Dougherty Valley HS Chemistry
Redox – Balancing Redox Rxns in Acidic and Basic Solutions

Worksheet #6

Name: **Period:** **Seat#:**

For Half-Reactions in Acidic Solution

Step One: Balance the atom being reduced/oxidized.
In our example, there is already one Mn on each side of the arrow, so this step is already done.

$$\mathbf{MnO_4^{1-} \rightarrow Mn^{2+}}$$

Step Two: Balance the oxygen's.
Do this by adding water molecules (as many as are needed) to the side needing oxygen. In our case, the left side has 4 oxygen's, while the right side has none, so:

$$\mathbf{MnO_4^{1-} \rightarrow Mn^{2+} + 4H_2O}$$

Notice that, when the water is added, hydrogen's also come along. There is nothing that can be done about this; we'll take care of it in the next step. A common question is: "Why can't I just add 4 oxygen atoms to the right side?" Quick answer: don't do it, it's wrong. There are not free oxygen atoms floating around in our reaction vessel. So you can't add them!

Step Three: Balance the hydrogen's.
Do this by adding hydrogen ions (as many as are needed) to the side needing hydrogen. In our example, we need 8 (notice the water molecule's formula, then consider 4 x 2 = 8).

$$\mathbf{8H^{1+} + MnO_4^{1+} \rightarrow Mn^{2+} + 4H_2O}$$

Step Four: Balance the total charge.
This will be done using electrons. It is ALWAYS the last step. First, a comment. You do not need to look at the oxidation number for each atom. You only need to look at the charge on the ion or molecule, then sum those up.

Left side of the reaction, total charge is +7. There are 8 H+, giving 8 x +1 = +8 and a minus one from the permanganate. (A very typical wrong answer for the left side is zero. Someone only counted the +1 and the -1, they forget the 8.)
Right side of the reaction, total charge is +2. The water molecule is neutral (zero charge) and the single Mn is +2.

$$\mathbf{5e^- + 8H^{1+} + MnO_4^{1-} \rightarrow Mn^{2+} + 4H_2O}$$

Five electrons reduces' the +7 to a +2 and the two sides are EQUAL in total charge.
The half-reaction is now correctly balanced!

For Half-Reactions in Basic Solution
Step One to Four: Balance the half-reaction AS IF it were in acid solution.
Just go about steps 1-4 business as usual! The half-reaction is actually in basic solution, but we are going to start out as if it were in acid solution and then fix it later to account for it being in a basic solution. Here are the 4 acid steps:

1) Balance the atom being reduced/oxidized.
2) Balance the oxygens.
3) Balance the hydrogens.
4) Balance the charge.

Let's use this half reaction for our example. Pretend we just finished steps 1-4: $\mathbf{2e^- + 2H^{1+} + PbO_2 \rightarrow PbO + H_2O}$

Step Five: Convert all H^+ to H_2O.
If you are in a basic solution you will not have H^+ floating around. It is combined with the hydroxides that are present from the base! To convert the H+ into water add the same number of OH^- ions as you have H^+ ions, BUT you have to add them to <u>both</u> sides to preserve the balancing! The side with the H+ will determine how many hydroxide to add. In our case, the left side has 2 hydrogen ions, while the right side has none, so:

$$\mathbf{2e^- + 2H_2O + PbO_2 \rightarrow PbO + H_2O + 2OH^-}$$

Notice that, when the two hydroxide ions on the left were added, they immediately reacted with the hydrogen ion present. The reaction is: $H^+ + OH^- \rightarrow H_2O$

Step Six: Remove any duplicate molecules or ions.
In our example, there are two water molecules on the left and one on the right. This means one water molecule may be removed from each side, giving:

$$\mathbf{2e^- + H_2O + PbO_2 \rightarrow PbO + 2OH^-}$$

The half-reaction is now correctly balanced!

By the way, notice the $2OH^-$. Be careful to read that as two hydroxide ions ($2\ OH^-$) and NOT twenty hydride ions ($2O\ H^{1-}$). People have been known to do that.

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Directions: In each of the following equations do the following on binder paper:

- Balance the reaction in the type of solution indicated.
- If you need/want some extra practice:
  - Indicate the element that has been oxidized and the one that has been reduced.
  - Label the oxidation state of each before and after the process.
  - Some answers have been provided in italics to the side.

Reaction done in Acidic Solution

1) $Re \rightarrow ReO_2$ *($2H_2O + Re \rightarrow ReO_2 + 4H^+ + 4e^-$)*
2) $Cl_2 \rightarrow HClO$ *($2H_2O + Cl_2 \rightarrow 2HClO + 2H^+ + 2e^-$)*
3) $NO_3^- \rightarrow HNO_2$ *($2e^- + 3H^+ + NO_3^- \rightarrow HNO_2 + H_2O$)*
4) $H_2GeO_3 \rightarrow Ge$ *($4e^- + 4H^+ + H_2GeO_3 \rightarrow Ge + 3H_2O$)*
5) $H_2SeO_3 \rightarrow SeO_4^{2-}$
6) $Au \rightarrow Au(OH)_3$
7) $H_3AsO_4 \rightarrow AsH_3$
8) $H_2MoO_4 \rightarrow Mo$
9) $NO \rightarrow NO_3^-$

Reaction done in Basic Solution

10) $NiO_2 \rightarrow Ni(OH)_2$ *($2e^- + 2H_2O + NiO_2 \rightarrow Ni(OH)_2 + 2OH^-$)*
11) $BrO_4^- \rightarrow Br^-$ *($8e^- + 4H_2O + BrO_4^- \rightarrow Br^- + 8OH^-$)*
12) $SbO_3^- \rightarrow SbO_2^-$ *($2e^- + H_2O + SbO_3^- \rightarrow SbO_2^- + 2OH^-$)*
13) $Cu_2O \rightarrow Cu$ *($2e^- + H_2O + Cu_2O \rightarrow 2Cu + 2OH^-$)*
14) $S_2O_3^{2-} \rightarrow SO_3^{2-}$
15) $Tl^+ \rightarrow Tl_2O_3$
16) $Al \rightarrow AlO_2^-$
17) $Sn \rightarrow HSnO_2^-$
18) $CrO_4^{2-} \rightarrow Cr(OH)_3$