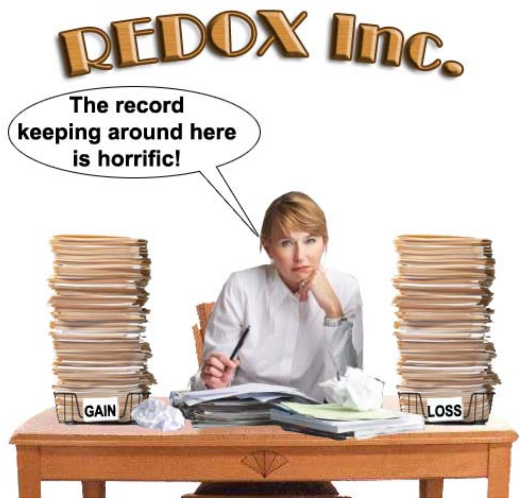
Chemistry: Form WS10.1.4A Name ________________________

ELECTROCHEMISTRY Date _________________ Period _____

Balancing Redox Reactions

Redox equations are often too complex to balance by inspection alone. Instead, they are balanced by the *half-reaction method* or *ion-electron method*. In redox reactions, the number of electrons lost is always equal to the number of electrons gained. Keeping track of the electrons helps to balance the parts of the equation that can’t be ba
lanced by inspection. This is done by the procedure outlined below.



Balance the following: $K_2Cr_2O_7 + H_2S + HCl \rightarrow CrCl_3 + KCl + S + H_2O$

Step 1: Write the ionic equation.

$2K^+ + Cr_2O_7^{2-} + 2H^+ + S^{2-} + H^+ + Cl^- \rightarrow Cr^{3+} + Cl^- + K^+ + Cl^- + S + H_2O$

Step 2: Determine the oxidation states.

$2K^+ + Cr_2O_7^{2-} + 2H^+ + S^{2-} + H^+ + Cl^- \rightarrow Cr^{3+} + Cl^- + K^+ + Cl^- + S + H_2O$

+1 +6 –2 +1 –2 +1 –1 +3 –1 +1 –1 0 +1 –2

Step 3: Write oxidation half reaction, balancing atoms and charge.

$H_2S \rightarrow 2H^+ + S^0 + 2e^-$

Step 4: Write reduction half reaction, balancing atoms and charge.

$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

Step 5: Conserve charge (electrons lost = electrons gained).

$3H_2S \rightarrow 6H^+ + 3S^0 + 6e^-$

$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

Step 6: Combine half reactions.

$Cr_2O_7^{2-} + 3H_2S + 8H^+ \rightarrow 2Cr^{3+} + 3S^0 + 7H_2O$

Step 7: Combine ions to form compounds in original equation.

$K_2Cr_2O_7 + 3H_2S + 8HCl \rightarrow 2CrCl_3 + 2KCl + 3S + 7H_2O$

In **Step 1** the ions are separated making the spectators easier to identify. In **Step 2** the oxidation states are determined so it is possible to tell what was oxidized and what was reduced. In **Steps 3** and **4**, half reactions are written showing the number of electrons transferred. Note that in the oxidation half, $2H^+$ are needed to balance the hydrogen in hydrogen sulfide. In the reduction half, $7H_2O$ are needed to balance the oxygen in the dichromate ion, and as a result $14H^+$ are needed on the reactant side. In **Step 5**, the half reactions are multiplied by the correct coefficients to make the number of electrons lost equal the number of electrons gained. In **Step 6**, note that the H^+ ions remaining are the net from the two half reactions where they are on opposite sides of the equation.

Balance the equations below by following the procedure above.

1. $H_2S(aq) + HNO_3(aq) \rightarrow S(s) + NO_2(g) + H_2O(\ell)$

2. $LiNO_3(aq) + FeCl_2(aq) + HCl(aq) \rightarrow NO(g) + LiCl(aq) + FeCl_3(aq) + H_2O(\ell)$

3. $Na_2Cr_2O_7(aq) + HI(aq) \rightarrow CrI_3(aq) + NaI(aq) + I_2(s) + H_2O(\ell)$

4. $NaClO_3 + HCl \rightarrow ClO_2 + NaClO_4 + NaCl + H_2O$

5. $PbS(s) + HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + S(s) + NO(g) + H_2O(\ell)$