Balancing Redox Equations

Using the Half-Reaction Method

Modified from Holt Modern Chemistry

Equation: $H_2S + HNO_3 \rightarrow H_2SO_4 + NO_2 + H_2O$

1. Assign oxidation numbers.
2. Determine which 2 elements are changing oxidation #.

- If a compound contains an element that changes oxidation #, write that compound in ionic form. (follow solubility rules)
- In the equation above, sulfur and nitrogen are changing. (sulfur from -2 to +6 and nitrogen from +5 to +4).
- For the time being, we'll disregard the elements that don't change oxidation #. That leaves us:

$H_2S + NO_3^- \rightarrow SO_4^{-2} + NO_2$

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3. **Write the oxidation half reaction.**

$H_2S \rightarrow SO_4^{-2}$

Be sure the element in question (sulfur in this case) is balanced (it already is). Then balance oxygen by adding water. Balance hydrogen by adding H^+ ions (in basic solutions add OH^- instead). We can add water or H^+ because we are assuming we are doing this in an acidic, aqueous solution where these species are readily available (or OH^- would be available if a basic solution)

$H_2S + 4H_2O \rightarrow SO_4^{-2} + 10H^+$

Add electrons to balance the charge. In this case there is a zero charge on the left and a +8 charge on the right, so we must add 8 electrons to the right.

$H_2S + 4H_2O \rightarrow SO_4^{-2} + 10H^+ + 8e^-$

The oxidation half reaction is now balanced. Phew.

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4. **Write the reduction half reaction.**

$NO_3^- \rightarrow NO_2$

Balance the atoms (N, H, and O).

$NO_3^- + 2H^+ \rightarrow NO_2 + H_2O$

Balance the charge (add electrons where needed)

$e^- + NO_3^- + 2H^+ \rightarrow NO_2 + H_2O$

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5. **Multiply half reactions so that the number of electrons lost equals the number gained.**

Ox.: $1(H_2S + 4H_2O \rightarrow SO_4^{-2} + 10H^+ + 8e^-)$

Red.: $8(e^- + NO_3^- + 2H^+ \rightarrow NO_2 + H_2O)$

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6. **Combine (add) the half reactions and cancel anything that is found on both sides**

$$H_2S + 4H_2O \rightarrow SO_4^{-2} + 10H^+ + 8e^-$$

$$8e^- + 8NO_3^- + 16H^+ \rightarrow 8NO_2 + 8H_2O$$

$$8e^- + 8NO_3^- + 16H^+ + H_2S + 4H_2O \rightarrow 8NO_2 + 8H_2O + SO_4^{-2} + 10H^+ + 8e^-$$

After canceling…

$$8NO_3^- + H_2S + 6H^+ \rightarrow 8NO_2 + 4H_2O + SO_4^{-2}$$

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7. **Re-combine the ions to form the compounds found in the original formula. Check to make sure that all other ions balance.**

We need to add 2 H^+ ions to the left so that we can make $8HNO_3$. Since we are adding 2 H^+ to the left, we must also add $2H^+$ to the right.

$$8HNO_3 + H_2S \rightarrow 8NO_2 + 4H_2O + SO_4^{-2} + 2H^+$$

Then we combine the $2H^+$ and SO_4^{-2}

$$8HNO_3 + H_2S \rightarrow 8NO_2 + 4H_2O + H_2SO_4$$

- Now double check to make sure the equation is properly balanced.
- H_2O, H^+, and OH^- can "come out of nowhere". H_2O in any solution, H^+ in an acidic solution, and OH^- in a basic solution.

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Brief Summary of Steps for Acidic Solution

After you write the two half reactions, do the following to each half reaction:

1. Balance element in question
2. Balance oxygen by adding water
3. Balance hydrogen by adding H^+
4. Balance charge by adding e^-
5. Multiply reactions so that e^- are equal
6. Add reactions together, recombine ions, cancel things on both sides, and make the equation look like the original

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Steps for a Basic Solution

After you write the two half reactions, do the following to each half reaction:

1. Follow same steps to balance as if it was acidic (hold off on e^- though)
2. Add OH^- to <u>both sides</u> to neutralize (make water) with any H^+
3. Cancel out any H_2O appearing on both sides
4. Make sure the charges balance (add e^-)
5. Multiply reactions so that e^- are equal
6. Add reactions together, recombine ions, cancel things on both sides, and make the equation look like the original