Review: Worksheet on Balancing Redox Equations

Two methods are often mentioned for balancing redox reactions: the half reaction method and the change in oxidation method. They actually involve the same procedure. In the first case you separate out the oxidation and reduction half reaction and in the second case, you do it all at once. I prefer the latter. The half reaction method is shown in the text but I will explain the change in oxidation method here.

Change in Oxidation Procedure:

a. Write out as much of the unbalanced reaction as possible
b. Assign oxidation numbers
c. Draw brackets to connect the atoms that are oxidized and the atoms that are reduced. Write the net increases and decreases in electrons.
d. Find the factors that create the least common multiple and use these to assign balanced stoichiometry for the reaction.

Example 1 (no O or H atoms in equation):

a.	$FeCl_2 + SnCl_4 \rightarrow SnCl_2 + FeCl_3$	Write out unbalanced equation
b.	$FeCl_2 + SnCl_4 \rightarrow SnCl_2 + FeCl_3$	Assign oxidation numbers
c.	$FeCl_2 + SnCl_4 \rightarrow SnCl_2 + FeCl_3$	Draw brackets connecting redox atoms

Complication: pH-dependent redox reactions.

Very often we will work with redox reactions that are dependent on the acidity or basicity of a reaction. When this occurs, we need to balance the numbers of O and H atoms that appear in H^+, OH^- and H_2O species in the reaction.

For example, in acid $MnO_4^- \rightarrow Mn^{2+}$ is a 5 electron process
in base $MnO_4^- \rightarrow MnO_2$ is a 3 electron process

Note the change in oxidation number means a different equivalent weight for the MnO_4^- depending on the reaction.

So we need to add an additional step in balancing redox reactions:

e. Rules for acid solution: balance O by adding H_2O, then balance H by adding H^+
Rules for basic solution: for each O, add two OH^- to side needing O and one H_2O to other side
for each H^+, add one H_2O to side needing H^+ and one OH^- to other side.
[Alternatively for base simply first balance in acid, then add enough OH^- to neutralize all the H^+ followed by cancelling out any H_2O that appears on each side of the reaction.]

Example 2. Balancing redox reactions in acid.

A.	$Fe^{2+} + MnO_4^- \rightarrow Mn^{2+} + Fe^{3+}$	Write out unbalanced equation
B.	$Fe^{2+} + MnO_4^- \rightarrow Mn^{2+} + Fe^{3+}$	Assign oxidation numbers
C.	$Fe^{2+} + MnO_4^- \rightarrow Mn^{2+} + Fe^{3+}$	Draw brackets connecting redox atoms
D.	$Fe^{2+} + MnO_4^- \rightarrow Mn^{2+} + Fe^{3+}$	Find common factor, assign stoichiometry
E.	$Fe^{2+} + MnO_4^- \rightarrow Mn^{2+} + Fe^{3+}$	Acidic solution balance of O atoms
F.	Fe2+ + MnO4- → Mn2+ + Fe3+	Acidic solution balance of H atoms

Example 3. Balancing redox in a basic solution

A.	$Fe^{2+} + MnO_4^- \rightarrow MnO_2 + Fe^{3+}$	Write out unbalanced equation
B.	$Fe^{2+} + MnO_4^- \rightarrow MnO_2 + Fe^{3+}$	Assign oxidation numbers
C.	$Fe^{2+} + MnO_4^- \rightarrow MnO_2 + Fe^{3+}$	Draw brackets connecting redox atoms
D.	$Fe^{2+} + MnO_4^- \rightarrow MnO_2 + Fe^{3+}$	Find common factor, assign stoichiometry
E.	$Fe^{2+} + MnO_4^- \rightarrow MnO_2 + Fe^{3+}$	Basic solution balance of O atoms
F.	$Fe^{2+} + MnO_4^- \rightarrow MnO_2 + Fe^{3+}$	Basic solution balance of H atoms

And as always, practice, practice, practice. There will be problems assigned on an upcoming worksheet. While you are doing them, remember that every balanced redox reaction you generate could be an electrochemical cell that has the potential to be a battery.