You are given 50.00 mL of a potassium hydroxide solution that is of unknown concentration. You perform a titration using 0.075M HNO_3 and find that you reach the equivalence point of the titration after 42.35 mL of the acid was added.

a) What was the molarity of the original potassium hydroxide solution?

$0.04235\,L\ HNO_3 \times \frac{0.075\,mol}{1\,L} = 3.176\times10^{-3}\,mol\ acid$

$\downarrow$

$3.176\times10^{-3}\,mol\ base\ originally$

$\frac{3.176\times10^{-3}\,mol}{0.05\,L} = 0.0635\,M\ KOH$

b) What was the pH of the original potassium hydroxide solution?

$[OH^-] \times [H_3O^+] = 1\times10^{-14}$ $\quad [H_3O^+] = \frac{1\times10^{-14}}{0.0635\,M} = 1.57\times10^{-13}\,M$

$pH = -\log(1.57\times10^{-13}) = 12.80$

c) What is the pH of the solution after the titration is finished?

7 - b/c neutral

In the presence of acid, potassium dichromate ($K_2Cr_2O_7$) reacts with with the Fe^{2+} ion to yield Fe^{3+}, water and the Cr^{3+} ion.

a) Write the balanced chemical equation for this process and identify what is being oxidized and what is being reduced.

$$Cr_2O_7^{2-}{}_{(aq)} + 14H_3O^+_{(aq)} + 6Fe^{2+}_{(aq)} \longrightarrow 2Cr^{3+}_{(aq)} + 21H_2O_{(l)} + 6Fe^{3+}_{(aq)}$$

Fe^{2+} oxidized to Fe^{3+}

Cr^{+6} reduced to Cr^{3+}

Note K^+ is a spectator

b) A titration with potassium dichromate can be used to measure the amount of iron (II) present in a solution. To do this, one uses an indicator that turns purple in the presence of the dichromate ion. You have a solution with a volume of 25.00 ml that contains an unknown amount of iron (II). You titrate it with a 0.250 M solution of potassium dichromate and the solution turns purple after you have added 17.34 mL of the dichromate. How many moles of Fe^{2+} were present in the original solution?

$$0.01734\,L \text{ dichromate soln} \times \frac{0.250 \text{ mol } Cr_2O_7^{2-}}{1\,L} \times \frac{6 \text{ mol } Fe^{2+}}{1 \text{ mol } Cr_2O_7^{-2}} = 0.02601 \text{ mol } Fe^{2+}$$