

How to Balance Equations for Oxidation-Reduction Reactions

Oxidation-reduction (redox) reactions are reactions in which oxidation numbers change. *Oxidation numbers* are either real charges or formal charges which help chemists keep track of electron transfer. In practice, oxidation numbers are best viewed as a bookkeeping device.

Oxidation cannot occur without reduction. In a redox reaction the substance which is *oxidized* contains atoms which *increase* in oxidation number. Oxidation is associated with *electron loss* (helpful mnemonic: *LEO = Loss of Electrons, Oxidation*). Conversely, the substance which is *reduced* contains atoms which *decrease* in oxidation number during the reaction. Reduction is associated with *electron gain* (helpful mnemonic: *GER = Gain of Electrons, Reduction*).

Chemists often talk about oxidizing and reducing agents. Be careful with these terms! An *oxidizing agent* is a substance which *oxidizes something else*: it itself is reduced! Also, a *reducing agent* is a substance that *reduces another reactant*: it itself is oxidized. A *disproportionation reaction* is a reaction in which the *same element* is *both* oxidized and reduced.

How to Assign Oxidation Numbers: The Fundamental Rules

Rules for assigning oxidation numbers are as follows:

- The oxidation number of any *pure element* is *zero*. Thus the oxidation number of H in H₂ is zero.
- The oxidation number of a *monatomic ion* is equal to its *charge*. Thus the oxidation number of Cl in the Cl⁻ ion is -1, that for Mg in the Mg⁺² ion is +2, and that for oxygen in O²⁻ ion is -2.
- The *sum* of the oxidation numbers in a compound is *zero if neutral*, or *equal to the charge if an ion*.
- The oxidation number of alkali metals in compounds is +1, and that of alkaline earths in compounds is +2. The oxidation number of F is -1 in all its compounds.
- The oxidation number of *H* is +1 in most compounds. Exceptions are H₂ (where H = 0) and the ionic hydrides, such as NaH (where H = -1).
- The oxidation number of *oxygen (O)* is -2 in most compounds. Exceptions are O₂ (where O = 0) and peroxides, such as H₂O₂ or Na₂O₂, where O = -1.
- For other elements, you can usually use rule (3) to solve for the unknown oxidation number.

Examples:

NO(g) has O = -2, so N = +2.

NO₂(g) has O = -2, so N = +4.

SO₄²⁻ has O = -2. Thus S + 4(-2) = -2. Solving the equation gives S = -2 + 8 = +6.

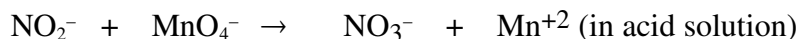
K₂Cr₂O₇ has K = +1 and O = -2. Thus 2(+1) + 2 Cr + 7(-2) = 0; 2 Cr = 12; Cr = +6.

How to Balance Redox Reactions Using the Method of Half-Reactions

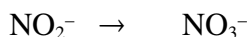
Oxidation-reduction reactions are often tricky to balance without using a systematic method. We shall use the method of half-reactions which is outlined in detail below.

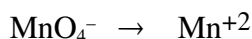
Method in Acidic (or Neutral) Solution

Suppose you are asked to balance the equation below:

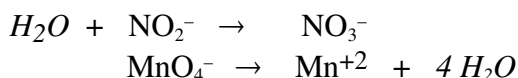


Begin by writing the *unbalanced* oxidation and reduction *half-reactions* (you do not need to know which is which):

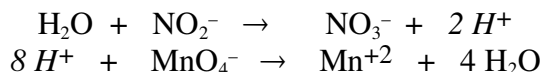




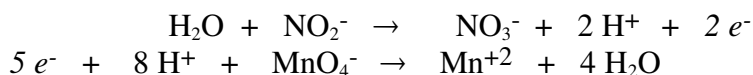
Next, **balance for atoms**. First do this for atoms *other than* O and H. (Both equations above are already balanced for N and Mn, so no change is needed in this example.) Then balance for O atoms by adding H_2O to the reaction side deficient in O:



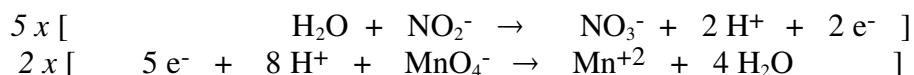
This leaves H atoms unbalanced. In *acidic (or neutral) solution*, balance for H atoms by adding H^+ to the side deficient in H:



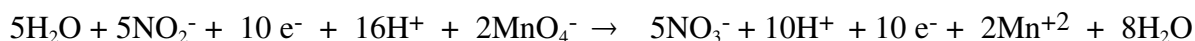
The next step is to **balance for charge**. To do this, add electrons (e^-) to the more positive side:



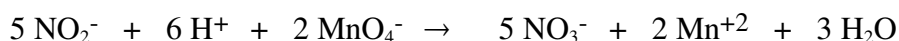
Now you need to **multiply the equations by appropriate factors** so that the number of electrons lost in the oxidation half-reaction (LEO) is equal to the number of electrons gained in the reduction half-reaction (GER):



Then, sum the above equations to obtain



Finally, **simplify** by subtracting out species that are identical on both sides. Our final balanced redox equation is



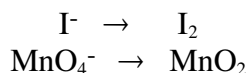
Check this equation to confirm that it is *balanced for atoms* and *balanced for charge*.

Method in Basic Solution

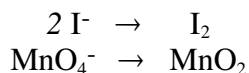
Suppose you are asked to balance the equation below:



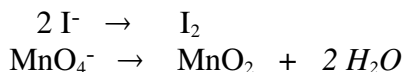
Begin by writing the *unbalanced* oxidation and reduction *half-reactions* (you do not need to know which is which):



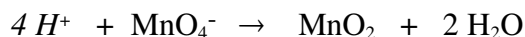
Next, **balance for atoms**. First do this for atoms *other than* O and H:



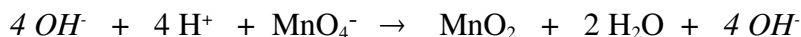
Then balance for O atoms by adding H_2O to the reaction side deficient in O:



This leaves H atoms unbalanced. In *basic solution* (just as in acidic or neutral solution) first balance for H atoms by adding H^+ to the side deficient in H:



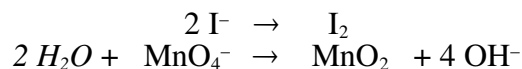
In *basic solution*, follow this step by neutralizing the H^+ ; do this by adding an equivalent amount of OH^- to both sides of the equation.



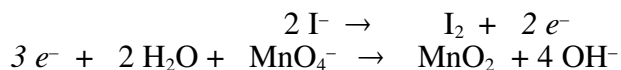
Then form water on the side which has both H^+ and OH^- (recall that $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$): in this case we form 4 H_2O on the left:



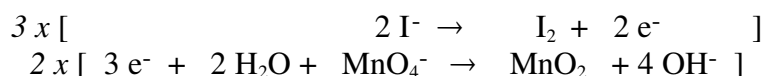
Next simplify the water by subtracting 2 H_2O from both sides. The half-reactions are now:



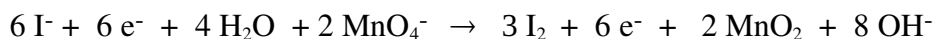
At this point the equations should be balanced for atoms. The next step is to **balance for charge**. To do this, add electrons (e^-) to the more positive side:



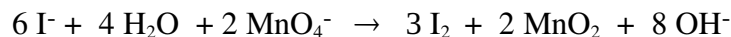
Now you need to **multiply the equations by appropriate factors** so that the number of electrons lost in the oxidation half-reaction (LEO) is equal to the number of electrons gained in the reduction half-reaction (GER):



Sum the equations to obtain



Finally, **simplify** by subtracting out species that are identical on both sides:

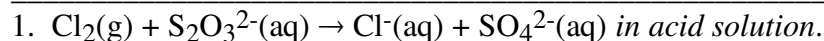


Check our final equation above to confirm that it is *balanced for atoms* and *balanced for charge*.

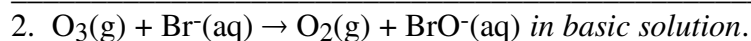
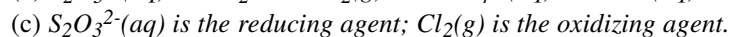
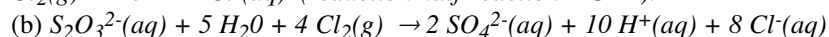
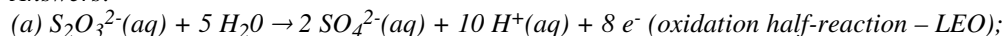
Exercises:

Balance the following redox reactions. In each case

- (a) give the balanced half-reactions; identify the oxidation half-reaction and the reduction half-reaction.
- (b) give the balanced net reaction.
- (c) identify the oxidizing agent and the reducing agent.



Answers:



Answers:

