Last name :________________Given name :______________________________

Student number:__________________________

Lab Section:__________________________Lab TA: ____________________.

Chemistry 1311D
Test 1 V1
October 9, 2014

1. **a) (4 points)** CFCs which contain only carbon, chlorine and fluorine wreak havoc with the ozone layer. A 2.55 g sample of a particular CFC is contained in a 1.50 L flask at 25°C and has a pressure of 262 torr. What is the molar mass of this compound?

$m = 2.55g$

$V_T = 1.50L$ (3.00L) (0.75L)

$T = 25°C = 298.2K$

$P_T = 262$ torr (131 torr) (524 torr) $= 262 \text{ torr} \times \frac{1 \text{ bar}}{750.06 \text{ torr}} = 0.3493$ bar 1

$PV = nRT$ 1

$\frac{n}{V} = \frac{P}{RT}$

$\frac{m}{V} = \frac{\mathcal{M}P}{RT}$

$\frac{2.55g}{1.5L} = \frac{\mathcal{M}(0.3493 \text{ bar})}{0.08314 \frac{L\,bar}{mol\,K}(298.2K)}$ 1

$1.7 = \mathcal{M}(0.01409)$ 1

$\mathcal{M} = 120.7$ g/mol 120.7 g/mol 120.7 g/mol

b) **(1 points)** Assuming the compound contains only one carbon atom, what is the molecular formula?

CCl_4 154 g/mol

CCl_3F 137.5 g/mol

$\boxed{CCl_2F_2 \quad 121 \text{ g/mol}}$ ← all versions

$CClF_3$ 104.5 g/mol

CF_3 88 g/mol

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2.a) **(4 points)** An aqueous nitric acid, HNO_3, solution can convert CuS into soluble copper (II) hydrogen sulphate. Balance the following reaction:

$$CuS(s) + NO_3^-(aq) \rightarrow NO(g) + Cu^{2+}(aq) + HSO_4^-(aq)$$

+2 −2   +5 −2   +2 −2   +2   +1 +6 −2   oxidation states

$$(CuS(s) + 4H_2O(l) \rightarrow Cu^{2+}(aq) + HSO_4^-(aq) + 7H^+ + 8e)\,3$$

$$(4H^+ + NO_3^-(aq) + 3e \rightarrow NO(g) + 2H_2O)\,8$$

$$3CuS(s) + \cancel{12H_2O(l)} + \overset{11}{\cancel{32H^+}}(aq) + 8NO_3^-(aq) + \cancel{24e} \rightarrow 3HSO_4^-(aq) + 3Cu^{2+}(aq) + \cancel{21H^+(aq)} + \cancel{24e} + 8NO(g) + \underset{4}{\cancel{16}}H_2O(l)$$

$$3CuS(s) + 8NO_3^-(aq) + 11H^+(aq) \rightarrow 3Cu^{2+}(aq) + 3HSO_4^-(aq) + 8NO(g) + 4H_2O(l)$$

b) **(4 points)** Calculate the mass of NO(g) if 500 ml of 2 M nitric acid (an aqueous solution of 2M nitric acid contains 2 M of $H^+(aq)$ and 2M of $NO_3^-(aq)$) react with 46.2 g of CuS

Limiting reagent!

$n_{H^+} = 0.500L(2M) = 1mole$ (0.250L, 0.5mole) (0.750 L, 1.5mole) 1

$n_{NO_3^-} = 1$ mole

$n_{CuS} = m/M = 46.2g/95.611g/mole = 0.483$ moles (23.1g, 0.242) (69.3g, 0.725)

IF all CuS used

$n_{H^+} = 0.483$ moles CuS $\times \frac{11}{3} = 1.77$ moles
$n_{NO_3^-}$ 0.483 moles CuS $\times \frac{8}{3} = 1.29$ moles

} needed   CuS is in excess, H^+ runs out first   1.5

$n_{NO} - n_{H^+} \times \frac{8NO}{11H^+} = 1 \times 8/11 = 0.7273$ mols (0.3636 moles) (1.0909 moles) 1

mass NO = n_{NO}(16+14.01 g/mol) = 21.8g (10.9g) (32.7g) 0.5

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3. A 6.47 g sample of compound containing only C, H and O combusts in air. $CO_2(g)$ is collected at 22.5 °C and 1.088 bar over water to a volume of 8.06 L. This combustion experiment also produced 6.30 ml of $H_2O(l)$. Using the appropriate data for vapour pressure and density of water from Table 3.1 below:

Temperature	Vapour Pressure (mm Hg)	Density (g/ml)
20	17.55	0.9982
21	18.66	0.9980
22	19.84 *	0.9978 *
23	21.09 *	0.9975 *
24	22.40	0.9970

a) **(2 points)** determine the amount (in moles) of hydrogen, H, in the 6.47 g sample

$V_{H_2O} = 6.30\,ml$

$d_{H_2O} = 0.9977\,g/ml$

$M = 18.02\,g/ml$

$m = d \times V = 6.30\,ml \times 0.9977\,g/ml = 6.2855\,g$

$n_{H_2O} = 6.2855\,g / 18.02\,g/mol = 0.3488\,mols$

$n_H = n_{H_2O} \times \frac{2H}{1H_2O} = 0.3488 \times 2 = 0.6976\,moles$ (0.703g)

b) **(3 points)** determine the amount (in moles) of carbon, C, in the 6.47 g sample

$P_T = 1.088\,bar$, $V_T = 8.06\,L$, $T = 22.5\,°C = 295.7\,K$

$P_{H_2O} = 20.465\,mmHg \times \frac{1\,bar}{750.06\,mmHg} \approx 0.027\,bar$

$P_{CO_2} = P_T - P_{H_2O} = 1.088 - 0.027 = 1.061\,bar$

$n_C = n_{CO_2} = \frac{PV}{RT} = \frac{1.061\,bar\,(8.06\,L)}{0.08314\,\frac{l\,bar}{mol\,K}\,(295.7\,K)} = 0.3478\,moles$ (4.178g)

c) **(1 point)** determine the amount (in moles) of oxygen, O, in the 6.47 g sample

$mass\ O = m_T - m_C - m_H = 6.47 - 0.703 - 4.178\,g = 1.589\,g$

$n_O = 1.589\,g / 15.999\,g/mol = 0.0993\,moles.$

d) **(2 points)** determine the empirical formula

$C_{\frac{0.3478}{0.0993}} H_{\frac{0.6976}{0.0993}} O_{\frac{0.0993}{0.0993}} = C_{3.5}H_{7.03}O \rightarrow C_7H_{14}O_2$

e) **(2 points)** if the experimentally determine molar mass was found to be 265 g/mol, what is the molecular formula of the compound.

$M_{C_7H_{14}O} = 7(12) + 14 + 2(16) = 130\,g/mol$

$Factor = \frac{265\,g/mol}{130\,g/mol} = 2.03 = 2$ $C_{14}H_{28}O_4$

(128, $C_7H_{14}O_2$)

(387, C_{21}, $H_{42}O_6$)

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4. Ammonia(g) reacts with O_2(g) to form nitric acid via the following overall process:

$$12\ NH_3(g) + 21\ O_2(g) \rightarrow 8HNO_3(g) + 4NO(g) + 14\ H_2O(g)$$

a) **(2 points)** Given the heats of formation:

$\Delta_f H°(NH_3(g)) = -45.9 kJ\ mol^{-1}$

$\Delta_f H°(HNO_3(g)) = -133.9 kJ\ mol^{-1}$

$\Delta_f H°(H_2O(g)) = -241.8 kJ\ mol^{-1}$

$\Delta_f H°(NO(g)) = 91.3 kJ\ mol^{-1}$

Determine the enthalpy of the reaction per mole of NH_3(g)

$\Delta_{RXN} H^0 = 14 \Delta_f H^0(H_2O) + 4 \Delta_f H^0(NO) + 8 \Delta_f H^0(HNO_3) - 12 \Delta_f H^0(NH_3) - 21 \Delta_f H^0(O_2)$

$= 14(-241.8) + 4(91.3) + 8(-133.9) - 12(-45.9) - 21(\phi)$

$= -3540.4$ kJ For 12 moles NH_3

$\Delta_{RXN} H^0$ (per mole) $= -3540.4 kJ/12 = -295.0$ kJ/mol NH_3.

b) **(6 points)** Heat given off is measured via calorimetry. If the temperature of a calorimeter containing 1000g of water with a colorimeter constant *C* of 1235 J °C^{-1} increases from 25°C to 46°C, determine the mass of NH_3(g) used. The specific heat capacity of water is 4.18 J mol^{-1} °C^{-1}.

1000g H_2O

C_{H_2O} 4.18 J mol^{-1} C^{-1}

C_{cal} 1235 J C^{-1}

$T_f = 46°C$ (67°C) (88°C)

$T_i = 25°C$

$q_{sys} = q_{rxn} + q_{H_2O} + q_{cal} = \phi$ 1

$-q_{rxn} = q_{H_2O} + q_{cal}$ 1.5 1.5

$= m C_{H_2O} \Delta T + C_{cal} \Delta T$

$= 1000g(4.18 \frac{J}{g°C})(46-25) + 1235(46-25)$

$= 113\ 715 J$

$= 113.7 kJ$ (227.4kJ) (341.1kJ)

$n_{NH_3} = \frac{q_{rxn}}{\Delta_{RXN} H^0 (\text{per mol})}$ 1

$= \frac{-113.7 kJ}{-295.0}$

$= 0.385$ moles (0.771 moles) (1.156 moles)

mass $NH_3 = n\mathcal{M} = 0.385$ moles (17.03g/mol)

$= 6.6g$ (13.1g) (19.7g)

1

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5. A solution is prepared by dissolving 27.575 g sodium carbonate, Na_2CO_3, in water to make 250 ml of a solution of density 1.103 g/ml. At these concentrations sodium carbonate dissolves completely.

a) **(1 point)** Balance the reaction for dissociation of sodium carbonate:

$Na_2CO_3\ (s) \rightarrow Na^+(aq) + CO_3^{2-}(aq)$

$$Na_2CO_3(s) \longrightarrow 2Na^+(aq) + CO_3^{2-}(aq)$$

0.193

b) **(1 point)** Determine the molarity of $Na^+(aq)$ in the solution.

0.097

$m = 27.575g$

$M_{Na_2CO_3} = 2(M_{Na}) + M_C + 3M_O = 105.99 g/mol$

$n_{Na_2CO_3} = 27.575g / 105.99 = 0.260$ moles

0.048

$n_{Na^+} = 2 \times n_{Na_2CO_3} = 0.520$ moles

$[Na^+] = \dfrac{0.520 \text{ moles}}{0.250L} = 2.08\ M$

~~1.97~~ M

1.93M

c) **(1 point)** Determine the molality of $Na^+(aq)$ in the solution.

$\text{molality} = \dfrac{\text{moles } Na^+}{\text{kg } H_2O}$

0.5 if they use kg soln

55.15

$\text{molality} = \dfrac{0.520}{.2482 \text{ kg}} = 2.10$ molal

1.93

1.93 molal

mass solution = $V \times d = 1.103 g/ml \times 250ml = 275.75g$

50.035

mass H_2O = mass soln − mass Na_2CO_3 = $275.75 - 27.575g = 248.175g$

100.075

d) **(2 point)** Determine the mole fraction of $Na^+(aq)$ in the solution.

$$X = \frac{n_{Na^+}}{n_{Na^+} + n_{CO_3^{2-}} + n_{H_2O}}$$

$$X_{Na^+} = \frac{0.520}{0.520 + 0.260 + 13.772}$$

$n_{Na^+} = 0.520$ moles

$= 3.57 \times 10^{-2}$

$n_{CO_3^{2-}} = n_{Na^+} \times \dfrac{1\ CO_3^{2-}}{2\ Na^+} = 0.260$ moles

3.32×10^{-2}

3.30×10^{-2}

$n_{H_2O} = \dfrac{\text{mass } H_2O}{M_{H_2O}} = \dfrac{248.175g}{18.02 g/ml} = 13.772$ moles

1 if no $n_{CO_3^{2-}}$

Bonus: What is the molarity of CO_2^{2-} ?

$$[CO_3^{2-}] = \frac{1}{2}[Na^+] = 1.04\ M$$

0.97 M

0.97 M.

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6. An element has three naturally occurring isotopes with the following molar masses: 27.97693 amu (natural abundance 92.23 %), 28.97649 amu (natural abundance 3.10 %), and 29.97376 amu.

a) **(2 points)** Determine the average atomic mass for this element.

% abundance 3rd isotope = 100 − 92.23 − 3.10 = 4.67 %

$\mathcal{M}$ = Σ natural abundace × molar mass

= 0.9223(27.97693 amu) + 0.0310(28.97649) + (0.0467)(29.97376) amu

= 28.10117 amu

Average atomic mass: 28.10117

b) **(1 point)** What is this element ? Si

c) **(1 point)** How many neutrons does the most abundant isotope of have? 14

(15)

(16)

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