AP Chemistry

# Exercise 18.1

## Balancing Redox Equations

Name: ________________________

Date:_________________ Per:_________

**How to Assign Oxidation Numbers: The Fundamental Rules**

Rules for assigning oxidation numbers are as follows:

1. The oxidation number of any pure element is zero. Thus the oxidation number of H in $H_2$ is zero.
2. The oxidation number of a monatomic ion is equal to its charge. Thus the oxidation number of Cl in the $Cl^-$ ion is -1, that for Mg in the $Mg^{2+}$ ion is +2, and that for oxygen in $O^{2-}$ ion is -2.
3. The sum of the oxidation numbers in a compound is zero if neutral, or equal to the charge if an ion.
4. The oxidation number of alkali metals in compounds is +1, and that of alkaline earths in compounds is +2. The oxidation number of F is -1 in all its compounds.
5. The oxidation number of H is +1 in most compounds. Exceptions are $H_2$ (where H = 0) and the ionic hydrides, such as NaH (where H = -1).
6. The oxidation number of oxygen (O) is -2 in most compounds. Exceptions are $O_2$ (where O = 0) and peroxides, such as $H_2O_2$ or $Na_2O_2$, where O = -1.
7. For other elements, you can usually use rule (3) to solve for the unknown oxidation number.

***DIRECTIONS: Determine the oxidation number for each element in the following formulas.***

a. $SnCl_4$ Sn _____ Cl _____
b. $Ca_3P_2$ Ca _____ P _____
c. SnO Sn _____ O _____
d. $Ag_2S$ Ag _____ S _____
e. HI H _____ I _____
f. $N_2H_4$ N _____ H _____
g. $Al_2O_3$ Al _____ O _____
h. $S_8$ S _____

i. $HNO_2$ H _____ N _____ O _____
j. $O_2$ O _____
k. $H_3O^+$ H _____ O _____
l. $ClO_3^-$ Cl _____ O _____
m. $S_2O_3^{2-}$ S _____ O _____
n. $KMnO_4$ K _____ Mn _____ O _____
o. $(NH_4)_2SO_4$ N _____ H _____ S _____ O _____

***DIRECTIONS: Label the oxidation and reduction process in each of the following redox reactions.***

1. $C + H_2SO_4 \rightarrow CO_2 + SO_2 + H_2O$
2. $HNO_3 + HI \rightarrow NO + I_2 + H_2O$
3. $KMnO_4 + HCl \rightarrow MnCl_2 + Cl_2 + H_2O + KCl$
4. $Sb + HNO_3 \rightarrow Sb_2O_3 + NO + H_2O$
5. $2\ Na + FeCl_2 \rightarrow 2\ NaCl + Fe$
6. $2\ C_2H_2 + 5\ O_2 \rightarrow 4\ CO_2 + 2\ H_2O$
7. $2\ PbS + 3\ O_2 \rightarrow 2\ SO_2 + 2\ PbO$
8. $2\ H_2 + O_2 \rightarrow 2\ H_2O$
9. $Cu + HNO_3 \rightarrow CuNO_3 + H_2$
10. $AgNO_3 + Cu \rightarrow CuNO_3 + Ag$

**Balancing Redox Reactions using the Half-Reaction Method**

1. Write the unbalanced oxidation and reduction half-reactions.
2. Balance the equation using atoms other than O & H.
3. Balance for O atoms by adding $H_2O$ to the reaction side deficient in O.
4. This leaves H atoms unbalanced. Balance H atoms by adding $H^+$ to the side deficient in H.
   a. *In basic solution*, follow this step by neutralizing the $H^+$ by adding an equivalent amount of $OH^-$ to both sides of the equation to neutralize the $H^+$.
   b. Then form water on the side which has both $H^+$ and $OH^-$ (recall that $H^+ + OH^- \rightarrow H_2O$).
   c. Next simplify the water by canceling equal amounts of water from each side.
5. Balance for charge. To do this, add electrons ($e^-$) to the more positive side.
6. Multiply the equations by appropriate factors so that the number of electrons lost in the oxidation half-reaction (OIL) is equal to the number of electrons gained in the reduction half-reaction (RIG).
7. Combine the equations and simplify by removing species common to both sides.
8. Check this equation to confirm that it is balanced for atoms and balanced for charge.

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***DIRECTIONS: Balance the following redox equations.***

11. $Mn^{2+} + BiO_3^- \rightarrow MnO_4^- + Bi^{3+}$ (in acidic solution)

12. $ClO_2 \rightarrow ClO_2^- + ClO_3^-$ (pH = 7.8)

13. $ClO_3^- + Cl^- \rightarrow Cl_2 + ClO_2$(pH = 4.7)

14. $MnO_4^- + C_2O_4^{2-} \rightarrow MnO_2 + CO_2$ (pH = 9.5)

15. $NO_2^- + MnO_4^- \rightarrow NO_3^- + Mn^{2+}$ (in acidic solution)

16. $Zn \rightarrow Zn(OH)_4^{2-} + H_2$ (pH = 10.1)

17. $MnO_4^- + S_2O_3^{2-} \rightarrow S_4O_6^{2-} + Mn^{2+}$(pH = 6.1)

18. $Cu(NH_3)_4^{2+} + S_2O_4^{2-} \rightarrow SO_3^{2-} + Cu + NH_3$ (in basic solution)