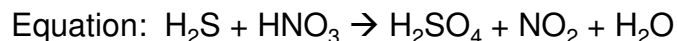
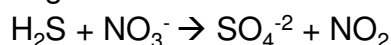


Balancing Redox Equations

Using the Half-Reaction Method Modified from Holt Modern Chemistry



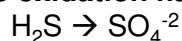
1. Assign oxidation numbers.
2. Determine which 2 elements are changing oxidation #.
 - If a compound contains an element that changes oxidation #, write that compound in ionic form. (follow solubility rules)
 - In the equation above, sulfur and nitrogen are changing. (sulfur from -2 to +6 and nitrogen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



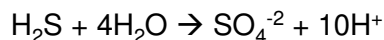
Balancing Redox Equations

original equation: $\text{H}_2\text{S} + \text{HNO}_3 \rightarrow \text{H}_2\text{SO}_4 + \text{NO}_2 + \text{H}_2\text{O}$

3. Write the oxidation half reaction.



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)



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.

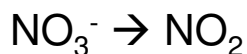


The oxidation half reaction is now balanced. Phew.

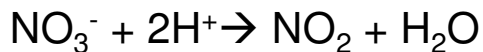
Balancing Redox Equations

original equation: $\text{H}_2\text{S} + \text{HNO}_3 \rightarrow \text{H}_2\text{SO}_4 + \text{NO}_2 + \text{H}_2\text{O}$

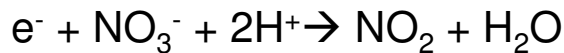
4. Write the reduction half reaction.



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



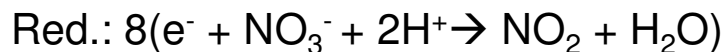
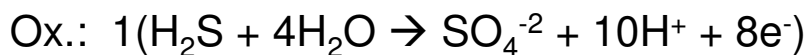
Balance the charge (add electrons where needed)



Balancing Redox Equations

original equation: $\text{H}_2\text{S} + \text{HNO}_3 \rightarrow \text{H}_2\text{SO}_4 + \text{NO}_2 + \text{H}_2\text{O}$

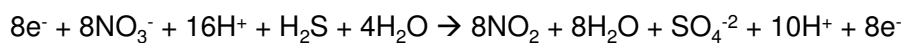
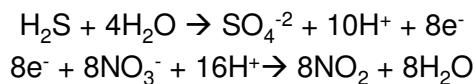
5. Multiply half reactions so that the number of electrons lost equals the number gained.



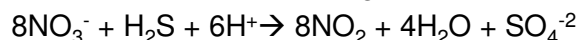
Balancing Redox Equations

original equation: $\text{H}_2\text{S} + \text{HNO}_3 \rightarrow \text{H}_2\text{SO}_4 + \text{NO}_2 + \text{H}_2\text{O}$

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



After canceling...

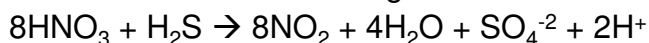


Balancing Redox Equations

original equation: $\text{H}_2\text{S} + \text{HNO}_3 \rightarrow \text{H}_2\text{SO}_4 + \text{NO}_2 + \text{H}_2\text{O}$

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.



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



- 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.

Balancing Redox Equations

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

Balancing Redox Equations

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 both sides 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