Worksheet 25 – CHEM 121 – Fall 2015 Monday Name: ________________________

Wednesday Name: _____________________

This worksheet is a little different: there are some typos in your notes for the course regarding balancing redox reactions. This worksheet will clean up the typos starting on page 69 of your notes through page ___ of your notes and contains the problems for you to work to learn the material.

Three Methods of Balancing Redox Reactions

Method 1	Method 2	Method 3
Oxidation Number Method	Oxidation Number Method for Aqueous Solutions	Half-Reactions for Aqueous Solutions Method

EXAMPLE 1

For the first example for ALL three methods (after you get through the homework, I don't care which method you utilize as long as you can do it), we'll use the following UNBALANCED equation:

$$Cr_2O_7^{2-} + HNO_2 + H^+ \rightarrow Cr^{3+} + NO_3^- + H_2O$$

Method 1: Oxidation Number Method

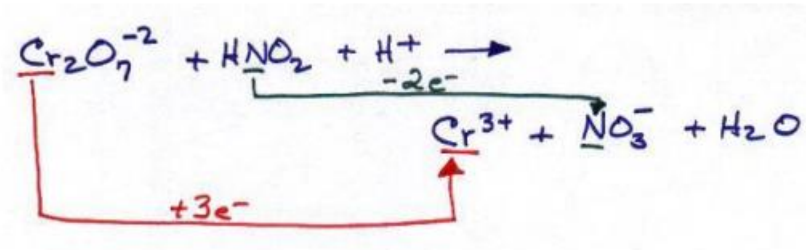
Step Number One: Assign oxidation numbers. Do this just as I set up the rules for oxidation numbers, above, in the multiple tables. Keep track of your charges.

$$(Cr_2O_7)^{-2} = 2Cr + 7O$$
$$-2 = 2x + -14 \quad (7 \cdot -2)$$
$$\frac{-2+14}{2} = x = +6$$

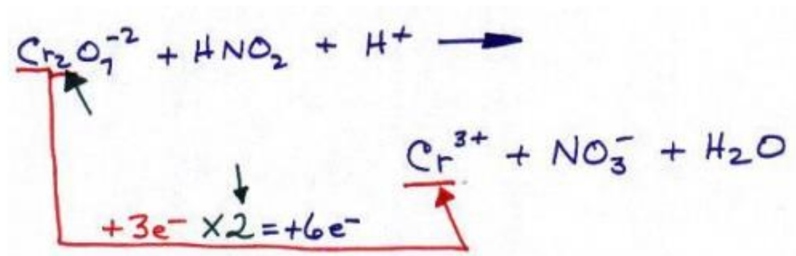
$$HNO_2 = 1H + 1N + 2O$$
$$0 = (1 \cdot 1) + x + (2 \cdot -2)$$
$$0 = 1 + x + (-4)$$
$$0 - 1 + 4 = x = +3$$

$$\overset{+6}{Cr_2}\overset{-2}{O_7}{}^{2-} + \overset{+1}{H}\overset{+3}{N}\overset{-2}{O_2} + H^+ \rightarrow Cr^{3+} + \overset{+5}{N}\overset{-2}{O_3}{}^- + \overset{+1}{H_2}\overset{-2}{O}$$

Step Number Two: You need to determine which reactant gains (oxidizing agent) or loses (reducing agent) electrons and identify how many electrons are lost per individual atom ONLY at this point.



Step Number three: Determine the gain or loss of electrons per formula unit -- this reaction is a great example: dichromate ion ($Cr_2O_7^{2-}$) has 2 Cr's. That means that EACH Cr (formula unit) gains 3 electrons as it's reduced for a total of 6 electrons gained (2 formula units times 3).



Step Number Four: Balance the left side of the reaction for gain and loss of electrons.

$$Cr_2O_7^{2-} + 3HNO_2 + H^+ \longrightarrow Cr^{3+} + NO_3^- + H_2O$$

Step Number Five: Balance the redox pair, now, by balancing the right side of the reaction.

$$Cr_2O_7^{2-} + 3HNO_2 + H^+ \longrightarrow 2Cr^{3+} + 3NO_3^- + H_2O$$

Step Number Six: Balance everything EXCEPT oxygen and hydrogen.

Step Number Seven: Balance the oxygens.

Step Number Eight: Balance the hydrogens.

Step Number Nine: Write the balanced equation/reaction.

$$Cr_2O_7^{-2} + 3HNO_2 + 5H^+ \rightarrow 2Cr^{3+} + 3NO_3^- + 4H_2O$$

Method 2: Oxidation Number Method for Aqueous Solutions

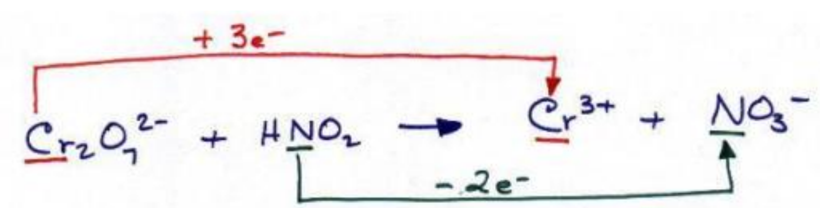
We'll still use the same example reaction, above. This method requires a slightly different approach.

Step Number One: write the net ionic reaction. This means to write the reaction that is going to occur between the redox pair ONLY.

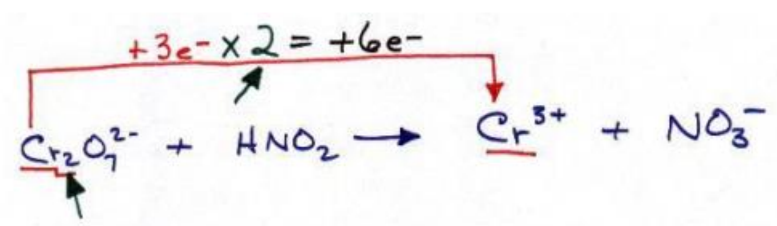
$$\overset{+6}{Cr_2}\overset{-2}{O_7^{2-}} + \overset{+1}{H}\overset{+3}{N}\overset{-2}{O_2} \rightarrow Cr^{3+} + \overset{+5}{N}\overset{-2}{O_3^-}$$

Step Number Two: Assign oxidation numbers just as we did in the first method.

Step Number Three: Determine which reactant gains and loses electrons.



Step Number Four: determine the loss and gain of electrons per formula unit -- this reaction is a great example: dichromate ion ($Cr_2O_7^{2-}$) has 2 Cr's. That means that EACH Cr (formula unit) gains 3 electrons as it's reduced for a total of 6 electrons gained (2 formula units times 3).



Step Number Five: Balance the left side of the reaction based on electron gain/loss.

$$Cr_2O_7^{2-} + 3HNO_2 \rightarrow Cr^{3+} + NO_3^-$$

Step Number Six: balance the right side of the reaction.

$$Cr_2O_7^{2-} + 3HNO_2 \rightarrow 2Cr^{3+} + 3NO_3^-$$

Step Number Seven: balance all but oxygen and hydrogen.

Step Number Eight: Add up the charges on both sides of the reaction.

$$Cr_2O_7^{2-} + 3HNO_2 \rightarrow 2Cr^{3+} + 3NO_3^-$$

$$-2 + 0 = -2 \qquad +3 = +6 - 3$$

Step Number Nine: balance the charges by adding H+ for positive charges and OH^- for negative charges.

$$Cr_2O_7^{2-} + 3HNO_2 + 5H^+ \rightarrow 2Cr^{3+} + 3NO_3^-$$

Step Number Ten: Balance the oxygens by putting water (H_2O) on the side opposite the hydroxide ions. Step Number Eleven: hydrogens ought to balance. Step Number Twelve: Write the balanced reaction.

$$Cr_2O_7^{2-} + 3HNO_2 + 5H^+ \rightarrow 2Cr^{3+} + 3NO_3^- + 4H_2O$$

Method 3: Half-Reactions for Aqueous Solutions Method

Step Number One: write out the unbalanced reaction.

Step Number Two: find the half reactions for each of the redox pair. This is done by looking them up in the appendix of traditional chemistry texts (Redox Potentials) or in the CRC Handbook of Chemistry and Physics. They will be given to you during the exams if they are needed and if they actually exist in the literature -- there are times there are no accessible half reactions for use; when that occurs, these problems are usually fairly simple.

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$
$$HNO_2 + H_2O \rightarrow NO_3^- + 3H^+ + 2e^-$$

Remember: half reactions are written in reduction half reaction form. One of the reactions will HAVE to be reversed so that you have a redox reaction. This is not unlike the thermochemical topics that you've studied previously.

Step Number Three: Balance all but the oxygen and hydrogen in each half reaction.

$$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$$
$$3(HNO_2 + H_2O \rightarrow NO_3^- + 3H^+ + 2e^-)$$

Step Number Four: Balance oxygens with water added to the oxygen deficient side of the reaction.

Step Number Five: Balance the hydrogens: 1) in ACID with H^+ ; 2) in BASE with 1 H_2O per needed hydrogen with an equal number of OH^- on the opposite side of the reaction.

Step Number Six: Balance the charges with electrons on the side of each half reaction with the least negative charge.

Step Number Seven: between the two half reactions, now balance electron gain and loss. In this step, it becomes necessary to turn one of the reactions around so that it will add to the other half-reaction, eventually. Note that two half reactions will make a whole reaction.

$$Cr_2O_7^{2-} + \overset{5}{\cancel{14}}H^+ + \cancel{6e^-} \rightarrow 2Cr^{3+} + \overset{4}{\cancel{7}}H_2O$$
$$3HNO_2 + \cancel{3H_2O} \rightarrow 3NO_3^- + \cancel{9H^+} + \cancel{6e^-}$$

Step Number Eight: Cancel out common elements, ions, and/or electrons between the 2 half reactions now balanced.

Step Number Nine: Add them up and write out the balanced reaction.

$$Cr_2O_7^{2-} + 3HNO_2 + 5H^+ \rightarrow 2Cr^{3+} + 3NO_3^- + 4H_2O$$

One key concept to keep in mind is that if you've balanced the same reaction by each method above, and you don't get the same answer each time, then there is an error in the balancing. Go back and find it -- it usually jumps right out at you.

EXAMPLE 2

Method 1: Oxidation Number Method

For this example for ALL three methods (after you get through the homework, I don't care which method you utilize as long as you can do it), we'll use the following UNBALANCED equation:

$$MnO_4^- + S^{2-} + H_2O \rightarrow MnO_2 \downarrow + S + OH^-$$

Note that the S on the right is in the elemental form and has an oxidation number of zero (0).

Step Number One: Assign oxidation numbers. Do this just as I set up the rules for oxidation numbers, above, in the multiple tables. Keep track of your charges.

$$\begin{aligned} (MnO_4)^{-1} &= 1\,Mn + 4\,O \\ -1 &= x + \underbrace{4(-2)}_{-8} \\ -1 + 8 &= +7 \end{aligned} \qquad \begin{aligned} MnO_2 &= 1\,Mn + 2\,O \\ 0 &= x + \underbrace{2(-2)}_{-4} \\ 0 + 4 &= +4 \end{aligned}$$

$$\overset{+7}{Mn}\overset{-2}{O_4^-} + S^{2-} + \overset{+1}{H_2}\overset{-2}{O} \rightarrow \overset{+4}{Mn}\overset{-2}{O_2}\downarrow + \overset{0}{S} + \overset{-2}{O}\overset{+1}{H^-}$$

Step Number Two: You need to determine which reactant gains (oxidizing agent) or loses (reducing agent) electrons and identify how many electrons are lost per individual atom ONLY at this point. (The sulfide loses 2 electrons and the manganese gains three electrons.)

$$MnO_4^- + S^{2-} + H_2O \rightarrow MnO_2\downarrow + S + OH^-$$

$+3e^-$ $\quad -2e^-$

gain of e^- $\quad$ loss of e^-

Step Number Three: Determine the gain or loss of electrons per formula unit -- this reaction is a bad example as there is only one formula unit for each -- the first of the examples illustrated this idea.

Step Number Four: Balance the left side of the reaction for gain and loss of electrons. This balancing is done by looking at the numbers of electrons gained and lost. They must be equal. Sometimes the easiest thing to do is to just multiply them together if there is no obvious lowest common multiple.

$$2MnO_4^- + 3S^{2-} + H_2O \rightarrow MnO_2\downarrow + S + OH^-$$

Step Number Five: Balance the redox pair, now, by balancing the right side of the reaction.

$$2MnO_4^- + 3S^{2-} + H_2O \rightarrow 2MnO_2\downarrow + 3S + OH^-$$

Make sure you've got as many atoms on the right as you do on the left.

Step Number Six: Balance everything EXCEPT oxygen and hydrogen.

Step Number Seven: Balance the oxygens. See below.

Step Number Eight: Balance the hydrogens. See below.

Step Number Nine: Write the balanced equation/reaction.

$$2MnO_4^- + 3S^{2-} + 4H_2O \longrightarrow 2MnO_2\downarrow + 3S + 8OH^-$$

Method 2: Oxidation Number Method for Aqueous Solutions

We'll still use the same example reaction, above. This method requires a slightly different approach.

Step Number One: write the net ionic reaction. This means to write the reaction that is going to occur between the redox pair ONLY.

$$MnO_4^- + S^{2-} \longrightarrow MnO_2\downarrow + S$$

Step Number Two: Assign oxidation numbers just as we did in the first method.

$$\overset{+7}{Mn}\overset{-2}{O_4^-} + S^{2-} \longrightarrow \overset{+4}{Mn}\overset{-2}{O_2}\downarrow + S^{\circ}$$

Step Number Three: Determine which reactant gains and loses electrons.

$$\underline{Mn}O_4^- + \underline{S}^{2-} \longrightarrow \underline{Mn}O_2\downarrow + \underline{S}$$

$-2e^-$

$+3e^-$

gains e^- loses e^-

Step Number Four: determine the loss and gain of electrons per formula unit (again, this is not a good example and was dealt with in the first example).

Step Number Five: Balance the left side of the reaction based on electron gain/loss.

$$2MnO_4^- + 3S^{2-} \rightarrow MnO_2\downarrow + S$$

Step Number Six: balance the right side of the reaction.

$$2MnO_4^- + 3S^{2-} \rightarrow 2MnO_2\downarrow + 3S$$

Step Number Seven: balance all but oxygen and hydrogen.

Step Number Eight: Add up the charges on both sides of the reaction.

$$2MnO_4^- + 3S^{2-} \rightarrow 2MnO_2\downarrow + 3S$$

$$-2 + -6 = -8 \qquad\qquad 0 = 0 + 0$$

$$(2\cdot -1) \quad (3\cdot -2) \qquad\qquad (2\cdot 0) \quad (3\cdot 0)$$

Step Number Nine: balance the charges by adding H^+ for positive charges and OH^- for negative charges.

$$2MnO_4^- + 3S^{2-} \rightarrow 2MnO_2\downarrow + 3S + 8OH^-$$

Note that I used OH^- to balance the negative charges on the right.

Step Number Ten: Balance the oxygens by putting water (H_2O) on the side opposite the hydroxide ions.

$$2MnO_4^- + 3S^{2-} + 4H_2O \rightarrow 2MnO_2\downarrow + 3S + 8OH^-$$

Step Number Eleven: hydrogens ought to balance.

Step Number Twelve: Write the balanced reaction, just as above.

$$2MnO_4^- + 3S^{2-} + 4H_2O \longrightarrow 2MnO_2\downarrow + 3S + 8OH^-$$

Method 3: Half-Reactions for Aqueous Solutions Method

Step Number One: write out the unbalanced reaction.

Step Number Two: find the half reactions for each of the redox pair. This done by looking them up in the appendix of traditional chemistry texts (Redox Potentials) or in the CRC Handbook of Chemistry and Physics. They will be given to you during the exams if they are needed and if they actually exist in the literature -- there are times there are no accessible half reactions for use; when that occurs, these problems are usually fairly simple. Note that for each reactant there is a half reaction – eventually, we'll add 'em up to make a whole reaction. Remember that half reactions are written in reduction reaction form. When you add them together, one must be reversed to the oxidation half reaction form so you have redox reactions. This is not unlike what you learned in thermochemistry.

$$MnO_4^- + 2H_2O + 3e^- \longrightarrow MnO_2\downarrow + 4OH^-$$

$$S^{2-} \longrightarrow S^\circ + 2e^-$$

Step Number Three: Balance all but the oxygen and hydrogen in each half reaction.

$$2(MnO_4^- + 2H_2O + 3e^- \longrightarrow MnO_2\downarrow + 4OH^-)$$

$$3(S^{2-} \longrightarrow S^\circ + 2e^-)$$

Step Number Four: Balance oxygens with water added to the oxygen deficient side of the reaction.

Step Number Five: Balance the hydrogens: 1) in ACID with H^+ ; 2) in BASE with 1 H_2O per needed hydrogen with an equal number of OH^- on the opposite side of the reaction.

Step Number Six: Balance the charges with electrons on the side of each half reaction with the least negative charge. Note that we multiply the top half reaction by the number of electrons from the bottom reaction and vice versa. While this works in most cases, remember that once the reaction is balanced, you must use the lowest possible coefficients, i.e., if the coefficients are divisible by the same number across the reaction, the final reaction must by divided by that number to yield numbers that are no longer divisible, yet consist of a whole number.

Step Number Seven: between the two half reactions, now balance electron gain and loss. In this step, it becomes necessary to turn one of the reactions around so that it will add to the other half-reaction, eventually. Note that two half reactions will make a whole reaction.

$$2MnO_4^- + 4H_2O + \cancel{6e^-} \rightarrow 2MnO_2\downarrow + 8\ OH^-$$
$$3\ S^{2-} \longrightarrow 3S + \cancel{6e^-}$$

Step Number Eight: Cancel out common elements, ions, and/or electrons between the 2 half reactions now balanced.

Step Number Nine: Add them up and write out the balanced reaction.

$$2MnO_4^- + 3S^{2-} + 4H_2O \rightarrow 2MnO_2\downarrow + 3S + 8OH^-$$

Problems

Table of Half Reactions

IMPORTANT: When necessary, turn the reactions around to fit your needs – do NOT, however, change the contents of the half-reactions

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$
$O_2 + 2H_2O + 2e^- \rightarrow H_2O_2 + 2OH^-$
$Zn^{2+} + 2e^- \rightarrow Zn$
$Cu^{2+} + 2e^- \rightarrow Cu$
$2IO_3^- + 12H^+ + 10\ e^- \rightarrow I_2 + 6H_2O$
$NO_3^- + 4H^+ + 3\ e^- \rightarrow NO + 2H_2O$
$ClO_3^- + 6H^+ + 6\ e^- \rightarrow Cl^- + 3H_2O$
$Cr_2O_7^{2-} + 14H^+ + 6\ e^- \rightarrow 2Cr^{3+} + 7H_2O$
$NO_3^- + 3H^+ + 2\ e^- \rightarrow HNO_2 + H_2O$
$PbO + H_2O + 2\ e^- \rightarrow Pb + 2OH^{-1}$.

1) Balance the following reactions by each of the first two methods of balancing redox reactions:

A) $Zn + NO_3^- \rightarrow Zn^{2+} + N_2\uparrow$

B) $NO_3^- + I_2 \rightarrow IO_3^- + NO_2\uparrow$

C) $Cu + NO_3^- \rightarrow Cu^{2+} + NO_2\uparrow$

D) $H_2O_2 + MnO_4^- \rightarrow Mn^{2+} + O_2\uparrow$

E) $CuS + NO_3^- \rightarrow Cu^{2+} + S + NO\uparrow$

F) $NO_3^- + Zn \rightarrow NH_3\uparrow + Zn(OH)_4^{2-}$

G) $ClO_3^- + I_2 \rightarrow IO_3^- + Cl^-$

H) $Cr_2O_7^{2-} + HNO_2 \rightarrow Cr^{3+} + NO_3^-$

l) $H_2SO_4 + HBr \rightarrow SO_2\uparrow + Br_2\uparrow$

J) $C + HNO_3 \rightarrow NO_2\uparrow + CO_2\uparrow$

2. Balance the following reactions from #1, above, by the third method: D, G and H.

3. Balance the following reactions by whichever methods you so desire:

A) $NO_3^- + Pb \rightarrow NO + PbO$

B) $Cl^- + Zn^{2+} \rightarrow ClO_3^- + Zn$