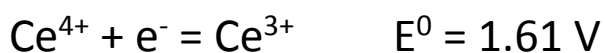


Another example of a redox titration.

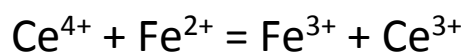
50.00 mL of 0.0100 M Fe^{2+} in 0.1 M H_2SO_4 is titrated with 0.0100 M Ce^{4+}

a) What is the reaction when Ce^{4+} is added to Fe^{2+} ?



$$E^0_{\text{cell}} = E_{\text{cath}} - E_{\text{anod}} = 1.61 \text{ V} - 0.771 \text{ V} = 0.839 \text{ V}$$

So the rxn is spontaneous.



b) Calculate E_{cell} with SCE and Pt at 25.00 mL titrant added.

c) 50.00 mL added

d) 75.00 mL added

b) 25.00 mL of 0.0100 M Ce^{4+}

Initial mol $\text{Fe}^{2+} = 50.00 \text{ mL}(0.0100 \text{ M}) = 0.500 \text{ mmol}$

Added mol $\text{Ce}^{4+} = 25.00 \text{ mL}(0.0100 \text{ M}) = 0.250 \text{ mmol}$

Ce^{4+}	+ Fe^{2+}	= Fe^{3+}	+ Ce^{3+}
0.250	0.500 mmol	0	0
-0.250	-0.250	+0.250	+0.250 mmol
<hr/>			
0	0.250	0.250	0.250 mmol

$\text{Fe}^{2+/3+}$ governs the potential at the Pt wire.

Total vol = 50.00 + 25.00 mL

$[\text{Fe}^{2+}] = 0.250 \text{ mmol} / 75.00 \text{ mL} = 3.33\text{e-}3 \text{ M}$

$[\text{Fe}^{3+}] = 0.250 \text{ mmol} / 75.00 \text{ mL} = 3.33\text{e-}3 \text{ M}$

$E = 0.771 - 0.0592 \log [\text{Fe}^{2+}] / [\text{Fe}^{3+}] = 0.771\text{V}$

$E_{\text{cell}} = 0.771 - 0.241 = 0.530 \text{ V}$

c) 50.00 mL of 0.0100 M Ce^{4+}

Added mol $\text{Ce}^{4+} = 50.00 \text{ mL}(0.0100 \text{ M}) = 0.500 \text{ mmol}$

This is the Eq. Pt.

At Eq. Pt. $[Ce^{4+}] = [Fe^{2+}]$ & $[Fe^{3+}] = [Ce^{3+}]$

2 Nernst Eqn.

$$E = 0.771 - 0.0592 \log [Fe^{2+}] / [Fe^{3+}]$$

$$+E = 1.61 - 0.0592 \log [Ce^{3+}] / [Ce^{4+}]$$

$$2E = 2.381 - 0.0592 \log [Fe^{2+}] / [Fe^{3+}] - 0.0592 \log [Ce^{3+}] / [Ce^{4+}]$$

$$2E = 2.381 - 0.0592 \log [Fe^{2+}][Ce^{3+}] / [Fe^{3+}][Ce^{4+}]$$

$$\text{Sub } 2E = 2.381 - 0.0592 \log [Fe^{2+}] [Fe^{3+}] / [Fe^{3+}][Fe^{2+}]$$

$$E = 1.19$$

$$E_{\text{cell}} = 1.19 - 0.241 = 0.95 \text{ V}$$

c) 75.00 mL of 0.0100 M Ce^{4+}

Added mol $Ce^{4+} = 75.00 \text{ mL}(0.0100 \text{ M}) = 0.750 \text{ mmol}$

Ce^{4+}	+ Fe^{2+}	= Fe^{3+}	+ Ce^{3+}
0.750	0.500 mmol	0	0
-0.500	-0.500	+0.500	+0.500 mmol
0.250	0	0.500	0.500 mmol

$$[Ce^{4+}] = 0.500 \text{ mmol} / 125.0 \text{ mL} = 4.00e-4 \text{ M}$$

$$[Ce^{3+}] = 0.250 \text{ mmol} / 125.0 \text{ mL} = 2.00e-4 \text{ M}$$

$$E = 1.61 - 0.0592 \log [Ce^{3+}] / [Ce^{4+}] = 1.61 - 0.0592 \log 2.00e-4 / 4.00e-4$$

$$E = 1.63 \text{ V}$$

$$E_{\text{cell}} = 1.63 - 0.241 \text{ V} = 1.39 \text{ V}$$