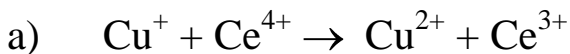


## Chapter 15 HW Problems

15-2, 15-5, 15-11, 15-14, 15-15, 15-18, 15-19, 15-24, 15-26, 15-27, 15-29

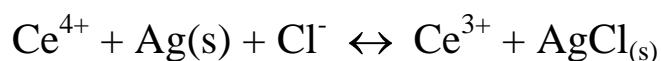
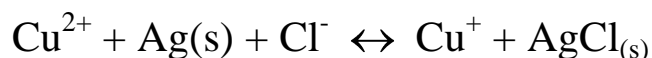
15-2



$$E^0_{(\text{Cu}^+/\text{Cu}^{2+})} = 0.161 \text{ V}$$

$$E^0_{(\text{Ce}^{3+}/\text{Ce}^{4+})} = 1.70 \text{ V}$$

b)



c)

$$E_{\text{cell}} = E^0_{(\text{Cu}^+/\text{Cu}^{2+})} - E_{(\text{Ag}/\text{AgCl}_{\text{ref}})} - 0.05916 \log\left(\frac{[\text{Cu}^+]}{[\text{Cu}^{2+}]}\right)$$

Or

$$E_{\text{cell}} = E^0_{(\text{Ce}^{3+}/\text{Ce}^{4+})} - E_{(\text{Ag}/\text{AgCl}_{\text{ref}})} - 0.05916 \log\left(\frac{[\text{Ce}^{3+}]}{[\text{Ce}^{4+}]}\right)$$

d)

It will take 25.0 mL to reach the equivalence point ( $M_1V_1 = M_2V_2$ )  
at 1.00 mL of  $\text{Cu}^+$  added

$$\text{mmol Ce}^{4+} =$$

$$(100 \text{ mL})(0.0100 \text{ M}) - (1.00 \text{ mL})(0.0400 \text{ M}) = 0.960 \text{ mmol}$$

$$\text{mmol Ce}^{3+} = (1.00 \text{ mL})(0.0400 \text{ M}) = 0.040 \text{ mmol}$$

$$E_{\text{cell}} = E^0_{(\text{Ce}^{3+}/\text{Ce}^{4+})} - E_{(\text{Ag}/\text{AgCl}_{\text{ref}})} - 0.05916 \log\left(\frac{[\text{Ce}^{3+}]}{[\text{Ce}^{4+}]}\right)$$

$$= 1.70 - 0.197 - 0.05916 \log(.04/.96)$$

$$= 1.58 \text{ V}$$

at 12.50 mL of  $\text{Cu}^+$  added (1/2 way!!!!)

$$\text{mmol Ce}^{4+} =$$

$$(100 \text{ mL})(0.0100 \text{ M}) - (12.50 \text{ mL})(0.0400 \text{ M}) = 0.500 \text{ mmol}$$

$$\text{mmol Ce}^{3+} = (12.50 \text{ mL})(0.0400 \text{ M}) = 0.500 \text{ mmol}$$

$$E_{\text{cell}} = E^0_{(\text{Ce}^{3+}/\text{Ce}^{4+})} - E_{(\text{Ag}/\text{AgCl}_{\text{ref}})} - 0.05916 \log\left(\frac{[\text{Ce}^{3+}]}{[\text{Ce}^{4+}]}\right)$$

$$= 1.70 - 0.197 - 0.05916 \log(0.500/0.500)$$

$$= 1.503 \text{ V}$$

at 25.0 mL of  $\text{Cu}^+$  added (at equiv. pt)

$$\begin{aligned} E_{\text{cell}} &= \{ (E^0_{(\text{Cu}^+/\text{Cu}^{2+})} + E^0_{(\text{Ce}^{3+}/\text{Ce}^{4+})}) / 2 \} - 0.196 \text{ V} = \\ &= (0.161 + 1.70)/2 - 0.196 \text{ V} \\ &= 0.734 \text{ V} \end{aligned}$$

at 25.50 mL of  $\text{Cu}^+$  added

mmol  $\text{Cu}^+$  =

$$(0.50 \text{ mL})(0.0400 \text{ M}) = 0.020 \text{ mmol}$$

mmol  $\text{Cu}^{2+}$  = (100 ml)(0.0100 M) = 1.00 mmol

$$\begin{aligned} E_{\text{cell}} &= E^0_{(\text{Cu}^+/\text{Cu}^{2+})} - E_{(\text{Ag}/\text{AgCl}_{\text{ref}})} - 0.05916 \log([\text{Cu}^+]/[\text{Cu}^{2+}]) \\ &= 0.161 - 0.197 - 0.05916 \log(0.02/1.00) \\ &= 0.065 \text{ V} \end{aligned}$$

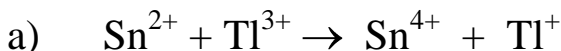
mmol  $\text{Cu}^+$  =

$$(25.00 \text{ ml})(0.0400 \text{ M}) = 1.00 \text{ mmol}$$

mmol  $\text{Cu}^{2+}$  = (100 ml)(0.0100 M) = 1.00 mmol

$$\begin{aligned} E_{\text{cell}} &= E^0_{(\text{Cu}^+/\text{Cu}^{2+})} - E_{(\text{Ag}/\text{AgCl}_{\text{ref}})} - 0.05916 \log([\text{Cu}^+]/[\text{Cu}^{2+}]) \\ &= 0.161 - 0.197 - 0.05916 \log(1.00/1.00) \\ &= -0.036 \text{ V} \end{aligned}$$

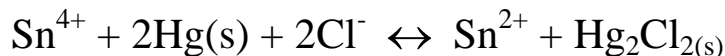
15.3



$$E^0_{(\text{Tl}^+/\text{Tl}^{3+})} = 0.77 \text{ V}$$

$$E^0_{(\text{Sn}^{2+}/\text{Sn}^{4+})} = 0.139 \text{ V}$$

b)



c)

$$E_{\text{cell}} = E^0_{(\text{Tl}^+/\text{Tl}^{3+})} - E_{(\text{SCE})} - 0.05916/2 \log([\text{Tl}^+]/[\text{Tl}^{3+}])$$

Or

$$E_{\text{cell}} = E^0_{(\text{Sn}^{2+}/\text{Sn}^{4+})} - E_{(\text{SCE})} - 0.05916 \log([\text{Sn}^{2+}]/[\text{Sn}^{4+}])$$

d)

It will take 5.00 mL to reach the equivalence point ( $M_1V_1 = M_2V_2$ )

at 1.00 mL of  $\text{Tl}^{3+}$  added

mmol  $\text{Sn}^{2+} =$

$$(25.0 \text{ mL})(0.0100 \text{ M}) - (1.00 \text{ mL})(0.0500 \text{ M}) = 0.200 \text{ mmol}$$

mmol  $\text{Sn}^{4+} = (1.00 \text{ mL})(0.0500 \text{ M}) = 0.0500 \text{ mmol}$

$$\begin{aligned} E_{\text{cell}} &= E^0_{(\text{Sn}^{2+}/\text{Sn}^{4+})} - E_{(\text{SCE})} - 0.05916/2 \log([\text{Sn}^{2+}]/[\text{Sn}^{4+}]) \\ &= 0.139 - 0.241 - 0.05916/2 \log(0.200/0.0500) \\ &= -0.120 \text{ V} \end{aligned}$$

at 2.50 mL of  $\text{Tl}^{3+}$  added (1/2 way!!!!)

mmol  $\text{Sn}^{2+} =$

$$(25.0 \text{ mL})(0.0100 \text{ M}) - (2.50 \text{ mL})(0.0500 \text{ M}) = 0.125 \text{ mmol}$$

mmol  $\text{Sn}^{4+} = (2.50 \text{ mL})(0.0500 \text{ M}) = 0.125 \text{ mmol}$

$$\begin{aligned} E_{\text{cell}} &= E^0_{(\text{Sn}^{2+}/\text{Sn}^{4+})} - E_{(\text{SCE})} - 0.05916/2 \log([\text{Sn}^{2+}]/[\text{Sn}^{4+}]) \\ &= 0.139 - 0.241 - 0.05916/2 \log(0.125/0.125) \\ &= -0.102 \text{ V} \end{aligned}$$

at 5.00 mL of  $\text{Tl}^{3+}$  added (at equiv. pt)

$$\begin{aligned} E_{\text{cell}} &= \{ (E^0_{(\text{Sn}^{2+}/\text{Sn}^{4+})} + E^0_{(\text{Tl}^+/\text{Tl}^{3+})}) / 2 \} - 0.241 \text{ V} = \\ &= (0.139 + 0.77)/2 - 0.241 \text{ V} \\ &= 0.21 \text{ V} \end{aligned}$$

at 5.10 mL of  $\text{Tl}^{3+}$  added

mmol  $\text{Tl}^{3+} =$

$$(0.10 \text{ mL})(0.0500 \text{ M}) = 0.0050 \text{ mmol}$$

mmol  $\text{Tl}^+ = (25.0 \text{ mL})(0.0100 \text{ M}) = 0.250 \text{ mmol}$

$$\begin{aligned} E_{\text{cell}} &= E^0_{(\text{Tl}^+/\text{Tl}^{3+})} - E_{(\text{SCE})} - 0.05916/2 \log([\text{Tl}^+]/[\text{Tl}^{3+}]) \\ &= 0.77 - 0.241 - 0.05916/2 \log(0.250/0.005) \\ &= 0.49 \text{ V} \end{aligned}$$

At 10.00 mL of  $\text{Tl}^{3+}$  added

mmol  $\text{Tl}^{3+} =$

$$(5.00 \text{ mL})(0.0500 \text{ M}) = 0.250 \text{ mmol}$$

mmol  $\text{Tl}^+ = (25.0 \text{ mL})(0.0100 \text{ M}) = 0.250 \text{ mmol}$

$$\begin{aligned} E_{\text{cell}} &= E^0_{(\text{Tl}^+/\text{Tl}^{3+})} - E_{(\text{SCE})} - 0.05916/2 \log([\text{Tl}^+]/[\text{Tl}^{3+}]) \\ &= 0.77 - 0.241 - 0.05916/2 \log(0.250/0.250) \\ &= 0.53 \text{ V} \end{aligned}$$

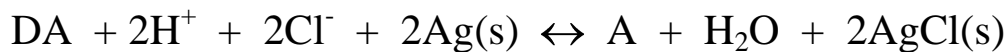
15-4

a)

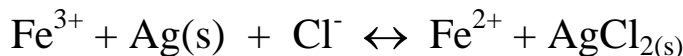
balanced rxn for the titration



b)



Or



c)

$$E_{\text{cell}} = E^0_{(\text{A}/\text{DA})} - E_{(\text{Ag}/\text{AgCl})} - 0.05916/2 \log([\text{A}]/[\text{D}][\text{H}^+]^2)$$

$$E_{\text{cell}} = E^0_{(\text{Fe}^{3+}/\text{Fe}^{2+})} - E_{(\text{Ag}/\text{AgCl})} - 0.05916 \log([\text{Fe}^{2+}]/[\text{Fe}^{3+}])$$

d) Equiv pt is 20.00 mL

at 5.0 mL of A added

mmol  $\text{Fe}^{3+} =$

$$(10.0 \text{ mL})(0.0200 \text{ M}) - (10.0 \text{ mL})(0.010 \text{ M}) = 0.10 \text{ mmol}$$

$$\text{mmol } \text{Fe}^{3+} = (10.0 \text{ mL})(0.010 \text{ M}) = 0.10 \text{ mmol}$$

$$\begin{aligned} E_{\text{cell}} &= E^0_{(\text{Fe}^{2+}/\text{Fe}^{3+})} - E_{(\text{Ag}/\text{AgCl})} - 0.05916 \log([\text{Fe}^{2+}]/[\text{Fe}^{3+}]) \\ &= 0.767 - 0.197 - 0.05916 \log(1) \\ &= 0.570 \text{ V} \end{aligned}$$

At 10.0 mL of A added we are at the equivalence pt

At the equiv. pt.

$$[\text{Fe}^{2+}] = [\text{D}], \text{ and } [\text{Fe}^{3+}] = [\text{A}]$$

To calculate  $E_{\text{cell}}$ , we can make use of this info by adding the following two equations

$$E_+ = E^0_{(\text{A}/\text{DA})} - 0.05916/2 \log([\text{A}]/[\text{D}][\text{H}^+]^2)$$

$$E_+ = E^0_{(\text{Fe}^{3+}/\text{Fe}^{2+})} - 0.05916 \log([\text{Fe}^{2+}]/[\text{Fe}^{3+}])$$

To effectively combine the log terms of the two equations when we add them, it is convenient to multiply the first equation by a factor of 2

$$\begin{aligned} 2E_+ &= 2E^0_{(\text{A}/\text{DA})} - 0.05916 \log([\text{A}]/[\text{D}][\text{H}^+]^2) \\ + \quad E_+ &= E^0_{(\text{Fe}^{3+}/\text{Fe}^{2+})} - 0.05916 \log([\text{Fe}^{2+}]/[\text{Fe}^{3+}]) \\ \hline 3E_+ &= 2E^0_{(\text{A}/\text{DA})} + E^0_{(\text{Fe}^{3+}/\text{Fe}^{2+})} - \\ &\quad 0.05916 \log([\text{A}][\text{Fe}^{2+}]/[\text{Fe}^{3+}][\text{D}][\text{H}^+]^2) \\ 3E_+ &= 2(.390) + 0.767 - 0.05916 \log(1/[\text{H}^+]^2) \\ \text{pH} &= 0.30 \rightarrow [\text{H}^+] = 0.501 \text{ M} \\ 3E_+ &= 1.511 \text{ V} \\ E_+ &= 0.504 \text{ V} \\ E_{\text{cell}} &= E^+ - E^0_{\text{Ag}/\text{AgCl}} = 0.504 - 0.197 = 0.307 \text{ V} \end{aligned}$$

At 15 mL added

mmol A =

$$(5.0 \text{ mL})(0.010 \text{ M}) = 0.050 \text{ mmol}$$

mmol D =

$$(10.0 \text{ mL Fe}^{3+})(0.0200 \text{ M Fe}^{3+})(1 \text{ mol D} / 2 \text{ mol Fe}^{3+}) = 0.100 \text{ mmol}$$

(look at stoichiometry of the overall rxn)

$$\begin{aligned} E_{\text{cell}} &= E_{(A/DA)}^0 - E_{(Ag/AgCl)} - 0.05916/2 \log([A]/[D][H^+]^2) \\ &= 0.390 - 0.197 - 0.05916/2 \log\{(.05)/(0.100)(.501)^2\} \\ &= 0.184 \text{ V} \end{aligned}$$

15-11

The Walden reductor uses the standard Ag/AgCl couple to reduce  $\text{Fe}^{3+}$ . The standard reduction potential for the Ag/AgCl couple is large enough (0.222 V) that  $\text{Cr}^{3+}$  and  $\text{TiO}^{2+}$  are not reduced.

Where as the reduction potential for the Zn/ $\text{Zn}^{2+}$  (Jones reducer) is much less  $-0.764 \text{ V}$ , and  $\text{Cr}^{3+}$  and  $\text{TiO}^{2+}$  are reduced.

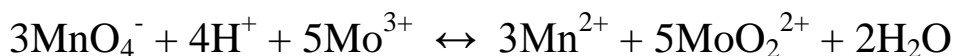
15-14

When 25.00 mL of unknown was passed through a Jones reductor,  $\text{MoO}_4^{2-}$  was converted to  $\text{Mo}^{3+}$ . The filtrate required 16.43 mL of 0.01033 M  $\text{KMnO}_4$  to reach an endpoint.

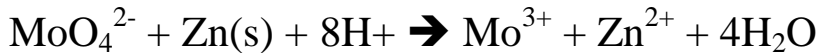


A blank required 0.04 mL. Balance the reaction and find the molarity of moly species in the unknown.

This is not a balanced redox reaction as written. Write two balanced-half reactions and add.



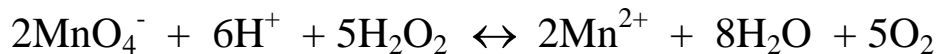
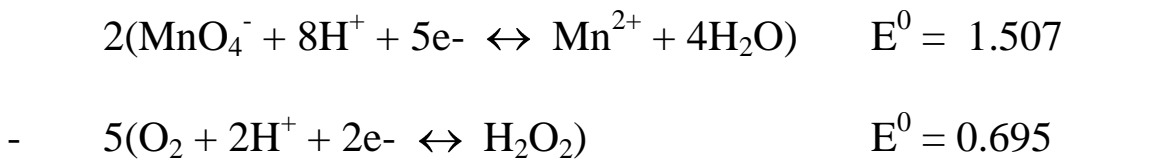
Jones reductor



$$\begin{aligned} [\text{MoO}_4^{2+}] &= (16.39 \text{ mL MnO}_4^-)(0.01033 \text{ M MnO}_4^-) * (5 \text{ mmol} \\ &\quad \text{MoO}_4^{2+} / 3 \text{ mmol MnO}_4^-) / (25.00 \text{ mL}) \\ &= 0.01129 \text{ M MoO}_4^{2+} \end{aligned}$$

15-15

A 25.00 mL aliquot of commercial a hydrogen peroxide solution was diluted to 250.0 mL in a volumetric flask. Then 25.00 mL of the diluted solution was mixed with 200 mL of water and 20 mL 3 M H<sub>2</sub>SO<sub>4</sub> and titrated with 0.02123 M KMnO<sub>4</sub>. The first pink color was observed at 27.66 mL of titrant added. A blank prepared from water in place of the diluted hydrogen peroxide solution required 0.04 mL to give a visible pink color. Find the molarity of the commercial hydrogen peroxide solution.



$$\text{mmol MnO}_4^- = (27.62 \text{ mL})(0.02123 \text{ M}) = 0.5863 \text{ mmol MnO}_4^-$$

$$\begin{aligned} [\text{H}_2\text{O}_2]_{\text{dil}} &= (0.5863) * (5 \text{ mol H}_2\text{O}_2 / 2 \text{ mol MnO}_4^-) / (25.00 \text{ mL}) \\ &= 0.05864 \text{ M} \end{aligned}$$

$$[\text{H}_2\text{O}_2] = (0.05864)(250 \text{ mL} / 25 \text{ mL}) = 0.5864 \text{ M}$$

dilution factor

15-18

An aqueous glycerol solution weighing 100.0 mg was treated with 50.0 mL of 0.0837 M  $\text{Ce}^{4+}$  in 4 M  $\text{HClO}_4$  at 60 C for 15 min to oxidize the glycerol to formic acid. The excess  $\text{Ce}^{4+}$  required 12.11 mL of 0.0448 M  $\text{Fe}^{2+}$  to reach a ferroin endpoint. What is the weight percent of glycerol in the unknown?

Write half rxn for the oxidation of glycerol to formic acid

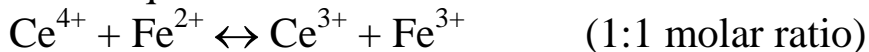


Write balanced rxn of glycerol with  $\text{Ce}^{4+}$



(8:1 molar ratio)

Write equation for the back titration



mmole  $\text{Ce}^{4+}$  needed to titrate glycerol (G) =

$$(50.0 \text{ mL})(0.0837 \text{ M}) - (12.11 \text{ mL})(0.0448 \text{ M}) = 3.6425 \text{ mmol } \text{Ce}^{4+}$$

$$\begin{aligned} \text{mmol } \text{C}_3\text{H}_8\text{O}_3 &= (3.6425 \text{ mmol } \text{Ce}^{4+})(1 \text{ mmol G}/8 \text{ mmol } \text{Ce}^{4+}) \\ &= 0.45531 \text{ mmol G} \end{aligned}$$

$$\text{mg G} = (0.45531 \text{ mmol G})(92.0938 \text{ mg G}/\text{mmol G}) = 41.931 \text{ mg}$$

$$\text{wt \% G} = [(41.931 \text{ mg G}) / (100.0 \text{ mg sample})]100 = 41.9 \%$$

15-19

mmole  $\text{Ce}^{4+}$  needed to titrate  $\text{NO}_2^-$  =

$$(50.0 \text{ mL})(0.1186 \text{ M}) - (31.13 \text{ mL})(0.04289 \text{ M}) = 4.59483 \text{ mmol } \text{Ce}^{4+}$$

(mmol  $\text{NO}_2^-$ )<sub>dil</sub> =

$$(4.59483 \text{ mmol } \text{Ce}^{4+})(1 \text{ mmol } \text{NO}_2^-/2 \text{ mmol } \text{Ce}^{4+}) = 2.29742 \text{ mmol } \text{NO}_2^-$$



$$(\text{mmol NO}_2^-)_{\text{sample}} = (2.29742 \text{ mmol NO}_2^-) * (500.0/25.00) = 45.94 \text{ mmol NO}_2^-$$

$$\begin{aligned} \text{mg G} &= (45.94 \text{ mmol NO}_2^-) * (68.995 \text{ mg NaNO}_2/\text{mmol NO}_2^-) \\ &= 3170.2 \text{ mg NaNO}_2 = 3.1702 \text{ g NaNO}_2 \end{aligned}$$

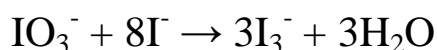
$$\% \text{ NaNO}_2 = (3.1702 \text{ g NaNO}_2)/(4.030 \text{ g sample}) * 100 = 78.67 \%$$

15.25

A potassium iodate solution was prepared by dissolving 1.022 g of  $\text{KIO}_3$  (FM 214.00) in a 500 mL flask. Then 50.00 mL of this solution was pipetted into a flask and treated with 2 g KI and 10 mL of 0.5 M  $\text{H}_2\text{SO}_4$ . How many moles of  $\text{I}_3^-$  are formed?

a.  $\text{mol IO}_3^- = (1.022 \text{ g})(1 \text{ mol}/214.00 \text{ g}) = 0.004776 \text{ mol}$

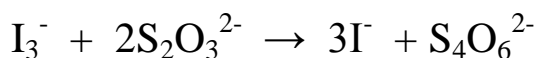
Reaction



50/500 mL taken or 0.4776 mmol

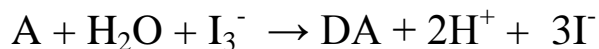
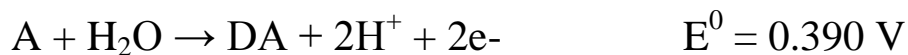
Thus, 1.433 mmol  $\text{I}_3^-$  is formed.

b. Titration Reaction



So...  $[\text{S}_2\text{O}_3^{2-}] = (2)(.001433)/(0.03766 \text{ L}) = 0.07609 \text{ M}$

c.



14.33 mmol  $\text{I}_3^-$  (or added to the ascorbic acid, 50 out of 500 mL)

$(14.22 \text{ mL})(0.07609 \text{ M}) = 1.082 \text{ mmol S}_2\text{O}_3^{2-}$  added. It will react with  $0.5410 \text{ mmol I}_3^-$ .

Therefore,  $1.433 - 0.5410 = 0.892 \text{ mmol I}_3^-$  reacted with  $0.892 \text{ mmol}$  of ascorbic acid.

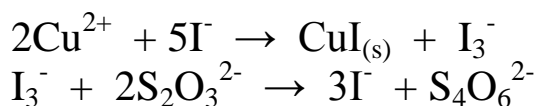
$$(0.892 \text{ mmol})(176.13 \text{ mg/mmol})/1000 = 0.157 \text{ g A}$$

$$\text{Wt \%} = (0.157/1.223) \cdot 100 = 12.8 \%$$

d) must add indicator right before the endpoint.

15-26

A  $3.026 \text{ g}$  portion of a copper(II) salt was dissolved in a  $250 \text{ mL}$  volumetric flask. A  $50.0 \text{ mL}$  aliquot was analyzed by adding  $1 \text{ g}$  of  $\text{KI}$  and titrating the liberated  $\text{I}_3^-$  with  $23.33 \text{ mL}$  of  $0.04668 \text{ M}$  thiosulfate std. Find the weight % of copper in the sample.



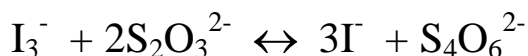
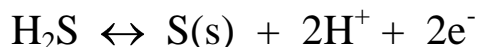
$$\text{Liberated I}_3^- = (23.33 \text{ mL})(0.04668 \text{ M})/2 = 0.5445 \text{ mmol I}_3^-$$

$$\text{mmol Cu}^{2+} \text{ in } 50.00 \text{ mL aliquot} = (0.5445)(2/1) = 1.089 \text{ mmol Cu}^{2+}$$

$$\text{mass Cu}^{2+} \text{ in sample} = (1.089 \text{ mmol})(63.546 \text{ mg/mmol Cu})(250/50)(1 \text{ g}/1000 \text{ mg}) = 0.3460 \text{ g}$$

$$\text{weight \%} = (0.3460/3.026) \cdot 100 = 11.44 \%$$

15-27



$$(\text{mmol I}_3^-)_{\text{tot}} = (25.00 \text{ mL})(0.01044 \text{ M}) = 0.26100 \text{ mmol I}_3^-$$

$$(\text{mmol I}_3^-)_{\text{S}_2\text{O}_3} = (14.44 \text{ mL})(0.009336 \text{ M}) * (1 \text{ mol I}_3^- / 2 \text{ mol S}_2\text{O}_3^{2-}) \\ = 0.067406 \text{ mmol I}_3^-$$

$$[\text{H}_2\text{S}] = (0.26100 - 0.067406) * (1 \text{ mol H}_2\text{S} / 1 \text{ mol I}_3^-) / 25.00 \text{ ml} \\ = 0.007744 \text{ M or } 7.744 \text{ mM}$$

Since we are titrating a solution that contain  $\text{I}_3^-$  before the equivalence point it is important that we wait until just before the equivalence point to add the starch indicator.

15-28

a)

Subtract the second and last equation



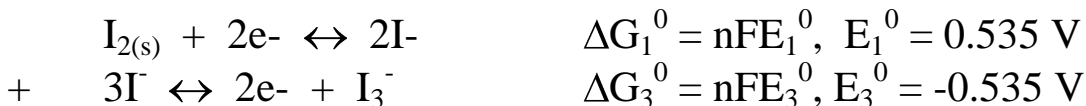
you can say:

$nFE_r^0 = (nFE_2^0 + nFE_3^0)$ , since both half rxns are two electron transfers  $\Rightarrow E_r^0 = E_2^0 + E_3^0 = 0.085 \text{ V}$

$$K = 10^{n(E_1^0 + E_3^0)/0.05916} = 10^{(2*0.085/0.05916)} = 7*10^2$$

b)

Subtract the first and last equation

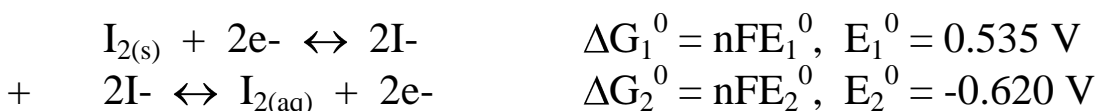


$$\begin{aligned}
 \Delta G_{\text{rxn}}^0 &= (nFE_1^0 + nFE_3^0) = 0 = -RT \ln K_{\text{eq}} \\
 K_{\text{eq}} &= 1.0
 \end{aligned}$$

$$\begin{aligned}
 nFE_r^0 &= (nFE_1^0 + nFE_3^0), \text{ since both half rxns are two electron} \\
 \text{transfers} &\Rightarrow E_r^0 = E_1^0 + E_3^0 = 0.000 \text{ V} \\
 K &= 10^{\frac{nF(E_1^0 + E_3^0)}{RT}} = 1.0
 \end{aligned}$$

c)

Subtract the first two equations



$$\begin{aligned}
 nFE_r^0 &= (nFE_1^0 + nFE_2^0), \text{ since both half rxns are two electron} \\
 \text{transfers} &\Rightarrow E_r^0 = E_1^0 + E_2^0 = -0.085 \text{ V} \\
 K &= 10^{\frac{n(E_1^0 + E_2^0)}{0.05916}} = 10^{(2 \cdot -0.085/0.05916)} = 1.338 \cdot 10^{-3}
 \end{aligned}$$

$$K = [\text{I}_{2(\text{aq})}] = 0.001338 \text{ mol/L} \Rightarrow 0.3 \text{ g/L}$$

15-31

WoW!!! This is a fun one!!!! You must break it down into steps.

There are four different reactions that are occurring in this experiment; the first two are given, the last two are not.

Let us write the last two out

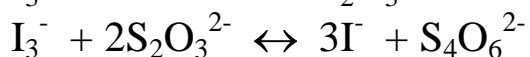
Rxn 3

The excess  $\text{Br}_2$  is converted to  $\text{Br}^-$ , producing  $\text{I}_3^-$



Rxn 4

$\text{I}_3^-$  is titrated with  $\text{S}_2\text{O}_3^{2-}$



OK!

moles of  $\text{I}_3^-$  produced from rxn 3 =

$$(8.83 \text{ ml})(0.05113 \text{ M}) * (1 \text{ mol } \text{I}_3^- / 2 \text{ mol } \text{S}_2\text{O}_3^{2-}) = 0.22574 \text{ mmol } \text{I}_3^-$$

moles of excess  $\text{Br}_2$  left over from rxn 2 =

$$(0.22574 \text{ mmol } \text{I}_3^-) * (1 \text{ mol } \text{Br}_2 / 1 \text{ mol } \text{I}_3^-) = 0.22574 \text{ mmol } \text{Br}_2$$

mmol  $\text{Br}_2$  produced from rxn 1 =

$$(25.00 \text{ ml})(0.02000 \text{ M}) * (3 \text{ mol } \text{Br}_2 / 1 \text{ mol } \text{BrO}_3^-) = 1.500 \text{ mmol } \text{Br}_2$$

mol  $\text{Br}_2$  reacted with  $\text{Al}(\text{C}_9\text{H}_6\text{ON})_3$  =

$$1.500 \text{ mmol} - 0.22574 \text{ mmol} = 1.2743 \text{ mmol } \text{Br}_2$$

mmol  $\text{Al}^{3+}$  in unknown =

$$(1.2743 \text{ mmol } \text{Br}_2) * (1 \text{ mol } \text{C}_9\text{H}_7\text{ON} / 2 \text{ mol } \text{Br}_2) * (1 \text{ mol } \text{Al}^{3+} / 3 \text{ mol } \text{C}_9\text{H}_7\text{ON}) = 0.21238 \text{ mmol } \text{Al}^{3+} \Rightarrow 5.730 \text{ mg Al}$$

In order to eliminate rounding errors it is important that you keep as many figures as we can in each of the intermediate calculations. Then, for the final answer you must go back through each step starting from the first one to determine the correct sig. figs that carry through to the answer.