4.11: Balancing Redox Equations

Oxidation-Reduction Reactions, or redox reactions, are reactions in which one reactant is oxidized and one reactant is reduced simultaneously. This module demonstrates how to balance various redox equations.

Identifying Redox Reactions

The first step in balancing any redox reaction is determining whether or not it is even an oxidation-reduction reaction. This requires that one and typically more species changing oxidation states during the reaction. To maintain charge neutrality in the sample, the redox reaction will entail both a reduction component and an oxidation components. These are often separated into independent two hypothetical **half-reactions** to aid in understanding the reaction. This requires identifying which element is oxidized and which element is reduced. For example, consider this reaction:

$$\mathrm{Cu(s)} + 2\,\mathrm{Ag^{+}(aq)} \rightarrow \mathrm{Cu^{2+}(aq)} + 2\,\mathrm{Ag(s)}$$

The first step in determining whether the reaction is a redox reaction is to split the equation into two hypothetical *half-reactions*. Let's start with the half-reaction involving the copper atoms:

$$\mathrm{Cu(s)} \rightarrow \mathrm{Cu^{2+}(aq)}$$

The oxidation state of copper on the left side is 0 because it is an element on its own. The oxidation state of copper on the right hand side of the equation is +2. The copper in this half-reaction is oxidized as the oxidation states increases from 0 in Cu to +2 in $\mathrm{Cu^{2+}}$. Now consider the silver atoms

$$2\,\mathrm{Ag^{+}(aq)} \rightarrow 2\,\mathrm{Ag(s)}$$

In this half-reaction, the oxidation state of silver on the left side is a +1. The oxidation state of silver on the right is 0 because it is an pure element. Because the oxidation state of silver decreases from +1 to 0, this is the reduction half-reaction.

Consequently, this reaction is a redox reaction as both reduction and oxidation half-reactions occur (via the transfer of electrons, that are not explicitly shown in equations 2). Once confirmed, it often necessary to balance the reaction (the reaction in equation 1 is balanced already though), which can be accomplished in two ways because the reaction could take place in neutral, acidic or basic conditions.

Balancing Redox Reactions

Balancing redox reactions is slightly more complex than balancing standard reactions, but still follows a relatively simple set of rules. One major difference is the necessity to know the half-reactions of the involved reactants; a half-reaction table is very useful for this. Half-reactions are often useful in that two half reactions can be added to get a total net equation. Although the half-reactions must be known to complete a redox reaction, it is often possible to figure them out without having to use a half-reaction table. This is demonstrated in the acidic and basic solution examples. Besides the general rules for neutral conditions, additional rules must be applied for aqueous reactions in acidic or basic conditions.

One method used to balance redox reactions is called the **Half-Equation Method**. In this method, the equation is separated into two half-equations; one for oxidation and one for reduction.

Half-Equation Method to Balance redox Reactions in **Acidic** Aqueous Solutions

Each reaction is balanced by adjusting coefficients and adding $\mathrm{H_2O}$, $\mathrm{H^+}$, and $\mathrm{e^-}$ in this order:

1. Balance elements in the equation other than O and H.
2. Balance the oxygen atoms by adding the appropriate number of water ($\mathrm{H_2O}$) molecules to the opposite side of the equation.
3. Balance the hydrogen atoms (including those added in step 2 to balance the oxygen atom) by adding $\mathrm{H^+}$ ions to the opposite side of the equation.
4. Add up the charges on each side. Make them equal by adding enough electrons ($\mathrm{e^-}$) to the more positive side. (Rule of thumb: $\mathrm{e^-}$ and $\mathrm{H^+}$ are almost always on the same side.)
5. The $\mathrm{e^-}$ on each side must be made equal; if they are not equal, they must be multiplied by appropriate integers (the lowest common multiple) to be made the same.
6. The half-equations are added together, canceling out the electrons to form one balanced equation. Common terms should also be canceled out.

The equation can now be checked to make sure that it is balanced.

Half-Equation Method to Balance redox Reactions in **Basic** Aqueous Solutions

If the reaction is being balanced in a basic solution, the above steps are modified with the the addition of one step between #3 and #4:

3b Add the appropriate number of OH^- to neutralize all H^+ and to convert into water molecules.

The equation can now be checked to make sure that it is balanced.

Neutral Conditions

The first step to balance any redox reaction is to separate the reaction into half-reactions. The substance being reduced will have electrons as reactants, and the oxidized substance will have electrons as products. (Usually all reactions are written as reduction reactions in half-reaction tables. To switch to oxidation, the whole equation is reversed and the voltage is multiplied by -1.) Sometimes it is necessary to determine which half-reaction will be oxidized and which will be reduced. In this case, whichever half-reaction has a higher reduction potential will by reduced and the other oxidized.

Example 4.11.1: Balancing in a Neutral Solution

Balance the following reaction

$$\mathrm{Cu^{+}(aq) + Fe(s) \rightarrow Fe^{3+}(aq) + Cu(s)}$$

Solution

Step 1: Separate the half-reactions. By searching for the reduction potential, one can find two separate reactions:

$$\mathrm{Cu^{+}(aq) + e^{-} \rightarrow Cu(s)}$$

and

$$\mathrm{Fe^{3+}(aq) + 3\,e^{-} \rightarrow Fe(s)}$$

The copper reaction has a higher potential and thus is being reduced. Iron is being oxidized so the half-reaction should be flipped. This yields:

$$\mathrm{Cu^{+}(aq) + e^{-} \rightarrow Cu(s)}$$

and

$$\mathrm{Fe(s) \rightarrow Fe^{3+}(aq) + 3\,e^{-}}$$

Step 2: Balance the electrons in the equations. In this case, the electrons are simply balanced by multiplying the entire $Cu^{+}(aq) + e^{-} \rightarrow Cu(s)$ half-reaction by 3 and leaving the other half reaction as it is. This gives:

$$\mathrm{3\,Cu^{+}(aq) + 3\,e^{-} \rightarrow 3\,Cu(s)}$$

and

$$\mathrm{Fe(s) \rightarrow Fe^{3+}(aq) + 3\,e^{-}}$$

Step 3: Adding the equations give:

$$\mathrm{3\,Cu^{+}(aq) + 3\,e^{-} + Fe(s) \rightarrow 3\,Cu(s) + Fe^{3+}(aq) + 3\,e^{-}}$$

The electrons cancel out and the balanced equation is left.

$$\mathrm{3\,Cu^{+}(aq) + Fe(s) \rightarrow 3\,Cu(s) + Fe^{3+}(aq)}$$

Acidic Conditions

Acidic conditions usually implies a solution with an excess of H^+ concentration, hence making the solution acidic. The balancing starts by separating the reaction into half-reactions. However, instead of immediately balancing the electrons, balance all the elements in the half-reactions that are not hydrogen and oxygen. Then, add H_2O molecules to balance any oxygen atoms. Next, balance the hydrogen atoms by adding protons (H^+). Now, balance the *charge* by adding electrons and scale the electrons (multiply by the lowest common multiple) so that they will cancel out when added together. Finally, add the two half-reactions and cancel out common terms.

Example 4.11.2: Balancing in a Acid Solution

Balance the following redox reaction in acidic conditions.

$$Cr_2O_7^{2-}(aq) + HNO_2(aq) \rightarrow Cr^{3+}(aq) + NO_3^-(aq)$$

Solution

Step 1: Separate the half-reactions. The table provided does not have acidic or basic half-reactions, so just write out what is known.

$$Cr_2O_7^{2-}(aq) \rightarrow Cr^{3+}(aq)$$

$$HNO_2(aq) \rightarrow NO_3^-(aq)$$

Step 2: Balance elements other than O and H. In this example, only chromium needs to be balanced. This gives:

$$Cr_2O_7^{2-}(aq) \rightarrow 2\,Cr^{3+}(aq)$$

and

$$HNO_2(aq) \rightarrow NO_3^-(aq)$$

Step 3: Add H_2O to balance oxygen. The chromium reaction needs to be balanced by adding 7 H_2O molecules. The other reaction also needs to be balanced by adding one water molecule. This yields:

$$Cr_2O_7^{2-}(aq) \rightarrow 2\,Cr^{3+}(aq) + 7\,H_2O(l)$$

and

$$HNO_2(aq) + H_2O(l) \rightarrow NO_3^-(aq)$$

Step 4: Balance hydrogen by adding protons (H^+). 14 protons need to be added to the left side of the chromium reaction to balance the 14 (2 per water molecule * 7 water molecules) hydrogens. 3 protons need to be added to the right side of the other reaction.

$$14\,H^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow 2\,Cr^{3+}(aq) + 7\,H_2O(l)$$

and

$$HNO_2(aq) + H_2O(l) \rightarrow 3\,H^+(aq) + NO_3^-(aq)$$

Step 5: Balance the charge of each equation with electrons. The chromium reaction has (14+) + (2-) = 12+ on the left side and (2 * 3+) = 6+ on the right side. To balance, add 6 electrons (each with a charge of -1) to the left side:

$$6\,e^- + 14\,H^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow 2\,Cr^{3+}(aq) + 7\,H_2O(l)$$

For the other reaction, there is no charge on the left and a (3+) + (-1) = 2+ charge on the right. So add 2 electrons to the right side:

$$HNO_2(aq) + H_2O(l) \rightarrow 3\,H^+(aq) + NO_3^-(aq) + 2\,e^-$$

Step 6: Scale the reactions so that the electrons are equal. The chromium reaction has $6e^-$ and the other reaction has $2e^-$, so it should be multiplied by 3. This gives:

$$6\,e^- + 14\,H^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow 2\,Cr^{3+}(aq) + 7\,H_2O(l).$$

and

$$3 \times \left[HNO_2(aq) + H_2O(l) \rightarrow 3\,H^+(aq) + NO_3^-(aq) + 2\,e^-\right]$$

$$3\,HNO_2(aq) + 3\,H_2O(l) \rightarrow 9\,H^+(aq) + 3\,NO_3^-(aq) + 6\,e^-$$

Step 7: Add the reactions and cancel out common terms.

$$3\,HNO_2(aq) + 3\,H_2O(l) \rightarrow 9\,H^+(aq) + 3\,NO_3^-(aq) + 6\,e^-$$

$$6\,e^- + 14\,H^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow 2\,Cr^{3+}(aq) + 7\,H_2O(l)$$

$$3\,HNO_2(aq) + \cancel{3\,H_2O(l)} + \cancel{6e^-} + 14\,H^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow 9\,H^+(aq) + 3\,NO_3^-(aq) + \cancel{6e^-} + 2\,Cr^{3+}(aq) + \overset{4}{\cancel{7}}\,H_2O(l)$$

The electrons cancel out as well as 3 water molecules and 9 protons. This leaves the balanced net reaction of:

$$3\,HNO_2(aq) + 5\,H^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow 3\,NO_3^-(aq) + 2\,Cr^{3+}(aq) + 4\,H_2O(l)$$

Basic Conditions

Bases dissolve into OH^- ions in solution; hence, balancing redox reactions in basic conditions requires OH^-. Follow the same steps as for acidic conditions. The only difference is adding hydroxide ions to each side of the net reaction to balance any H^+. OH^- and H^+ ions on the same side of a reaction should be added together to form water. Again, any common terms can be canceled out.

✔ Example 4.11.1: Balancing in Basic Solution

Balance the following redox reaction in basic conditions.

$$Ag(s) + Zn^{2+}(aq) \rightarrow Ag_2O(aq) + Zn(s)$$

Solution

Go through all the same steps as if it was in acidic conditions.

Step 1: Separate the half-reactions.

$$Ag(s) \rightarrow Ag_2O(aq)$$

$$Zn^{2+}(aq) \rightarrow Zn(s)$$

Step 2: Balance elements other than O and H.

$$2\,Ag(s) \rightarrow Ag_2O(aq)$$

$$Zn^{2+}(aq) \rightarrow Zn(s)$$

Step 3: Add H_2O to balance oxygen.

$$H_2O(l) + 2\,Ag(s) \rightarrow Ag_2O(aq)$$

$$Zn^{2+}(aq) \rightarrow Zn(s)$$

Step 4: Balance hydrogen with protons.

$$H_2O(l) + 2\,Ag(s) \rightarrow Ag_2O(aq) + 2\,H^+(aq)$$

$$Zn^{2+}(aq) \rightarrow Zn(s)$$

Step 5: Balance the charge with e^-.

$$H_2O(l) + 2\,Ag(s) \rightarrow Ag_2O(aq) + 2\,H^+(aq) + 2\,e^-$$

$$Zn^{2+}(aq) + 2\,e^- \rightarrow Zn(s)$$

Step 6: Scale the reactions so that they have an equal amount of electrons. In this case, it is already done.

Step 7: Add the reactions and cancel the electrons.

$$H_2O(l) + 2\,Ag(s) + Zn^{2+}(aq) \rightarrow Zn(s) + Ag_2O(aq) + 2\,H^+(aq).$$

Step 8: Add OH^- to balance H^+. There are 2 net protons in this equation, so add 2 OH^- ions to each side.

$$H_2O(l) + 2\,Ag(s) + Zn^{2+}(aq) + 2\,OH^-(aq) \rightarrow Zn(s) + Ag_2O(aq) + 2\,H^+(aq) + 2\,OH^-(aq).$$

Step 9: Combine OH^- ions and H^+ ions that are present on the same side to form water.

$$\cancel{H_2O(l)} + 2\,Ag(s) + Zn^{2+}(aq) + 2\,OH^-(aq) \longrightarrow Zn(s) + Ag_2O(aq) + \cancel{2}H_2O(l)$$

Step 10: Cancel common terms.

$$2\,Ag(s) + Zn^{2+}(aq) + 2\,OH^-(aq) \rightarrow Zn(s) + Ag_2O(aq) + H_2O(l)$$

References

1. Petrucci, Ralph, William Harwood, Geoffrey Herring, and Jeffry Madura. General Chemistry: Principles & Modern Applications. 9th edition. Upper Saddle River, New Jersey: Pearson Prentince Hall, 2007.
2. Helmenstine, Anne Marie. "How to Balance Redox Reactions - Balancing Redox Reactions." *Balancing Redox Reactions - Half-Reaction Method* (2009): n. pag. Web. 1 Dec 2009. http://chemistry.about.com/od/genera...s/redoxbal.htm
3. Stanitski, Conrad L. "Chemical Equations." *Chemistry Explained Foundations and Applications*. 1st. Chemistry Encyclopedia, 2009. Print.
4. "How to Balance Redox Equations." *Youtube*. Web. 1 Dec 2009.

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