

BALANCING OXIDATION-REDUCTION EQUATIONS

As you balance oxidation-reduction equations, keep track along the way of electrons involved. The amount of electrons being transferred must balance in the equation containing both half-reactions. It is important to use a method involving electrons being transferred to guarantee correct balancing of the entire reaction equation.

Method 1: Balancing Oxidation-Reduction Equations

1. Write the separate but not complete half-reactions, making balanced equations for the oxidation and reduction.
2. Balance each half-reaction.
 - a. First balance elements other than O and H.
 - b. Next, balance atoms necessary adding H₂O (if atoms are not balanced by the reaction).
 - c. Then balance the H atoms by adding H⁺ (if atoms are not balanced by the reaction).
 - d. Finally, balance the charge by adding e⁻ to the side with the greater total charge.
3. Multiply each half-reaction by an integer so that the number of electrons lost in one half-reaction equals the number gained in the other.
4. Add the two half-reactions and simplify where possible by cancelling species appearing on both sides of the equation.
5. Check the equation to make sure that there are the same number of atoms of each element on each side and that the charge is the same on both sides.

NOTE: In acidic conditions the final balanced equation must have balanced species and charges.

- a. Balance conditions that should be noted: solutions must have only species specified. This is done by adding H₂O, H⁺, and OH⁻ as appropriate to balance the reaction.

Example 1:

Complete and balance the following equation by the method of half-reactions.



Remember that within the approach procedure always:



Step 1:



Step 2: Notice the number of electrons produced (oxidation) and the number of electrons consumed (reduction).



Step 3: Remember the reduction number is being lost so we double the oxidation number of the oxidation product.



Step 4: Remember the electrons produced must be consumed the electrons.



10. Oxidation: Fe²⁺ to Fe³⁺ and H₂O₂ to H₂O. Total electron difference is 2e⁻.



Step 5: Notice that number of electrons lost (oxidation) and number of electrons gained (reduction).

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