

Redox Stoichiometry (Titrations)

- In a redox titration, a known concentration of an oxidizing agent is used to find an unknown concentration of a reducing agent, or vice-versa.
- Applications include finding the iron content of water or the Vitamin C content in foods.

General Steps:

1. Balance the redox equation
2. Organize data for the given and required substances
3. Find the moles of the given substance
If given volume and concentration, use $n = Cv$
If given mass, find molar mass, then use $n = m/M$
4. Find the moles of the required substances, using mole ratio
5. Convert moles to concentration ($C = n/v$) or volume ($v = n/C$) or mass ($m = nM$) of required substance

Example 1:

The following example involves the titration of a solution of sodium oxalate, $\text{Na}_2\text{C}_2\text{O}_4(\text{aq})$ with a solution of potassium permanganate, $\text{KMnO}_4(\text{aq})$.

Important Note:

- The permanganate ion, MnO_4^- , is often used as an oxidizing agent because of its deep purple color.
- It acts as its own indicator.
- It is reduced to pale pink Mn^{2+} by the oxalate ions.
- Once all oxalate ions have been oxidized the next drop of permanganate ions turns the solution a faint purple color, signaling the endpoint of the titration.

Find the **volume** of 0.200 M solution of MnO_4^- which will react with 50.0 mL of 0.100 M solution of $\text{C}_2\text{O}_4^{2-}$.



Example 2:

A 23.30 ml sample of KMnO_4 solution is decolorized by 0.1111g of oxalic acid ($\text{H}_2\text{C}_2\text{O}_4$). The products are Mn^{2+} and CO_2 gas. Calculate the concentration of KMnO_4 using the following equation:

Example 3:

The oxidation of Fe^{2+} ions by permanganate in acidic solution is accompanied by the formation of Fe^{3+} ions and Mn^{2+} ions. How many moles of FeSO_4 would be oxidized by 100.0 ml of 0.02118 mol/L KMnO_4 solution.



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