Unit 8, Lesson 03: Balancing Redox Reactions in Acidic Conditions

These youtube videos by Tyler DeWitt are excellent. Before reading the notes, please watch:

1. How to Balance Redox Equations in Acidic Solution
2. How to Balance Redox Equations in Acidic Solution Example 1
3. How to Balance Redox Equations in Acidic Solution Example 2 (Advanced)

Redox reactions can be extremely complex and difficult to balance. There are many systematic approaches to balancing these reactions. We will use the "half-reaction" method because this seems to be what is used at most universities.

- we usually work from a net ionic equation; the spectator ions do not participate so we ignore them
- water and its ions (H^+ and OH^-) are often involved, so they will be added as needed

eg. Balance this redox reaction: $HSO_3^{1-} + IO_3^{1-} \rightarrow I_2 + SO_4^{2-}$

Step 1: Assign oxidation numbers to all atoms to determine which species are being oxidized and reduced.

- H is +1 and O is –2 (these do not change)
- S changes from +4 in HSO_3^{1-} to +6 in SO_4^{2-} so it is oxidized
- I changes from +4 in IO_3^{1-} to zero in I_2 so it is reduced

Step 2: Divide the net ionic reaction into oxidation and reduction half reactions.

$$HSO_3^{1-} \rightarrow SO_4^{2-}$$

$$IO_3^{1-} \rightarrow I_2$$

Step 3: In the half reactions, balance all atoms EXCEPT oxygen and hydrogen.

$$HSO_3^{1-} \rightarrow SO_4^{2-}$$

$$2\ IO_3^{1-} \rightarrow I_2$$

Step 4: Balance oxygen by adding H_2O to whatever side needs more oxygen.

$$H_2O + HSO_3^{1-} \rightarrow SO_4^{2-}$$

$$2\ IO_3^{1-} \rightarrow I_2 + 6\ H_2O$$

Step 5: Balance hydrogen by adding H+ to whatever side needs more hydrogen.

$$H_2O + HSO_3^{1-} \rightarrow SO_4^{2-} + 3\ H^{1+}$$

$$12\ H^{1+} + 2\ IO_3^{1-} \rightarrow I_2 + 6\ H_2O$$

Step 6: Balance the charges by adding electrons to whatever side needs them.

$$H_2O + HSO_3^{1-} \rightarrow SO_4^{2-} + 3\ H^{1+} + 2\ e^-$$

$$10\ e^- + 12\ H^{1+} + 2\ IO_3^{1-} \rightarrow I_2 + 6\ H_2O$$

Step 7: Multiply the half reactions above by the lowest common multiple (LCM) to make the number of electrons lost in the oxidation half reaction equal to the number of electrons gained in the reduction half reaction.

$$5\ (\ H_2O + HSO_3^{1-} \rightarrow SO_4^{2-} + 3\ H^{1+} + 2\ e^-\ )$$

$$10\ e^- + 12\ H^{1+} + 2\ IO_3^{1-} \rightarrow I_2 + 6\ H_2O$$

Step 8: Multiply (distribute) the LCM multiple(s) through both half reactions then add both half reactions together.

$$5\ H_2O \quad + \quad 5\ HSO_3^{1-} \quad \rightarrow \quad 5\ SO_4^{2-} \quad + \quad 15\ H^{1+} \quad + \quad 10\ e^-$$

$$10\ e^- \quad + \quad 12\ H^{1+} \quad + \quad 2\ IO_3^{1-} \quad \rightarrow \quad I_2 \quad + \quad 6\ H_2O$$

Step 9: Simplify the reaction by canceling out species that appear on both sides of the equation. Please draw a box around your final, balanced equation:

$$\boxed{5\ HSO_3^{1-} \quad + \quad 2\ IO_3^{1-} \quad \rightarrow \quad 5\ SO_4^{2-} \quad + \quad I_2 \quad + \quad 3\ H^{1+} \quad + \quad H_2O}$$

Step 10: Double check that all atoms and charges are correctly balanced.

$$\boxed{5\ HSO_3^{1-} \quad + \quad 2\ IO_3^{1-} \quad \rightarrow \quad 5\ SO_4^{2-} \quad + \quad I_2 \quad + \quad 3\ H^{1+} \quad + \quad H_2O}$$

5 (H)	5 (H)
5 (S)	5 (S)
21 (O)	21 (O)
2 (I)	2 (I)
charge: –7	charge: –7

Summary of Steps:

1. Assign oxidation numbers, if necessary. I do not need to see this work, so you can sometimes go straight to step 2.
2. Write two half reactions.
3. Balance all atoms except oxygen and hydrogen using species from the original reaction (if necessary).
4. Balance oxygen by adding water where needed (if necessary).
5. Balance hydrogen by adding H+ where needed (if necessary).
6. Balance charge by adding electrons where needed.
7. Multiply the half reactions by a lowest common multiple so the number of electrons in each is equal.
8. Add the half reactions together.
9. Simplify.
10. Double check that all atoms and charges are balanced.

Steps 1 to 7 can be done without re-writing the half reactions. It is probably a good idea to re-write the reactions when you multiply (distribute) the LCM through the two half reactions, before you add them together. I also do not need to see your double check, but make sure you do it. I start by making sure the charges are balanced as this is the most common spot to make a mistake. If you have made a mistake, start right back at step 3 and correct every step moving forward from there.

eg. Balance this redox reaction: $Cr_2O_7^{2-} \quad + \quad C_2H_4O \quad \rightarrow \quad HC_2H_3O_2 \quad + \quad Cr^{3+}$

You can find the half reactions quickly: one has the carbon atoms and one has the chromium atoms

$$6\ e^- \quad + \quad 14\ H^{1+} \quad + \quad Cr_2O_7^{2-} \quad \rightarrow \quad 2\ Cr^{3+} \quad + \quad 7\ H_2O$$

$$3\ (\ H_2O \quad + \quad C_2H_4O \quad \rightarrow \quad HC_2H_3O_2 \quad + \quad 2\ H^{1+} \quad + \quad 2\ e^-\ )$$

Re-write both half reaction, and multiply (distribute) through the second half reaction:

$$6\ e^- \quad + \quad 14\ H^{1+} \quad + \quad Cr_2O_7^{2-} \quad \rightarrow \quad 2\ Cr^{3+} \quad + \quad 7\ H_2O$$

$$3\ H_2O \quad + \quad 3\ C_2H_4O \quad \rightarrow \quad 3\ HC_2H_3O_2 \quad + \quad 6\ H^{1+} \quad + \quad 6\ e^-$$

Simplify, add & double-check

$$\boxed{8\ H^{1+} \quad + \quad Cr_2O_7^{2-} \quad + \quad 3\ C_2H_4O \quad \rightarrow \quad 3\ HC_2H_3O_2 \quad + \quad 2\ Cr^{3+} \quad + \quad 4\ H_2O}$$