Last name: ______________________ First name: ______________________

Student Number: ______________________

Lab Day and Section**(1 point)** : ______________________

Lab TA name**(1 point)** : ______________________

Chemistry 1311 D -2
Midterm October 2012

1. Nitric acid is used in the production of fertilizers and explosives. One step in the industrial production of nitric acid is the reaction of ammonia with molecular oxygen to form nitrogen monoxide.

In a study of this reaction, a chemist mixed 220 g ammonia (NH_3) with 100 L of oxygen at 30°C and 201.5 kPa and allowed them to react to completion. The unbalanced chemical reaction is:

$$NH_3 + O_2 \longrightarrow NO + H_2O \quad (1)$$

a) **(1 point)** balance reaction (1) above

$(4e + 4H^+ + O_2 \rightarrow 2H_2O)\,5$

$(H_2O + NH_3 \rightarrow NO + 5H^+ + 5e)\,4$

$5O_2 + 4NH_3 \rightarrow 4NO + 6H_2O$ or $2NH_3 + \frac{5}{2}O_2 + 2NO + 3H_2O$

b) **(4 points)** What masses of NO and H_2O are produced?

NH_3 220g 17g/mol $n = \frac{220}{17} = 12.94$ moles

O_2 $n = \frac{PV}{RT} = \frac{201.5\times10^3 \times 100 L \times 10^{-3} m^3}{303\,(8.314)} = 7.999$ mols.

If I were to use all the NH3

$n_{O_2} = n_{NH_3} \times \frac{5O_2}{4NH_3} - 16.18$ moles O_2 needed, only have 8

O_2 limiting.

$n_{NO} = n_{O_2} \times \frac{4NO}{5O_2} = 7.999 \times \frac{4}{5} = 6.4$ moles NO

$mass_{NO} = n_{NO} \times MM_{NO} = 6.4 \times 30 g/mol = 192 g$ NO

$mass_{H_2O} = n_{H_2O} \times MM_{H_2O} = n_{O_2} \times \frac{6H_2O}{5O_2} \times 18 g/mol = 8 \times \frac{6}{5} \times 18 = 173 g\ H_2O$

2. a) **(4 points)** Assign oxidation states and balance the following reaction in acid:

1. $MnO_4^-\ (aq) + H_2O_2\ (aq) \rightarrow Mn^{2+}\ (aq) + O_2(g)$

$\overset{+7}{Mn}\overset{-2}{O_4^-}$, $\overset{+1}{H_2}\overset{-1}{O_2}$, $\overset{+2}{Mn^{2+}}$, $\overset{0}{O_2}$

2. $(5e + MnO_4^- + 8H^+ \longrightarrow Mn^{2+} + 4H_2O)2$

3. $(H_2O_2 \longrightarrow O_2 + 2H^+ + 2e)5$

$\cancel{10}e + 2MnO_4^- + \cancel{16}H^+ + 5H_2O_2 \longrightarrow 2Mn^{2+} + 8H_2O + 5O_2 + \cancel{10H^+} + \cancel{10}e$ 4.

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$$2MnO_4^- + 5H_2O_2 + 6H^+ \longrightarrow 2Mn^{2+} + 8H_2O + 5O_2$$

b) **(4 points)** Consider the above reaction, if the original concentration of H_2O_2 was 0.52 M, what volume of 0.124 M $KMnO_4$ is required to titrate (to equivalence) 10.0 ml of $H_2O_2(aq)$?

10 ml H_2O_2 0.52 M

V? 0.124 M

$KMnO_4 \longrightarrow K^+(aq) + MnO_4^-(aq)$

n_{MnO_4} ?

$n_{H_2O_2} = 0.52\,M \times 10\times10^{-3}\,l = 5.2\times10^{-3}$ moles

$n_{MnO_4^-} = n_{H_2O_2} \times \dfrac{2MnO_4^-}{5H_2O_2}$

$= 5.2\times10^{-3} \times 2/5 = 2.08\times10^{-3}$ moles

$[\cancel{MnO_4^-}] = 2.08\times10^{-3} = 0.124\,M \times V$

$V = 0.0168\,l$; 16.8 ml

3. Combustion analysis of a 1.70 g of a compound formed as a by-product of the industrial synthesis of polymers produced 1.32 g of CO_2 and 0.631 g of H_2O. The mass percent of iodine in the compound was determined by converting the iodine in a 0.850 g sample of the compound into 1.15 g of lead(II)iodide, PbI_2.

a) **(2 points)** What is the % by mass of C in the sample?

$n\,CO_2 = \frac{1.32}{44} = 0.03$ moles ; mass C = 0.03 × 12 = 0.36g

% C = 21.18%

b) **(2 points)** What is the % by mass of H in the sample?

$n_{H_2O} = \frac{0.631}{18 g/mol} = 0.03505$ moles ; mass H = 0.03505 × 1.008 × 2
= 0.0353g × 2
= 0.0701

% H = ~~2.08%~~ 4.12%

c) **(2 points)** What is the % by mass of I in the sample?

PbI_2 = 207.2 + 2(126.9) = 461.

$n\,PbI_2$ = 1.15g/461 g/mol = 2.495×10^{-3} $n_I = 2\times2.495\times10^{-3} = 4.989\times10^{-3}$

mass I = $4.989\times10^{-3}\times126.9$ = 0.633g

% I = 0.633/0.850 = 74.48

d) **(1 points)** Could this compound also contain oxygen? Explain your answer.

100 − 21.18 − ~~4.08~~ 4.12 − 74.48 =
100 − 99.78 = 0.22%
No → %C + %H + %I ≈ 100.

e) **(2 points)** What is the empirical formula of the compound?

100g C: 21.18g = 1.765 moles ÷ .5869 = 3.007
H 4.12g = 4.087 ÷ .5869 = 6.96
I 74.48g = 0.5869 ÷ .5869 = 1

C_3H_7I

f) **(1 points)** If the molar mass of the compound was found experimentally to be 527 g what is it's molecular formula.

$\frac{MM}{EM} = \frac{527 g/mol}{170} = 3.1 \sim 3$

$C_9H_{21}I_3$

5. **(7 points)** The following initial rate information was collected at 25 °C for the reaction of I^- with $S_2O_8^{2-}$:

$2\,I^-(aq) + S_2O_8^{2-}(aq) \rightarrow I_2(aq) + 2\,SO_4^{2-}(aq)$

$[I^-]_o$ (M)	$[S_2O_8^{2-}]_o$ (M)	Initial Rate M/min	
0.125	0.150	4.4×10^{-2}	1
0.375	0.150	2.3×10^{-1}	2
0.125	0.050	1.5×10^{-2}	3

Determine the rate law, order and evaluate the rate constant for this reaction.

I^-

$$\frac{R_1}{R_2} = \left(\frac{.125}{.375}\right)^x = \left(\frac{4.4\times10^{-2}}{2.3\times10^{-1}}\right)$$

$$(.333)^x = (0.191)$$

$$x \log .333 = \log .191$$

$$x(-0.4815) = -0.7189$$

$$x = 1.5$$

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$S_2O_8^{2-}$

$$\frac{R_1}{R_3} = \frac{4.4}{1.5} = \left(\frac{.150}{.050}\right)^x$$

$$2.93 = 3^x$$

$$x = 1$$

2

$$R = k[S_2O_8^{2-}]^1[I^-]^{1.5}$$

$M^{2.5}$

1.5

$$k = \frac{R}{[S_2O_8][I^-]^{1.5}} = \frac{4.4\times10^{-2}}{.15(.125)^{1.5}} = 6.637$$

$$= \frac{2.3\times10^{-1}}{.15(.375)^{1.5}} = 6.677$$

$$k = 6.65\ M^{-1.5}S^{-1}$$

order 2.5

4. You have three beakers each containing a 200 ml of water. To the first you add 850 mg of solid LiCl, to the second you add 6.8 g of solid $Pb(NO_3)_2$ and to the third, you add 2.4 g of solid $MgSO_4$.

a) **(1 point)** What species will you find in the first beaker ?

$Li^+(aq)$ $Cl^-(aq)$ H_2O

b) **(1 point)** in the second ?

$Pb^{2+}(aq)$ $NO_3^-(aq)$ H_2O

c) **(1 point)** in the third ?

$Mg^{2+}(aq)$ $SO_4^{2-}(aq)$

d) **(3 point)** Next you mix all three beakers together, identify any precipitates that could form.

$$Pb^{2+}(aq) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s)$$

$$Pb^{2+}(aq) + Cl^-(aq) \longrightarrow PbCl_2(s)$$

e) **BONUS:** Next you throw in a handful of $(NH_4^+)_3PO_4$. What happens ?

Puts $NH_4^+(aq)$ & $PO_4^{3-}(aq)$ into sol^n

$$\left.\begin{aligned} n \text{ : } LiCl &= \frac{0.850}{(7+35.5)} = 0.02 \text{ moles} \\ n\ Pb^{2+} &= \frac{6.8g}{(207.2+2(14+3(16))} = 0.02 \text{ moles.} \\ n\ SO_4^{2-} &= \frac{2.4g}{(32+4(16))} = 0.025 \text{ moles} \end{aligned}\right\}$$

Cl^- gets rid of $\frac{1}{2}$ Pb^{2+}

SO_4^{2-} gets of rest.

PO_4^{3-} will ppt Mg^{2+}

$Mg_3(PO_4)_2(s)$

6. **(4 points)** Calculate the heat in kilojoules to vaporize 40 g CCl_4 initially at 25 ^oC.
You might be interested that:
the boiling point of CCl_4 is 77 ^oC,
the melting point is -23 ^oC
the molar heat capacity of CCl_4 (s) is 44.22 J $mol^{-1}\,^oC^{-1}$;
the molar heat capacity of CCl_4(l) is 132.4 J $mol^{-1}\,^oC^{-1}$ ←
the molar heat capacity of CCl_4(g) is 82.7 J $mol^{-1}\,^oC^{-1}$.

MM = 153.81 g/mol
n_{CCl_4} = 40/153.8
= 0.260 moles

$C(graphite) + 2\,Cl_2\,(g) \rightarrow CCl_4(g)$ $\Delta H = -102.9$ kJ mol^{-1}

$C(graphite) + 2\,Cl_2\,(g) \rightarrow CCl_4\,(l)$ $\Delta H = -135.4$ kJ mol^{-1}

CCl4 is a liquid.

$$
\begin{array}{lll}
C(grap) + 2Cl_2(g) \rightarrow CCl_4(g) & & -102.9 \\
CCl_4(l) \rightarrow C(grap) + 2Cl_2(g) & & +135.4 \\
\hline
CCl_4(l) \rightarrow CCl_4(g) & & 32.5
\end{array}
$$

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$$heat = \underset{warm}{nC_m\Delta T} + n\Delta H$$

$= 0.260(132.4\,J\,mol^{-1}\,C^{-1})(77-25) + 0.260(325\times10^3/mol)$

$= 1.79\,kJ + 8.45\,kJ$

$= \underline{\underline{10.24\,kJ}}$

Question	Points Possible	Points Earned
Title page	2	
1	5	
2	8	
3	10	
4	6	
5	7	
6	4	
Total	40	