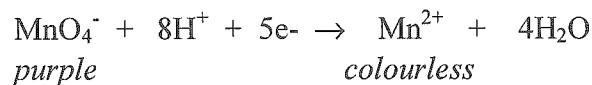


4. Redox Titrations

Find an unknown concentration by slowly adding a known concentration until the equivalence point.

a) Find Concentration of a Reducing Agent

- i) **Procedure:** 1. Titrate the reducing agent with KMnO_4 (powerful oxidizing agent)
2. The reducing agent will change MnO_4^- into Mn^{2+} :

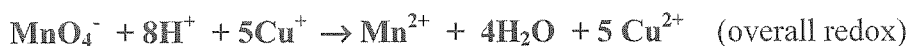
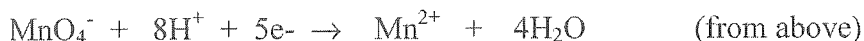


3. Equivalence point is when you see the first hint of purple in the flask.
There is no longer any reducing agent left to convert the purple MnO_4^- into colourless Mn^{2+}

ii) Calculation Sample (very similar to acid-base titration)

25.00 ml of unknown $[\text{Cu}^+]$ is titrated to equivalence point with 18.73 ml of 0.200 M KMnO_4 solution. What is the $[\text{Cu}^+]$?

1st write the overall reaction and balance



2nd Find moles MnO_4^-

$$\text{moles} = \text{M} \times \text{L} = (0.200 \text{ M})(0.01873 \text{ L}) = 0.003746 \text{ mol } \text{MnO}_4^-$$

3rd Convert to moles Cu^+

$$0.003746 \text{ moles } \text{MnO}_4^- \times \frac{5 \text{ mol } \text{Cu}^+}{1 \text{ mol } \text{MnO}_4^-} = 0.01873 \text{ mol } \text{Cu}^+$$

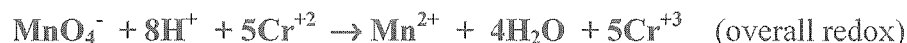
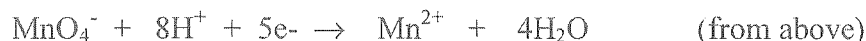
4th Find $[\text{Cu}^+]$

$$[\text{Cu}^+] = \text{mol/L} = 0.01873 \text{ mol} / 0.02500 \text{ L} = \mathbf{0.749 \text{ M}}$$

iii) Calculation Example

25.00 ml of unknown $[\text{Cr}^{+2}]$ is titrated to equivalence point with 28.45 ml of 0.150 M KMnO_4 solution. What is the $[\text{Cr}^{+2}]$?

1st write the overall reaction and balance



2nd Find moles MnO_4^-

$$\text{moles} = \text{M} \times \text{L} = (0.150 \text{ M})(0.02845 \text{ L}) = 0.004268 \text{ mol } \text{MnO}_4^-$$

3rd Convert to moles Cr^{+2}

$$0.004268 \text{ moles } \text{MnO}_4^- \times \frac{5 \text{ mol } \text{Cr}^{+2}}{1 \text{ mol } \text{MnO}_4^-} = 0.02134 \text{ mol } \text{Cr}^{+2}$$

4th Find $[\text{Cr}^{+2}]$

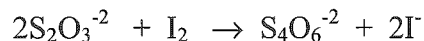
$$[\text{Cr}^{+2}] = \text{mol/L} = 0.02134 \text{ mol} / 0.02500 \text{ L} = \mathbf{0.854 \text{ M}}$$

b) Find Concentration of an Oxidizing Agent

- i) Procedure:**
1. Titrate the oxidizing agent with I^-
 2. The oxidizing agent will change I^- into I_2 :



3. It is difficult to detect the disappearance of I^-
4. We then titrate the I_2 with $\text{Na}_2\text{S}_2\text{O}_3$



5. Use starch as the indicator. Starch is blue in presence of I_2 !
6. Equivalence point is when you see the blue colour in the flask disappear. There is no longer any I_2 left to "blue" the starch.
7. If know moles of $\text{S}_2\text{O}_3^{2-}$, we can find moles of I_2 . If know moles of I_2 , can find moles of oxidizing agent from balanced redox reaction.

ii) Calculation Sample

A 20.00 ml sample of NaOCl is reacted with NaI according to the reaction:



What is the [OCl⁻] if the I₂ produced required 34.77 mL of 0.500 M Na₂S₂O₃ to reach the equivalence point?

1st Find moles of S₂O₃⁻²

$$\text{moles} = M \times L = (0.500 \text{ M})(0.03477 \text{ L}) = 0.01738 \text{ mol S}_2\text{O}_3^{-2}$$

2nd Convert to moles I₂ (using 2S₂O₃⁻² + I₂ → S₄O₆⁻² + 2I⁻)

$$0.01738 \text{ moles S}_2\text{O}_3^{-2} \times \frac{1 \text{ mol I}_2}{2 \text{ mol S}_2\text{O}_3^{-2}} = 0.00869 \text{ mol I}_2$$

3rd Conver to moles OCl⁻

$$0.00869 \text{ mol I}_2 \times \frac{1 \text{ mol OCl}^-}{1 \text{ mol I}_2} = 0.00869 \text{ mol OCl}^-$$

4th Find [OCl⁻]

$$[\text{OCl}^-] = \text{mol} / \text{L} = 0.00869 \text{ mol} / 0.02000 \text{ L} = \mathbf{0.434 \text{ M}}$$

iii) Calculation Example

A 15.00 ml sample of NaOCl is reacted with KI according to the reaction:



What is the [OCl⁻] if the I₂ produced required 19.50 mL of 0.240 M Na₂S₂O₃ to reach the equivalence point?

$$\text{moles S}_2\text{O}_3^{-2} = M \times L = (0.240 \text{ M})(0.01950 \text{ L}) = 0.00468 \text{ mol S}_2\text{O}_3^{-2}$$

$$0.00468 \text{ moles S}_2\text{O}_3^{-2} \times \frac{1 \text{ mol I}_2}{2 \text{ mol S}_2\text{O}_3^{-2}} = 0.00234 \text{ mol I}_2$$

$$0.00234 \text{ mol I}_2 \times \frac{1 \text{ mol OCl}^-}{1 \text{ mol I}_2} = 0.00234 \text{ mol OCl}^-$$

$$[\text{OCl}^-] = \text{mol} / \text{L} = 0.00234 \text{ mol} / 0.01500 \text{ L} = \mathbf{0.156 \text{ M}}$$