydrogen half-cell is the anode. Determine E° for the reaction:

    $Pd \rightarrow Pd^{+2} + 2e^-$

24. Use oxidation numbers to balance the reaction between ferrous ion, Fe^{+2}, and permanganate ion, MnO_4^-, in acid solution to produce ferric ion, Fe^{+3}, and manganous ion, Mn^{+2}.

25. By use of half-reactions, give a balanced equation for each of the following reactions:
    a. $H_2O_2 + I^- + H^+$ gives $H_2O + I_2$
    b. $Cr_2O_7^{-2} + Fe^{+3} + H^+$ gives $Cr^{+3} + Fe^{+3} + H_2O$
    c. $Cu + NO_3^- + H^+$ gives $Cu^{+2} + NO + H_2O$
    d. $MnO_4^- + Sn^{+2} + H^+$ gives $Mn^{+2} + Sn^{+4} + H_2O$

26. By use of oxidation numbers, give a balanced equation for each of the following reactions:
    a. $HBr + H_2SO_4$ gives $SO_2 + Br_2 + H_2O$
    b. $NO_3^- + Cl^- + H^+$ gives $NO + Cl_2 + H_2O$
    c. $Zn + NO_3^- + H^+$ gives Zn^{+2} _ $NO_2 + H_2O$
    d. BrO^- gives $Br^- + BrO_3^-$

27. Show the arbitrariness of oxidation numbers by balancing the reaction discussed in Problem 25 with the assumption that the oxidation number of manganese in MnO_4^- is +2. Compare with the result obtained in Problem 25.

28. In each of the following equations, identify (1) the atom or ion oxidized, (2) the atom or ion reduced, (3) the oxidizing agent, (4) the reducing agent, (5) the oxidation number of all atoms or ions in each compound, (6) the change in oxidation number associated with each oxidation process, (7) the change in oxidation number associated with each reduction process. (8) Then use the information obtained in answering part 6 and 7 to balance each equation:
    a. $CH_4 + O_2 \rightarrow CO_2 + H_2O$
    b. $Mn + Cl_2 \rightarrow MnCl_2$
    c. $PbO_2 + Pb + H_2SO_4 \rightarrow PbSO_4 + H_2O$
    d. $NH_3 + O_2 \rightarrow NO + H_2O$
    e. $HNO_3 + H_2S \rightarrow NO_{(g)} + S + H_2O$
    f. $CuO + NH_3 \rightarrow N_2 + H_2O + Cu_{(s)}$
29. When a metallic silver strip is placed in a solution of $Au(NO_3)_3$, metallic gold plates out on the silver strip and Ag^+ ions go into solution. Which is the better competitor for electrons: Au^{+++} or Ag^+?
30. In an experiment similar to Experiment 27 in the Laboratory Manual, strips of gold, silver, and tin are placed in beakers containing solutions of Au^{+++}, Ag^+, and Sn^{++} ions. The following results are observed:

    Au^{+++} + Sn: metallic gold is deposited on tin strip
    Au + Ag^+: no reaction
    Sn + Ag^+: metallic silver is deposited on tin strip

Arrange the ions above in order of *decreasing* tendency to attract electrons (that is, place the ion with the greatest attraction for electrons on top).

31. Balance the following oxidation-reduction equations, using the oxidation number method. Be sure charge and atoms are conserved.
    a. $Cu + Ag^+ \rightarrow Ag + Cu^{++}$
    b. $MnO_4^- + H_2S + H^+ \rightarrow S + Mn^{++} + H_2O$
    c. $Cu_2S + O_2 \rightarrow Cu_2O + SO_2$
    d. $FeCl_3 + SnCl_2 \rightarrow SnCl_4 + FeCl_2$
32. The potential E° for two half-cell processes is:

    $Cu \rightarrow Cu^{++} + 2e^-$ E° = -0.34
    $Zn \rightarrow Zn^{++} + 2e^-$ E° = +0.76
    What is the E° for the process:

    $$Cu^{++} + Zn \rightarrow Zn^{++} + Cu$$

    Does E° show that the process is spontaneous? Explain.