

Electrochemistry

Review Oxidation-Reduction (redox) reactions, assigning oxidation numbers.

****LEO goes GER**



e⁻ transferred from Fe to H⁺

Oxidizing agent (oxidant) = e⁻ taker (gets reduced)

Reducing agent (reductant) = e⁻ giver (gets oxidized)

Half Reactions



Balancing equations by the half reaction method.

Often, all of the ions present at the beginning of a redox will not be represented, due to the fact that they are spectators and are not important to the redox process.

Also, redox rxns may be accompanied by the terms "in acid solution or acidic environment" or "in base or basic environment."

➤ If acid, use H⁺ and H₂O to balance the equation.

➤ If base, use OH⁻ and H₂O to balance the equation.

Acidic env. $\text{MnO}_4^-(\text{aq}) + \text{C}_2\text{O}_4^{2-}(\text{aq}) \rightarrow \text{Mn}^{2+}(\text{aq}) + \text{CO}_2(\text{g})$

Step 1: Write basic half reactions.



Step 2: Each half rxn is balanced with respect to the

a. elements oxidized and reduced and

b. other atoms

c. charge

Reduction: $8\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$ (atoms now balance

net chg +7

net chg +2 using H⁺ and H₂O)

The e⁻ change is 5.

So: $5e^- + 8\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$

Oxidation: $\text{C}_2\text{O}_4^{2-} \rightarrow 2\text{CO}_2$ (atoms balance)

net chg -2

net chg 0

The e⁻ change is 2.

So: $\text{C}_2\text{O}_4^{2-} \rightarrow 2\text{CO}_2 + 2e^-$

Step 3: Balance e⁻ in the 2 half rxns by multiplying by correct values to give a common factor.

e⁻ lost must = e⁻ gained.



Oxidation Reduction And Electrochemistry Section Review

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