Determination of Iron(II) by Titration with Permanganate

Fe^{2+} is oxidized to Fe^{3+} by MnO_4^{1-} in acidic solution. The MnO_4^{1-} is converted to Mn^{2+} in the reaction. Before the beginning of lab, balance the equation for the reaction.

You will be provided with ~0.02 M standard MnO_4^{1-} solution, an unknown Fe^{2+} solution, a buret, pipette, and other necessary glassware. Using good quantitative techniques, transfer, via pipette, 10.00 mL of the iron solution to a titration flask. Add 50 mL of 1M H_2SO_4 to the flask and begin the titration by adding a few mL of the MnO_4^{1-} titrant. When a distinct yellow color develops, add 3 mL of 85% H_3PO_4 and continue the titration. The end point is when the first faint purple color of the MnO_4^{1-} persists in the flask. Repeat the titration at least two more times. Your report will consist of the data sheet on the next page.

Answer the following question prior to lab: If a 10.00 mL sample of Fe^{2+} solution is titrated with 47.25 mL of 0.02453 M MnO_4^{1-}, what is the molar concentration of Fe^{2+} in the iron solution?

Conc. of Standard [MnO_4^{1-}]=				
	mL titrant	moles MnO_4^{1-}	moles Fe^{2+}	[Fe^{2+}]
Titration 1				
2				
3				
4				
5				
Average [Fe^{2+}]				
Standard Deviation for [Fe^{2+}]				