Unit 4 – Chapter 4: Types of Reactions **Name** KEY

Pre-Lab Questions: Permanganate Determination of the Iron Sample **Period** ______

Redox prelab problems:

1) A solution of permanganate was standardized by titration with oxalic acid ($H_2C_2O_4$). It required 28.97 mL of the permanganate solution to react completely with 0.1058 grams of oxalic acid. The unbalanced equation for the reaction is:

$$MnO_4^-{}_{(aq)} + H_2C_2O_{4(aq)} \rightarrow Mn^{2+}{}_{(aq)} + CO_{2(aq)}$$

Calculate the molarity of the permanganate solution.

$2 H \times 1 = 2$
$2 C \times 12 = 24$
$4 O \times 16 = 64$
$90g\ H_2C_2O_4$

$$6H^+ + 2MnO_4^- + 5H_2C_2O_4 \rightarrow 2Mn^{2+} + 8H_2O + 10CO_2$$

28.97mL ?M   0.1058g

$$0.1058g\ H_2C_2O_4 \times \frac{1\ mol\ H_2C_2O_4}{90g\ H_2C_2O_4} \times \frac{2\ mol\ MnO_4^-}{5\ mol\ H_2C_2O_4} = 0.00047\ mol\ MnO_4^-$$

$$M = \frac{mol\ solute}{L\ sol'n}$$

$$M = \frac{0.00047\ mol\ MnO_4^-}{0.02897\ L} = \boxed{0.01622M\ MnO_4^-}$$

2) A 50.00 mL sample of solution containing Fe^{2+} ions is titrated with a 0.0216 *M* $KMnO_4$ solution. It required 20.62 mL of the $KMnO_4$ solution to oxidize all the Fe^{2+} ions to Fe^{3+} ions by the reaction:

$$MnO_4^-{}_{(aq)} + Fe^{2+}{}_{(aq)} \rightarrow Mn^{2+}{}_{(aq)} + Fe^{3+}{}_{(aq)} \quad \text{(unbalanced)}$$

a. What was the concentration of the Fe^{2+} ions in the sample solution?

b. What volume of 0.0150 *M* $K_2Cr_2O_7$ solution would it take to do the same titration? The reaction is:

$$Cr_2O_7^{2-}{}_{(aq)} + Fe^{2+}{}_{(aq)} \rightarrow Cr^{3+}{}_{(aq)} + Fe^{3+}{}_{(aq)}$$

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

0.0216M 20.62mL   50.00mL

$$0.0216M = \frac{X\ mol\ MnO_4^-}{0.02062L} \rightarrow X = 0.000445\ mol\ KMnO_4 \times \frac{5\ mol\ Fe^{2+}}{1\ mol\ KMnO_4} = 0.00223\ mol\ Fe^{2+}$$

$$\frac{0.00223\ mol\ Fe^{2+}}{0.0500\ L} = \boxed{0.0446M\ Fe^{2+}}$$

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

50.0mL   0.0150M

From part a) → 0.00223 mol

$$0.00223\ mol\ Fe^{2+} \times \frac{1\ mol\ Cr_2O_7^{2-}}{6\ mol\ Fe^{2+}} = 0.000372\ mol\ Cr_2O_7^{2-}$$

$$0.0150M\ Cr_2O_7^{-2} = \frac{0.000372\ mol}{X\ L\ sol'n} \qquad \boxed{X = 24.8mL}$$

3) The iron content of iron ore can be determined by titration with standard $KMnO_4$ solution. The iron ore is dissolved in HCl and all the iron is reduced to Fe^{2+} ions. This solution is then titrated with $KMnO_4$ solution, producing Fe^{3+} and Mn^{2+} ions in acidic solution. If it required 41.95 mL of 0.0205 *M* $KMnO_4$ to titrate a solution made from 0.6128 grams of iron ore, what is the mass percent of iron in the iron ore?

$$8H^+ + \underset{\substack{41.95\,mL \\ 0.0205\,M}}{MnO_4^-} + 5\,Fe^{2+} \rightarrow Mn^{2+} + 4\,H_2O + 5\,Fe^{3+}$$

$$\frac{0.0205\ mol\ MnO_4^-}{1\ L} \times 0.04195\ L = 0.000860\ mol\ KMnO_4 \times \frac{5\ mol\ Fe^{2+}}{1\ mol\ KMnO_4} = 0.00430\ mol\ Fe^{2+}$$

$$0.00430\ mol\ Fe^{2+} \times \frac{56\ g\ Fe^{2+}}{1\ mol\ Fe^{2+}} = 0.24\ g\ Fe^{2+}$$

$$\frac{0.24\ g\ Fe}{0.6128\ g\ ORE} \times 100\% = \boxed{39.2\%\ Fe\ IN\ ORE}$$