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Balancing Redox Reactions: Acidic Conditions

- In the example given, the two reactants are permanganate (MnO_4^-) and iron (Fe^{2+}), and the products are manganese (Mn^{2+}) and iron (Fe^{3+}). It is given that the reaction is occurring under acidic conditions, so we add H^+ ions to the left side and water on the right side to balance the oxygen atoms that are present.

- The initial set-up of the two half reactions is:

$$MnO_4^- \rightarrow Mn^{2+}$$

$$Fe^{2+} \rightarrow Fe^{3+}$$

Iron half-reaction (oxidation): $Fe^{2+} \rightarrow Fe^{3+}$

- Because the charge goes from +2 to +3, iron 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 +1 charge greater than the reactant side, we add one electron to the product side. Both sides are now balanced with a +2 charge.

$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

* Note: **If electrons appear on the left side of this reaction, iron is undergoing reduction and is thus an oxidizing agent. If electrons appear on the right side, iron is undergoing oxidation and must be a reducing agent.**

Permanganate half-reaction (reduction): $MnO_4^- \rightarrow Mn^{2+}$

- Because the charge goes from +7 $[x + 4(-2) = -1]$ to +2, manganese is being reduced. This is the reduction half reaction.

- Because there are four oxygen atoms in MnO_4^- that are not present in the final Mn^{2+} form, we must add water to the product side. We add four waters to the side opposite of the four oxygen atoms to balance the oxygens:

$$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$$

- The left side of the equation is now lacking hydrogen atoms. **Since it is acidic conditions, we add the necessary number of H^+ ions** on the side opposite of the waters (the reactant side) to balance hydrogens. We must add 8 H^+ ions to balance:

$$MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$$

- Next, we must balance the charges by adding electrons. The left side of the reaction has a net charge of +7 (-1 and +8) while the right side only has +2 from the Mn^{2+}. Since this is the reduction half-reaction we add 5 negatively charged electrons to the left side so both sides have a +2 charge.

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

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

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$$

$$Fe^{2+} \rightarrow Fe^{3+} + e^-$$

- 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 (as seen above), and they must have the same number of electrons. The first reaction has 5 electrons while the second reaction has only 1. Therefore, we must multiply the entire second equation by 5 to balance those electrons.

$$5\ (Fe^{2+} \rightarrow Fe^{3+} + e^-)$$

$$5Fe^{2+} \rightarrow 5Fe^{3+} + 5e^-$$

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

Reduction half reaction: $MnO_4^- + 8H^+ + \cancel{5e^-} \rightarrow Mn^{2+} + 4H_2O$

Oxidation half reaction: $5Fe^{2+} \rightarrow 5Fe^{3+} + \cancel{5e^-}$

Overall reaction: $\mathbf{MnO_4^- + 8H^+ + 5Fe^{2+} \rightarrow Mn^{2+} + 4H_2O + 5Fe^{3+}}$

This redox reaction has all atoms and charge balanced.