Name: Key

Pd:_____ Date:_____

Chemistry **12 Lesson #7 Redox Titrations**

Balance the following redox reaction (in acid solution):

$MnO_4^{1-} + Fe^{2+} \rightarrow Mn^{2+} + Fe^{3+}$

$5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O$

$(Fe^{2+} \rightarrow Fe^{3+} + 1e^-)\ 5$

$8H^+ + MnO_4^- + 5Fe^{2+} \longrightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

When MnO_4^{1-} is reduced to Mn^{2+} there is a colour change. MnO_4^{1-} is purple and Mn^{2+} is colourless. So in a titration the MnO_4^- undergoes reduction and becomes colourless. When you see the colour purple in the solution you have reached the equivalence point, and you STOP titrating.

Note: Spectator Ions are chemical species that are present in a reaction but they do not participate. As an analogy, before you get you learn to drive your parents must bring you to your sporting events. They are there but they do not play in the game.

Examples: for this unit: $K^+ / Cs^+ / Li^+ / Rb^+ / Na^+$

$NO_3^- / SO_4^{2-} / F^- / Cl^-$

When 20.0 mL of unknown concentration of $Fe(NO_3)_2$ is titrated with 15.25 mL of 0.100 M $KMnO_4$ (under acid conditions colour change is observed. What is the concentration of the $Fe(NO_3)_2$?

Step 1. Balanced redox equation

note: $Fe(NO_3)_2 \rightarrow Fe^{2+} + 2NO_3^-$ $\quad KMnO_4 \rightarrow MnO_4^- + K^+$

$8H^+ + MnO_4^- + 5Fe^{2+} \longrightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$

Step 2. Determine the moles of known

$0.01525\ L \times \frac{0.100\ mol\ MnO_4^-}{1\ L} = 0.00125\ mol\ MnO_4^-$

Step 3. Cross the mole bridge to determine moles of unknown

$0.00125\ mol\ MnO_4^- \times \frac{5\ mol\ Fe^{2+}}{1\ mol\ MnO_4^-} \times \frac{1\ mol\ Fe(NO_3)_2}{1\ mol\ Fe^{2+}} = 0.006125\ mol\ Fe(NO_3)_2$

Step 4. Calculate the unknown concentration

$M = \frac{0.006125\ mol\ Fe(NO_3)_2}{0.0200\ L} \therefore$ **0.381 M $Fe(NO_3)_2$**

Another common substances used in redox titrations is the reduction of $Cr_2O_7^{2-}$ to Cr^{3+} under acid conditions.

$Cr_2O_7^{2-}$ (orange) → Cr^{3+} (colourless)

A titration is performed with an unknown volume of a 0.50 M sample of $SnCl_2$ in an acidic solution. The titration requires 15.0 mL if 0.030 M $K_2Cr_2O_7$. What is the volume of the $SnCl_2$ used if the UNBALANCED redox reaction is:

note: $SnCl_2 \rightarrow Sn^{2+} + 2e^-$ $\quad K_2Cr_2O_7 \rightarrow 2K^+ + Cr_2O_7^{2-}$

$Cr_2O_7^{2-} + Sn^{2+} \rightarrow Sn^{4+} + Cr^{3+}$

$6e^- + 14H^+ + Cr_2O_7^{2-} \rightarrow Cr^{3+} + 7H_2O$

Step 1. Balanced redox equation: $(Sn^{2+} \rightarrow Sn^{4+} + 2e^-)3$

$$14H^+ + Cr_2O_7^{2-} + 3Sn^{2+} \rightarrow 3Sn^{4+} + 2Cr^{3+} + 7H_2O$$

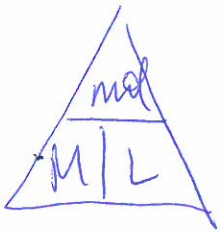
Step 2. Determine the moles of known

$$0.0150\ L \times \frac{0.030\ mol\ Cr_2O_7^{2-}}{1\ L} = 0.00045\ mol\ Cr_2O_7^{2-}$$

Step 3. Cross the mole bridge to determine moles of unknown

$$4.5 \times 10^{-4}\ mol\ Cr_2O_7^{2-} \times \frac{3\ mol\ Sn^{2+}}{1\ mol\ Cr_2O_7^{2-}} \times \frac{1\ mol\ SnCl_2}{1\ mol\ Sn^{2+}} =$$

Step 4. Calculate the unknown volume



$$L = \frac{0.00135\ mol\ SnCl_2}{0.50\ M\ SnCl_2}$$

$$= 0.0027\ L \text{ or } 2.7\ mL$$