Balancing Redox Reactions: Basic Conditions

Problem: Balance the following reaction, and assume it takes place in basic solution:

$$MnO_4^- + S^{2-} \rightarrow S + MnO_2$$

- In the example given, the two reactants are permanganate (MnO_4^-) with sulfide (S^{2-}), and the products are sulfur (S) and manganese(IV) oxide (MnO_2). It is given that the reaction is under basic conditions, so we would expect hydroxide ions to be part of our final reaction.
- The initial set-up for the two reactions is:

$$MnO_4^- \rightarrow MnO_2$$

$$S^{2-} \rightarrow S$$

Sulfide Half-Reaction (oxidation): $S^{2-} \rightarrow S$

- Because the charge goes from -2 to 0, sulfide is being oxidized. Thus, this is your oxidation half reaction.
- First, balance the charges by adding electrons (e-) to either side of the reaction. Because the product side has a +2 charge greater than the reactant side, you would add 2 electrons to the product side. Both sides are now balanced in terms of charges, with each side having a -2 charge.

$$S^{2-} \rightarrow S + \underline{2e-}$$

  - **If electrons appear on the left side of the reaction, sulfide, or whichever reactant you are dealing with, is undergoing reduction and is thus an oxidizing agent. If electrons appear on the right side, sulfide is undergoing oxidation and is thus a reducing agent.**
  - In this case, since the electrons appear on the right side of the reaction, the sulfide ion is undergoing oxidation and is thus the reducing agent.

<u>Permanganate Half-Reaction (reduction):</u> $MnO_4^- \rightarrow MnO_2$

- Because the charge of manganese is going from +7 [x + 4(-2) = -1] to +4 [x + 2(-2) = 0], manganese is being reduced. This is the reduction half-reaction.
- Because there are two oxygen atoms in MnO_4^- that are not present in the final MnO_2 form, we would add water to the product side. **To balance the oxygens, we add two water molecules, one for each oxygen atom needed, to the side that needs oxygen:**

$$MnO_4^- \rightarrow MnO_2 + \underline{2H_2O}$$

- The left side of the equation is now lacking 4 hydrogen atoms, so we now need to balance the hydrogens. **Since the reaction is in basic conditions, we add the necessary number of H_2O molecules, one for each H atom needed, to the side that needs hydrogen, and the same number of OH- ions to the opposite side.** Then cancel repetitive species that appear on both sides of the reaction--in this case, $2H_2O$.

$$MnO_4^- + \underline{4H_2O} \rightarrow MnO_2 + 2H_2O + \underline{4OH^-}$$

$$MnO_4^- + \underline{2H_2O} \rightarrow MnO_2 + 4OH^-$$

- Next, we must balance the charges by adding electrons. The left side of the reaction only has a net charge of -1 (from the MnO_4^- ion) while the right side has a net charge of -4 (from the 4 OH- ions). Since this is a reduction half-reaction we add 3 electrons to the left side so that both sides have a net charge of -4.

$$MnO_4^- + 2H_2O + \underline{3e^-} \rightarrow MnO_2 + 4OH^-$$

  - Since the electrons appear on the left side of the reaction, permanganate is undergoing reduction and is thus the oxidizing agent.

Combining the Half-Reactions

- Now we have to combine the two half-reactions and balance them.

$$MnO_4^- + 2H_2O + 3e- \rightarrow MnO_2 + 4OH-$$
$$S^{2-} \rightarrow S + 2e-$$

- To cancel out compounds, they must be on opposite sides. **To balance a redox reaction, the electrons must also cancel out.**

- For the electrons to cancel, they must be on opposite sides, and each side must have the same number of electrons. The first half-reaction has 3 electrons while the second one has 2 electrons. The least common multiple of 3 and 2 is 6, so to match the number of electrons we must multiply the first reaction by 2 and the second reaction by 3.

$$2(MnO_4^- + 2H_2O + 3e- \rightarrow MnO_2 + 4OH-) \qquad 3(S^{2-} \rightarrow S + 2e-)$$
$$2MnO_4^- + 4H_2O + 6e- \rightarrow 2MnO_2 + 8OH- \qquad 3 S^{2-} \rightarrow 3S + 6e-$$

- When we add the two half-reactions the electrons cancel. Since there are no other species that cancel, we can write down the rest of the reactants and products to get our overall reaction.

Reduction Half-Reaction: $2MnO_4^- + 4H2O + \cancel{6e-} \rightarrow 2MnO_2 + 8OH-$

Oxidation Half-Reaction: $3S^{2-} \rightarrow 3S + \cancel{6e-}$

Overall Reaction: $\mathbf{2MnO_4^- + 4H_2O + 3 S^{2-} \rightarrow 2MnO_2 + 8OH- + 3S}$

This redox reaction has all atoms and charge balanced.