

Chemistry 1154 Fall Spring 2011 Test 3

Wednesday, March 23, 2010

Time: 1 hour 50 minutes

Name: ANSWERS

Student Number: _____

This exam consists of seven pages of questions, a periodic table, and the formula sheet. Please ensure that you have a complete paper and, if you do not, obtain one from me **immediately**. There are 39.5 marks available. Good luck!

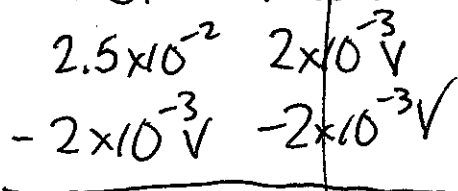
1) [4 marks total] A 10.0-mL aliquot of 2.50×10^{-3} M HCl is to be titrated with 2.00×10^{-3} M NaOH. An indicator with $pK_{in} = 3.602$ is to be used.

a) [3 marks] At what volume of added base will the endpoint of the titration be reached? Give your answer in mL.

2.5×10^{-2} mmol HCl. If V is mL NaOH, then $2 \times 10^{-3} V$ = mmol base. $pK_{ind} = 3.602$ means acid in excess, and

$$[HCl] = [H^+] = 10^{-3.602} = 2.5 \times 10^{-4} \text{ M}$$

HCl + NaOH \rightarrow products



$$\text{mmol HCl in excess} = 2.5 \times 10^{-2} - 2 \times 10^{-3} V$$

$$[HCl] = \frac{2.5 \times 10^{-2} - 2 \times 10^{-3} V}{V + 10} = 2.5 \times 10^{-4}$$

$V = 10.0 \text{ mL}$

b) [1 mark] Is the indicator a suitable one for the titration? How do you know?

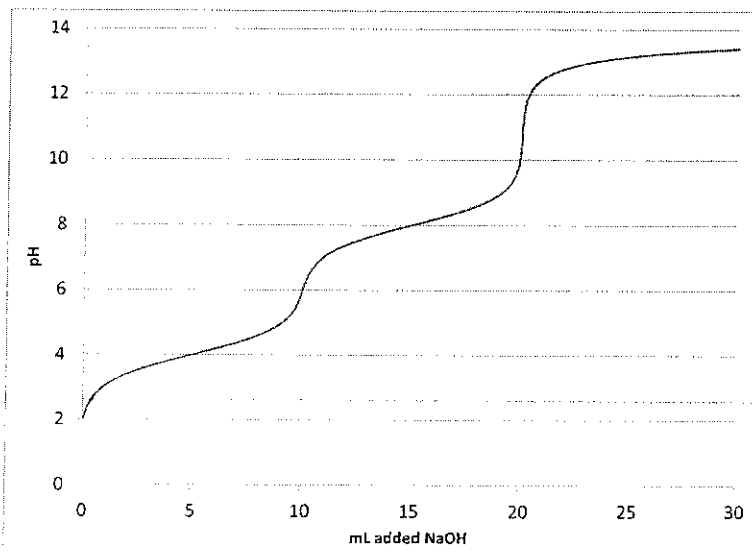
Want end point when moles acid = moles base

$$2.5 \times 10^{-2} = 2 \times 10^{-3} V$$

$$\Rightarrow V = 12.5 \text{ mL}$$

Indicator changes colour 2.5 mL too soon - error of 20%. No good.

- 2) [5 marks total] The titration curve for a weak acid is shown below. Assume that the titration curve is complete as shown, and that all equivalence points are easily visible.

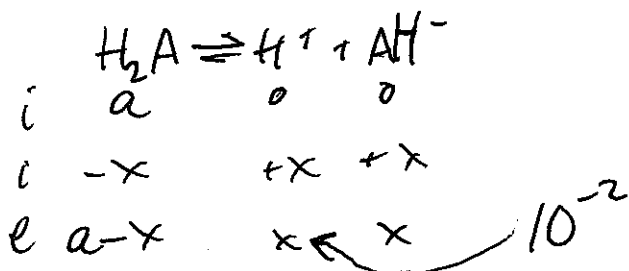


- a) [1 mark] How many acidic protons does the acid have? **2**
- b) [1 mark] How many buffer regions are shown on the graph? **2**
- c) [1 mark] What is or are the approximate pK_a value or values for the acid?

4 + 8

- d) [2 marks] Estimate the concentration of the acid.

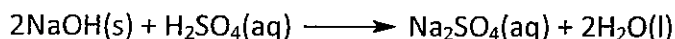
pH is ≈ 2 at start of titration, so:



So "a" is $\approx 1\text{M}$

$$\frac{x^2}{a-x} = 10^{-4}$$

- 3) [4 marks] A 1.00-gram sample of NaOH (40.0 g/mol) was added to 200. mL of 0.100 M H_2SO_4 ($S = 4.184 \text{ J/g}\cdot^\circ\text{C}$, $D = 1.00 \text{ g/mL}$) at 21.000°C .



After the reaction, the temperature of the solution had increased to 23.378°C . Calculate ΔH for the reaction.

LR check: $1.00 \text{ g NaOH} \times \frac{1 \text{ mol}}{40.0 \text{ g}} \times \frac{1 \text{ rxn}}{2 \text{ NaOH}} = 0.0125 \text{ mol rxn}$ NaOH is LR.

$0.2 \text{ L H}_2\text{SO}_4 \times \frac{0.1 \text{ mol}}{1 \text{ L}} \times \frac{1 \text{ rxn}}{1 \text{ H}_2\text{SO}_4} = 0.02 \text{ mol rxn}$

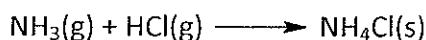
$$q_{\text{rxn}} = -1999.859952 \text{ J}$$

$$+ q_{\text{soln}} = 201 \text{ g} \times 4.184 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \times (23.378 - 21.000)^\circ\text{C} = 1999.859952 \text{ J}$$

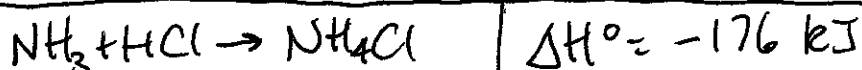
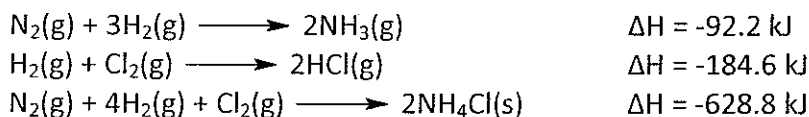
$$\therefore \Delta H = \frac{-1999.859952 \text{ J}}{0.0125} = -159988.7962 \text{ J}$$

or $\sim -160 \text{ kJ}$

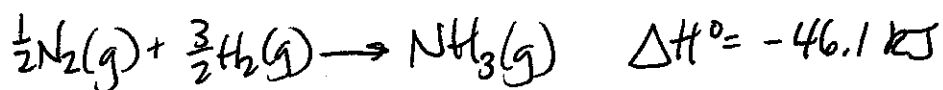
- 4) [2 marks] Calculate ΔH for the reaction:



Given the following data:



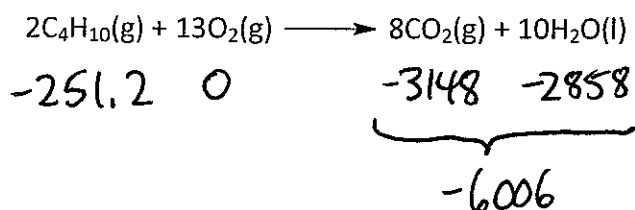
5) [1 mark] What is the enthalpy of formation of $\text{NH}_3(\text{g})$? Give your answer in kJ/mol.



6) [2 marks] Given the following data:

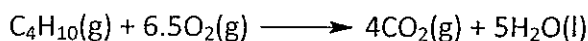
Compound	$\Delta H^\circ_{\text{f},298}$ (kJ/mol)
$\text{C}_4\text{H}_{10}(\text{g})$	-125.6
$\text{CO}_2(\text{g})$	-393.5
$\text{H}_2\text{O}(\text{l})$	-285.8

Calculate ΔH° for the reaction:



$$\Delta H^\circ = -6006 - (-251.2) = \boxed{-5754.8 \text{ kJ}}$$

7) [3 marks] Given that ΔH° for the reaction



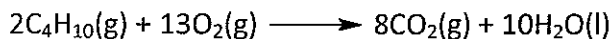
is -2877.4 kJ, how many kJ of heat would be released by the reaction of 5.80 grams of C_4H_{10} (58.0 g/mol) with 10.4 grams of O_2 (32.0 g/mol)?

$$5.80 \text{ g} \times \frac{1 \text{ mol}}{58.0 \text{ g}} \times \frac{-2877.4 \text{ kJ}}{1 \text{ mol}} = -287.74 \text{ kJ}$$

$$10.4 \text{ g} \times \frac{1 \text{ mol}}{32.0 \text{ g}} \times \frac{-2877.4 \text{ kJ}}{6.5 \text{ mol}} = \boxed{-143.87 \text{ kJ}}$$

\therefore 143.87 kJ of heat would be released.

- 8) [3 marks] A 5.8122-gram sample of C_4H_{10} (58.122 g/mol) was burned in excess oxygen in a bomb calorimeter with a heat capacity of 28.687 kJ/°C. As a result of the combustion the temperature of the bomb calorimeter increased by 10.000°C. Calculate ΔH_{298} for the reaction:



$$q_{\text{cal}} = 28.687 \frac{\text{kJ}}{^\circ\text{C}} \times 10.000^\circ\text{C} = 286.87 \text{ kJ}$$

$$+ q_{\text{rxn}} = -286.87 \text{ kJ}$$

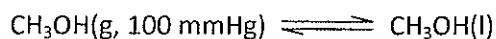
0

$$n_{\text{rxn}} = 5.8122 \text{ g} \times \frac{1 \text{ mol}}{58.122 \text{ g}} \times \frac{1 \text{ rxn}}{2 C_4H_{10}} = 0.05 \text{ moles rxn}$$

$$\Delta E = \frac{-286.87 \text{ kJ}}{0.05} = -5737.4 \text{ kJ}$$

$$\Delta H = -5737.4 + (-7)(8.314472)(298) \times \frac{1 \text{ kJ}}{1000 \text{ J}} = \boxed{-5754.7 \text{ kJ}}$$

- 9) [3 marks] At 21.2°C, the vapour pressure of methanol (CH_3OH) is 100 mmHg. Complete the following chart for the reaction:



Your choices are positive (greater than zero), negative (less than zero), or zero.

	ΔG	ΔH	ΔS
15	-	-	-
20	-	-	-
25	+	-	-

10) [4.5 marks] The following hypothetical process has been proposed to inflate a balloon from 2 litres to 4 litres:

- Perform the inflation against a constant external pressure of 1.50 atm.
- Add 1500 J of heat to the balloon.
- Perform the inflation at 300 K

a) [2 marks] Calculate q , w , ΔE , and ΔH for the process.

$$q = +1500 \text{ J}$$

$$w = -1.50 \text{ atm}(4\text{L}-2\text{L}) \times \frac{8.314472 \text{ J}}{0.0820575 \text{ L}\cdot\text{atm}} = -304 \text{ J}$$

$$\Delta E = q + w = 1196 \text{ J}$$

$$\Delta H = q = +1500 \text{ J}$$

b) [1.5 marks] $\Delta S_{\text{sys}} = +5.76 \text{ J/K}$ for the expansion. Calculate ΔS_{surr} , ΔS_{univ} , and ΔG for the process.

$$\Delta S_{\text{surr}} = \frac{-1500 \text{ J}}{300 \text{ K}} = -5 \frac{\text{J}}{\text{K}}$$

$$\Delta S_{\text{univ}} = 5.76 - 5 = 0.76 \frac{\text{J}}{\text{K}}$$

$$\Delta G = -T\Delta S_{\text{univ}} = -300 \times (0.76) = -228 \text{ J}$$

c) [1 mark] Will the process work as proposed? How do you know?

The process will work as proposed:

$$\Delta G < 0 \text{ or } \Delta S_{\text{univ}} > 0$$

11) [8 marks total] Given the following data:

Compound	$\Delta H^\circ_{f,298}$ (kJ/mol)	S°_{298} (J/mol·K)
CO(g)	-110.5	197.7
O ₂ (g)		205.1
CO ₂ (g)	-393.5	213.7

and the following reaction:



a) [3 marks] Calculate ΔH° , ΔS° , and $\Delta G^\circ_{350\text{K}}$ for the reaction.

$$\Delta H^\circ = 2(-393.5) - 1(0) - 2(-110.5) = -566 \text{ kJ}$$

$$\Delta S^\circ = 2(213.7) - 1(205.1) - 2(197.7) = -173.1 \frac{\text{J}}{\text{K}} \text{ or } -0.1731 \frac{\text{kJ}}{\text{K}}$$

$$\Delta G^\circ = -566 \text{ kJ} - 350(-0.1731) = -505.415 \text{ kJ}$$

b) [2 marks] $\Delta G^\circ_{700\text{K}} = -444.83 \text{ kJ}$. Calculate $\Delta G_{700\text{K}}$ for $P_{\text{CO}} = 1.0 \times 10^{-10} \text{ atm}$, $P_{\text{O}_2} = 1.0 \times 10^{-5} \text{ atm}$, and $P_{\text{CO}_2} = 1.0 \times 10^{+5} \text{ atm}$.

$$Q = \frac{(1 \times 10^5)^2}{(1 \times 10^{-10})^2 (1 \times 10^{-5})} = 1 \times 10^{35}$$

$$\Delta G = -444.83 + \left(8.314472 \times 700 \times \ln(1 \times 10^{35}) \right) \times \frac{1 \text{ kJ}}{1000 \text{ J}}$$

$$= 21.2 \text{ kJ}$$

c) [1 mark] Is the reaction spontaneous at 700 K when carried out under the conditions given in (c)? How do you know?

No. $\Delta G < 0$.

d) [2 marks] Calculate $K_{p,700}$ for the reaction.

$$-444830 = -8.314472 \times 700 \times \ln K$$

$$\Rightarrow K = 1.56 \times 10^{33}$$