

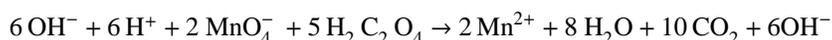
Step 4: Check your work. Once again, you should double-check to be sure your equation is balanced, for atoms and for charge.

This completes big piece two. The result is the balanced full redox equation, for acid solution.

Big piece three: If necessary, convert the equation to basic solution

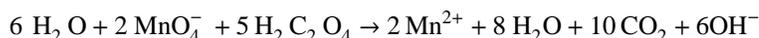
If the redox process is actually carried out in basic solution, then the redox equation needs to be converted to its form in excess OH⁻.

Step 1: Add OH⁻ to both sides of the equation, to neutralize any H⁺

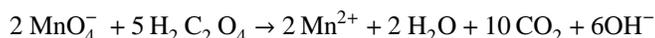


Because OH⁻ is added to both sides of the equation, the equation is still balanced.

Step 2: Combine OH⁻ and H⁺ to form H₂O



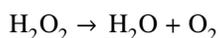
Step 3: Cancel H₂O as possible. There are 6 H₂O on the left and 8 H₂O on the right, for a net of 2 H₂O on the right.



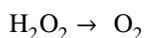
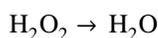
Step 4: Check your work. One last time, make sure that your equation is balanced.

Disproportionation

It is possible for one species to be simultaneously oxidized and reduced. This is called disproportionation. An example is



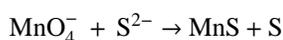
In such a case, one species appears in both the oxidation half-reaction and the reduction half-reaction.



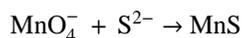
Using oxidation numbers as a guide

A feature of the half-reaction method is that oxidation numbers are not used. Sometimes, though, it is unclear exactly how to form the half-reaction skeletons. In such cases, oxidation number changes make clear what is going on.

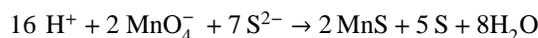
The following problem, 4-50d from Zumdahl, *Chemical Principles*, 2e, is such a case. The unbalanced equation is



The Mn oxidation number changes from +7 (in MnO_4^-) to +2 (in MnS). The sulfur oxidation number changes from -2 (in S^{2-}) to 0 (in S), but it stays the same at -2 in MnS. This means the half-reaction skeletons are



The S^{2-} in the first equation is just a spectator. Using these skeletons, the balanced full equation turns out to be

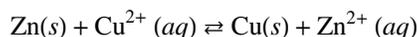


This example shows that the combination of oxidation-number changes and the half-reaction method can be very useful.

■ Harnessing electron flow in redox processes

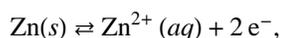
If we place a sheet of zinc in a beaker filled with solution of copper sulfate, immediately the zinc becomes covered with solid copper. After some minutes, so much copper will be formed that it falls to the bottom of the beaker. As the process continues, the copper sulfate solution, originally deep blue, becomes more and more pale.

What is happening is that the copper ions, Cu^{2+} , in the solution are picking up electrons from zinc atoms on the surface of the metal, and so depositing as neutral copper atoms there. The zinc atoms, having lost their electrons to the copper atoms, go into solution as zinc ions, Zn^{2+} . The net result is the redox process

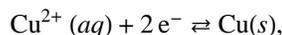


in which electrons are *spontaneously transferred* from zinc to copper.

We can harness this spontaneous transfer of electrons by arranging for the oxidation,



and the reduction,



to take place in separate *half-cells*, connected together as an *electrochemical cell*.

The $\text{Zn} | \text{Zn}^{2+}$ half-cell consists of a zinc electrode in a solution of zinc ions, say from zinc sulfate, and the $\text{Cu}^{2+} | \text{Cu}$ half-cell consists of a copper electrode in a solution of copper ions, say from copper sulfate. If we connect the solutions in the two half-cells by a *salt-bridge* filled with an ionic solution (say, of sodium chloride), then the excess positive charge that builds up in the $\text{Zn} | \text{Zn}^{2+}$ half-cell, due to the formation of zinc ions, can be balanced by flow of chloride ions into the half-cell from the salt bridge; similarly, the excess negative charge that builds up in the $\text{Cu}^{2+} | \text{Cu}$ half-cell, due to the consumption of copper ions, can be balanced by flow of sodium ions into the half-cell from the salt bridge. If we then connect the electrodes with a wire, electrons will flow along the wire from the zinc to the copper.

The spontaneity of the overall redox process is what causes the electrons to flow. A measure of this spontaneity is the voltage between the two electrodes. At 25°C, if the zinc and copper ions are each