Name ______________________ Section ______ Date ______

Chapter 18: Basic Review Worksheet

1. What is meant by the term *oxidation*? What is meant by the term *reduction*? Answer these in terms of electrons, and in terms of oxidation states.

2. Determine the oxidation states of the atoms in the following substances
   a. Cr_2O_3
   b. $FeCl_2$
   c. Na_3PO_4

3. What is meant by *oxidizing agent*? What is meant by *reducing agent*? Explain the statement "an oxidizing agent is reduced and a reducing agent is oxided".

4. What are half-reactions? Why do we use them?

5. Answer the following:
   a. When a half-reaction is reversed, what happens to its potential?
   b. When the coefficients of a half-reaction are multiplied by a factor, what happens to its potential?

6. How is balancing oxidation-reduction equations similar to the method for balancing equations you learned in Chapter 7? How is it different?

7. Balance the following oxidation-reduction equations.
   a. $Mg(s) + Hg^{2+}(aq) \rightarrow Mg^{2+}(aq) + Hg_2^{2+}(aq)$
   b. $Zn(s) + Ag^{+}(aq) \rightarrow Zn^{2+}(aq) + Ag(s)$

8. What is an *anode*? What is a *cathode*?

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1. Identify the elements that are oxidized and the elements that are reduced in the following equations.
   a. $2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$
   b. $2K(s) + I_2(s) \rightarrow 2KI(s)$
   c. $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

2. Determine the oxidation states of the atoms in the following.
   a. $KMnO_4$
   b. $HCrO_4^-$
   c. BiO^+

3. Identify the oxidizing and reducing agents in the following equations.
   a. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
   b. $2Cu(s) + S(s) \rightarrow Cu_2S(s)$

4. Why must the number of electrons be balanced in an oxidation-reduction reaction?

5. Balance the following oxidation-reduction equations in acidic solution.
   a. $NO_3^-(aq) + Br^-(aq) \rightarrow NO(g) + Br_2(l)$
   b. $Ni(s) + NO_3^-(aq) \rightarrow Ni^{2+}(aq) + NO_2(g)$

6. Why must we separate the oxidizing agent from the reducing agent when constructing an electrochemical cell?

7. Consider the oxidation-reduction reaction

$$Cu(s) + Ag^+(aq) \rightarrow Cu^{2+}(aq) + Ag(s)$$

   a. Balance the oxidation-reduction equation.
   b. Label the oxidizing agent and the reducing agent.
   c. Sketch a galvanic cell that uses this reaction and label the anode and the cathode.

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1. Determine the oxidation states of the atoms in the molecule OF_2.

2. Not all oxidation-reduction reactions involve a metal and nonmetal. Give an example that does not and explain how it is an oxidation-reduction reaction.

3. Balance the following oxidation-reduction equations for the conditions specified.
   a. $S_2O_8^{2-}(aq) + Cr^{3+}(aq) \rightarrow SO_4^{2-}(aq) + Cr_2O_7^{2-}(aq)$ (acidic solution)
   b. $ClO_4^-(aq) + Cl^-(aq) \rightarrow ClO_3^-(aq) + Cl_2(g)$ (acidic solution)
   c. $Cl_2(g) \rightarrow Cl^-(aq) + ClO^-(aq)$ (basic solution)

5. Examples will vary. Consider the following:

$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$ $\quad K_{sp} = [Ag^+][Cl^-]$

$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$ $\quad K_{sp} = [Pb^{2+}][Cl^-]^2$

$BaSO_4(s) \rightleftharpoons Ba^{2+}(aq) + SO_4^{2-}(aq)$ $\quad K_{sp} = [Ba^{2+}][SO_4^{2-}]$

6. $8.0 \times 10^{-2} g \times \dfrac{1 \text{ mol}}{62.31 g\ MgF_2} = 1.3 \times 10^{-3}$ mol

$$K_{sp} = [Mg^{2+}][F^-]^2 = (1.3 \times 10^{-3})(2.6 \times 10^{-3})^2 = 8.8 \times 10^{-9}$$

Chapter 18: Basic Review Worksheet

1. Oxidation may be defined as the loss of electrons by an atom or as an increase in the oxidation state of the atom. Reduction may be defined as the gain of electrons by an atom or as a decrease in the oxidation state of the atom.

2. a. Cr = +3, O = +2
   b. Fe = +2, Cl = -1
   c. Na = +1, P = +5, O = -2

3. An oxidizing agent is a molecule, atom, or ion that causes the oxidation of some other species; it does so by accepting one or more electrons. A reducing agent is a molecule, atom, or ion that causes reduction of some other species; it does so by giving up one or more electrons. Therefore, an oxidizing agent, by taking electrons, is reduced. A reducing agent, by giving electrons, is oxidized.

4. A half-reaction is the partial equation of an overall oxidation-reduction equation which represents either oxidation or reduction. We use them because they make it easier to balance oxidation-reduction equations.

5. a. When a half-reaction is reversed, the cell potential is reversed.
   b. When the coefficients of a half-reaction are multiplied by a factor, the cell potential is unchanged.

6. Like those covered in Chapter 7, oxidation-reduction equations must be balanced by atoms (the total number of each type of atom on each side of the equation must be the same). With oxidation-reduction equations, we must also balance the charge (the number of electrons gained by one species must be lost by another; there are no "free" electrons).

7. a. $Mg(s) + 2Hg^{2+}(aq) \rightarrow Mg^{2+}(aq) + Hg_2^{2+}(aq)$
   b. $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$

8. The anode is the electrode where oxidation occurs in a galvanic cell; the cathode is the electrode where reduction occurs in a galvanic cell.

Chapter 18: Review Worksheet

1. a. aluminum is oxidized; chlorine is reduced
   b. potassium is oxidized; iodine is reduced
   c. nitrogen is oxidized; oxygen is reduced

2. a. K = +1, Mn = +7, O = -2
   b. H = +1, Cr = +6, O = -2
   c. Bi = +3, O = -2

3. a. CH_4 is the reducing agent, O_2 is the oxidizing agent.
   b. Cu is the reducing agent, S is the oxidizing agent.

4. Under ordinary conditions it is not possible to have "free" electrons that are not part of some atom, molecule, or ion. Thus, the number of electrons lost by one species must be gained by another species (the electrons must be balanced).

5. a. $8H^+(aq) + 2NO_3^-(aq) + 6Br^-(aq) \rightarrow 2NO(g) + 3Br_2(l) + 4H_2O(l)$
   b. $4H^+(aq) + Ni(s) + 2NO_3^-(aq) \rightarrow Ni^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$

6. By separating the oxidizing agent and reducing agent, the electron that transfers from the reducing agent to the oxidizing agent must go through a wire. The current produced in the wire can do useful work.

7. a. $Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$
   b. Cu is the reducing agent and Ag^+ is the oxidizing agent.
   c. The cell should look similar to Figure 18.5 in the text with Cu as the anode and Ag as the cathode.

Chapter 18: Challenge Review Worksheet

1. O = +2, F = -1. Normally the oxidation state of O is -2. However, the fluorine atom is more electronegative than the oxygen atom, so the oxidation state of F must be lower than that of O.

2. Answers will vary. The combustion of methane is discussed in Section 18.3 of the text; students should choose a different example.

3. a. $7H_2O(l) + 3S_2O_8^{2-}(aq) + 2Cr^{3+}(aq) \rightarrow 6SO_4^{2-}(aq) + Cr_2O_7^{2-}(aq) + 14H^+(aq)$
   b. $2H^+(aq) + ClO_4^-(aq) + 2Cl^-(aq) \rightarrow ClO_3^-(aq) + Cl_2(g) + H_2O(l)$
   c. $2OH^-(aq) + Cl_2(g) \rightarrow Cl^-(aq) + ClO^-(aq) + H_2O(l)$

Chapter 19: Basic Review Worksheet

1. The protons and neutrons are present in the nucleus.

Particle	Relative Mass	Relative Charge
proton	1.0000	1+
neutron	1.0016	none