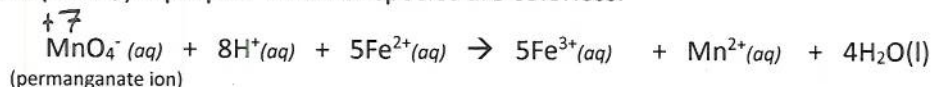


Name Key Reox titration problem set II

When $\text{KMnO}_4(\text{aq})$ is mixed with a solution of $\text{FeSO}_4(\text{aq})$ and sulfuric acid (H_2SO_4), the balanced reaction can be written as shown below. Note that the spectator ions, K^+ and SO_4^{2-} , can be left out of the equation. The permanganate ion (MnO_4^-) is purple. All other species are colorless.



1. What is being oxidized in the above equation? $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$

2. What is being reduced? $\overset{+7}{\text{Mn}} + 5\text{e}^- \rightarrow \text{Mn}^{2+}$

3. What is the oxidizing agent? MnO_4^-

4. What is the reducing agent? Fe^{2+}

5. If it takes 10.00 mL of 0.020 M KMnO_4 to oxidize 10.00 mL of FeSO_4 solution, what is the concentration (M) of the FeSO_4 solution?

$$\frac{0.020 \text{ mol KMnO}_4}{1000 \text{ mL solution}} \times 10.00 \text{ mL} = \frac{0.0002 \text{ mol KMnO}_4}{1 \text{ mol MnO}_4^-} \times \frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol MnO}_4^-} = 0.001 \text{ mol Fe}^{2+}$$

$$\frac{0.001 \text{ mol FeSO}_4}{0.01 \text{ L solution}} = 0.10 \text{ M FeSO}_4$$

6. If it takes 12.45 mL of 0.025 M KMnO_4 to oxidize 20.00 mL of FeSO_4 solution, what is the concentration (M) of the FeSO_4 solution?

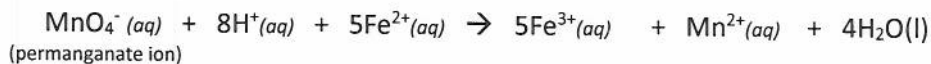
$$\frac{0.025 \text{ mol MnO}_4^-}{1000 \text{ mL solution}} \times 12.45 \text{ mL solution} = \frac{0.00031125 \text{ mol MnO}_4^-}{1 \text{ mol MnO}_4^-} \times \frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol MnO}_4^-} = 0.001556 \text{ mol Fe}^{2+}$$

$$\frac{0.001556 \text{ mol FeSO}_4}{0.02 \text{ L solution}} = 0.0778 \text{ M} = 0.078 \text{ M FeSO}_4$$

7. How many mL of 0.25 M KMnO_4 solution would be required to oxidize 34.00 mL of 0.10 M FeSO_4 ?

$$\frac{0.10 \text{ mol FeSO}_4}{1000 \text{ mL solution}} \times 34.00 \text{ mL} = \frac{0.0034 \text{ mol Fe}^{2+}}{5 \text{ mol Fe}^{2+}} \times \frac{1 \text{ mol MnO}_4^-}{1 \text{ mol MnO}_4^-} = 0.00068 \text{ mol MnO}_4^-$$

$$\frac{0.00068 \text{ mol MnO}_4^-}{0.25 \text{ mol}} \times 1000 \text{ mL solution} = 2.72 \text{ mL KMnO}_4 \text{ solution}$$



8. How many mL of 0.15 M KMnO_4 solution would be required to oxidize 5.00 mL of 0.20 M FeSO_4 ?

$$\frac{0.20 \text{ mol Fe}^{2+}}{1000 \text{ ml solution}} \times 5.00 \text{ ml solution} = 0.001 \text{ mol Fe}^{2+}$$

$$\frac{1 \text{ mol MnO}_4^-}{5 \text{ mol Fe}^{2+}} = 0.0002 \text{ mol MnO}_4^-$$

$$\frac{0.0002 \text{ mol MnO}_4^-}{0.15 \text{ mol MnO}_4^- / 1000 \text{ ml solution}} = 1.33 \text{ mL or } 1.3 \text{ mL KMnO}_4 \text{ solution}$$

9. If 30.00 mL of 0.01 M KMnO_4 are mixed with 30.00 mL of 0.01 M FeSO_4 , what will be the color of the resulting solution? Explain!

$$\frac{0.01 \text{ mol}}{1000 \text{ mL}} \times 30.00 \text{ mL} = 0.0003 \text{ mol MnO}_4^-$$

$$0.0003 \text{ mol Fe}^{2+}$$

MnO_4^- is in excess and Fe^{2+} is limiting, so it will be purple.

10. If 20.00 mL of 0.10 M KMnO_4 are mixed with 20.00 mL of 1.00 M FeSO_4 , what will be the color of the resulting solution? Explain!

$$\frac{0.10 \text{ mol}}{1000 \text{ ml}} \times 20.00 \text{ ml} = 0.002 \text{ mol MnO}_4^-$$

$$\frac{1 \text{ mol}}{1000 \text{ ml}} \times 20.00 \text{ mL} = 0.02 \text{ mol Fe}^{2+}$$

The equation requires 5X the Fe^{2+} , but we have 10X as many moles of Fe^{2+} as MnO_4^- , so Fe^{2+} is in excess - it will be clear.

A buret filled with a 0.025 M solution of KMnO_4 is used to titrate a FeSO_4 solution of unknown concentration. The data are shown below.

Molarity of KMnO_4 solution	0.025 M
Initial buret reading (mL)	12.00
Final buret reading (mL)	30.00
Volume of FeSO_4 solution in flask under buret	10.00 mL

11. What is the concentration (M) of the FeSO_4 solution?

$$\frac{0.025 \text{ mol KMnO}_4}{1000 \text{ ml solution}} \times 18 \text{ ml KMnO}_4 \text{ solution} = 0.00045 \text{ mol KMnO}_4$$

$$\frac{5 \text{ mol Fe}^{2+}}{1 \text{ mol KMnO}_4} = 0.00225 \text{ mol Fe}^{2+}$$

$$\frac{0.00225 \text{ mol Fe}^{2+}}{0.01 \text{ L FeSO}_4 \text{ solution}} = 0.225 \text{ M FeSO}_4$$