

## Chemistry 1105 R10 Spring 2011 Test 3

Friday, March 25, 2011

Time: 1 hour 50 minutes

Name: ANSWERS

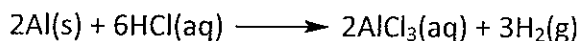
Student Number: \_\_\_\_\_

This test consists of **eight** pages of questions, a page of useful constants, and a periodic table. Please ensure that you have a complete paper and, if you do not, obtain one from me **immediately**. There are **33** marks available. Good luck!

- 1) [4 marks] The following data were collected for the molar mass of Aluminum, carried out in the same way as the molar mass of magnesium experiment:

Mass of Aluminum (g)	
Atmospheric pressure (mmHg)	756.3
Volume of H <sub>2</sub> evolved (mL)	76.0
Temperature of solution (°C)	22.0
Vapour pressure of solution (mmHg)	19.8
Height of column of water (mm)	136.0
Density of water (kg/m <sup>3</sup> )	1000.

Aluminum reacts with HCl according to the equation:



How many grams of aluminum were used in the experiment?

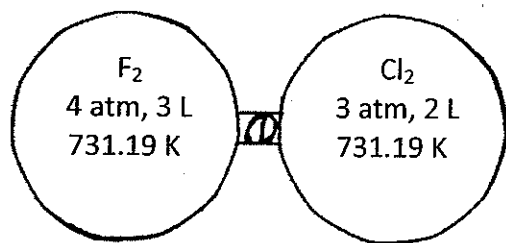
$$P_{\text{H}_2} = 756.3 - 19.8 - \frac{1000 \times 9.80665 \times 0.136 \times 76.0}{101325} = 726.5 \text{ mmHg}$$

$$n_{\text{H}_2} = \frac{(726.5)(76.0)}{(62.3637)(295.15)} = 3 \text{ mmol H}_2$$

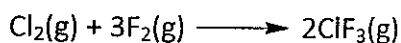
$$3 \text{ mmol H}_2 \times \frac{2 \text{ Al}}{3 \text{ H}_2} = 2 \text{ mmol Al}$$

$$2 \text{ mmol Al} \times \frac{26.98 \text{ g}}{1 \text{ mol}} \times \frac{1 \text{ g}}{1000 \text{ mg}} = \boxed{0.0540 \text{ g}}$$

2) [4 marks] The following apparatus was assembled:



When the valve was opened, the following reaction occurred:



Calculate the partial pressures of all species present after reaction.

$$n_{F_2} = \frac{12}{RT} ; n_{Cl_2} = \frac{6}{RT}$$

LR check:

$$\frac{12}{RT} \text{ moles } F_2 \times \frac{2ClF_3}{3F_2} = \frac{8}{RT} \text{ moles } ClF_3$$

$$\frac{6}{RT} \text{ moles } Cl_2 \times \frac{2ClF_3}{1Cl_2} = \frac{12}{RT} \text{ moles } ClF_3$$

Moles  $Cl_2$  excess:

$$\frac{6}{RT} \text{ moles } Cl_2 - \frac{12}{RT} \text{ moles } F_2 \times \frac{1Cl_2}{3F_2} = \frac{2}{RT}$$

$$\therefore P_{F_2} = 0 ; P_{Cl_2} = \frac{\frac{2}{RT} \times RT}{5} = 0.4 \text{ atm} ; P_{ClF_3} = \frac{\frac{8}{RT} \times RT}{5} = 1.6 \text{ atm}$$

- 3) [3 marks] A 60-gram piece of copper ( $S = 0.385 \text{ J/g}\cdot^\circ\text{C}$ ) at  $10.00^\circ\text{C}$  was placed in 20.0 grams of water ( $S = 4.184 \text{ J/g}\cdot^\circ\text{C}$ ) at  $75.00^\circ\text{C}$ . The water was contained in a cup with heat capacity  $20.0 \text{ J/}^\circ\text{C}$ . What was the final temperature of the water?

$$q_{\text{Cu}} = 60 \text{ g} \times 0.385 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \times (x - 10)^\circ\text{C} = 23.1(x - 10)$$

$$+ q_{\text{H}_2\text{O}} = 20 \text{ g} \times 4.184 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \times (x - 75)^\circ\text{C} = 83.68(x - 75)$$

$$+ q_{\text{Cup}} = 20 \frac{\text{J}}{^\circ\text{C}} \times (x - 75)^\circ\text{C} = 20(x - 75)$$

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0

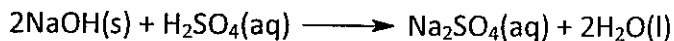
$$\text{So: } 23.1(x - 10) + 83.68(x - 75) + 20(x - 75) = 0$$

$$23.1x - 231 + 83.68x - 6276 + 20x - 1500 = 0$$

$$126.78x - 8007 = 0$$

$$\Rightarrow x = 63.2^\circ\text{C}$$

- 4) [4 marks] A 1.00-gram sample of NaOH (40.0 g/mol) was added to 200. mL of 0.100 M  $\text{H}_2\text{SO}_4$  ( $S = 4.184 \text{ J/g}\cdot^\circ\text{C}$ ,  $D = 1.00 \text{ g/mL}$ ) at  $21.000^\circ\text{C}$ .



After the reaction, the temperature of the solution had increased to  $23.378^\circ\text{C}$ . Calculate  $\Delta H$  for the reaction.

LR check:

$$1 \text{ g NaOH} \times \frac{1 \text{ mol}}{40.0 \text{ g}} \times \frac{1 \text{ rxn}}{2 \text{ NaOH}} = \boxed{0.0125 \text{ moles rxn}}$$

$$200 \text{ L H}_2\text{SO}_4 \times \frac{0.1 \text{ mol}}{\text{L}} \times \frac{1 \text{ rxn}}{1 \text{ H}_2\text{SO}_4} = 0.02 \text{ moles rxn}$$

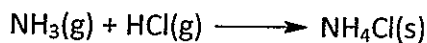
$$q_{\text{rxn}} = -2 \text{ kJ}$$

$$+ q_{\text{soln}} = 201 \text{ g} \times 4.184 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \times (23.378 - 21.000)^\circ\text{C} = \sim 2 \text{ kJ}$$

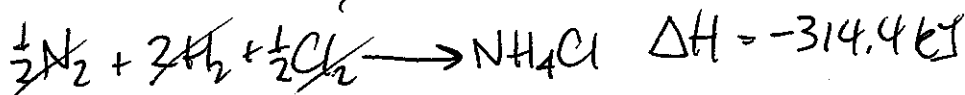
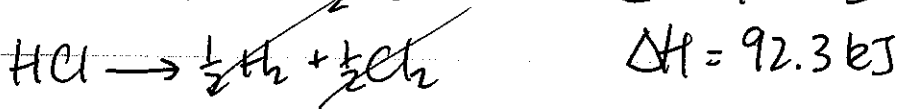
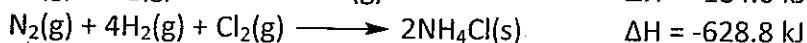
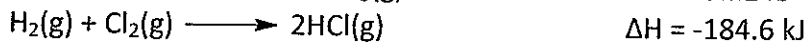
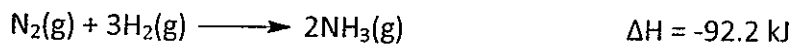
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$$\therefore \Delta H = \frac{-2 \text{ kJ}}{0.0125 \text{ mol}} = \boxed{-160 \text{ kJ}}$$

5) [2 marks] Calculate  $\Delta H$  for the reaction:



Given the following data:



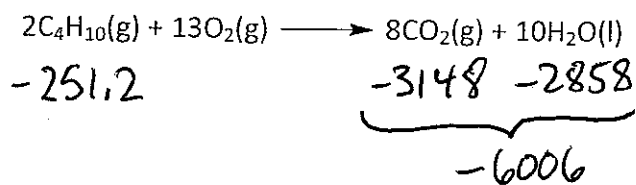
6) [1 mark] What is the enthalpy of formation of  $\text{NH}_3(\text{g})$ ? Give your answer in kJ/mol.

$$\boxed{-46.1 \frac{\text{kJ}}{\text{mol}}}$$

7) [2 marks] Given the following data:

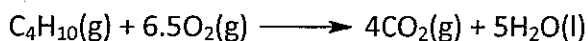
Compound	$\Delta H^\circ_{f,298}$ (kJ/mol)
$\text{C}_4\text{H}_{10}(\text{g})$	-125.6
$\text{CO}_2(\text{g})$	-393.5
$\text{H}_2\text{O}(\text{l})$	-285.8

Calculate  $\Delta H^\circ$  for the reaction:



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

8) [3 marks] Given that  $\Delta H^\circ$  for the reaction



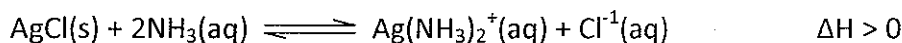
is  $-2877.4 \text{ kJ}$ , how many kJ of heat would be released by the reaction of 5.80 grams of  $\text{C}_4\text{H}_{10}$  (58.0 g/mol) with 10.4 grams of  $\text{O}_2$  (32.0 g/mol)?

$$5.80 \text{ g C}_4\text{H}_{10} \times \frac{1 \text{ mol}}{58.0 \text{ g}} \times \frac{-2877.4 \text{ kJ}}{1 \text{ mol}} = -287.74 \text{ kJ}$$

$$10.4 \text{ g O}_2 \times \frac{1 \text{ mol}}{32.0 \text{ g}} \times \frac{-2877.4 \text{ kJ}}{6.5 \text{ mol}} = -143.87 \text{ kJ}$$

So 143.87 kJ of heat would be released.

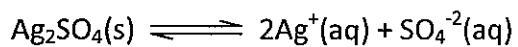
9) [4 marks] Given the following equilibrium:



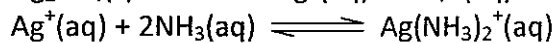
Predict the effect of the following changes on the value of  $K_c$  and of the concentration of  $\text{Cl}^-$ . Your choices are Increase from the current value, Decrease from the current value, or Not Change from the current value. Each change is carried out on a fresh system at equilibrium, and is carried out at constant temperature unless explicitly stated otherwise.

	Effect on:					
	$K_c$			$\text{Cl}^-$		
Heating the reaction mixture	I	D	NC	I	D	NC
Adding $\text{NH}_3$	I	D	NC	I	D	NC
Adding $\text{AgCl}$	I	D	NC	I	D	NC
Adding water	I	D	NC	I	D	NC

10) [2 marks] Given the following equilibria:

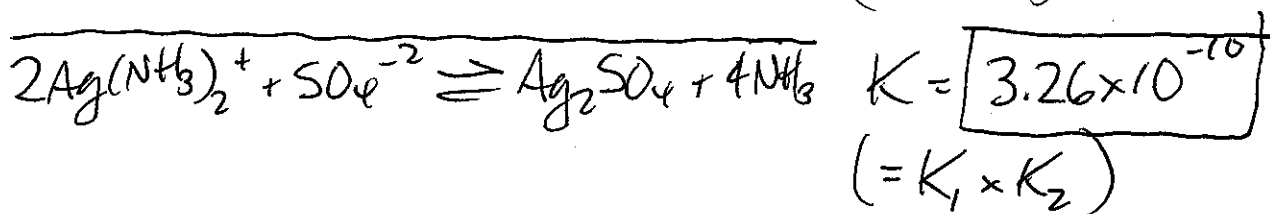
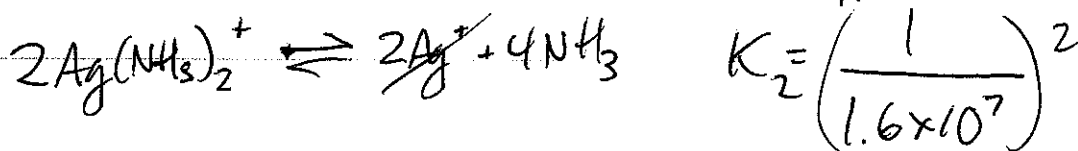
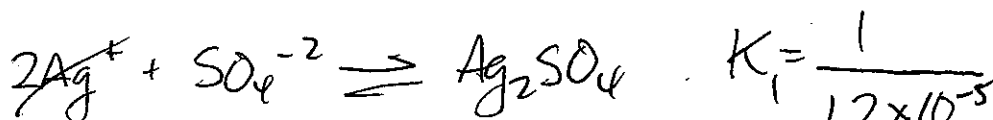
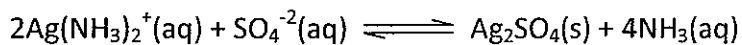


$$K_{\text{sp}} = 1.20 \times 10^{-5}$$

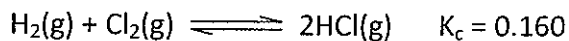


$$K_{\text{f}} = 1.60 \times 10^7$$

Calculate the equilibrium constant for:



- 11) [4 marks total] A 2.00-litre flask was charged with 2.00 moles of  $\text{H}_2$ , 2.00 moles of  $\text{Cl}_2$ , and 4.00 moles of  $\text{HCl}$ , and the equilibrium

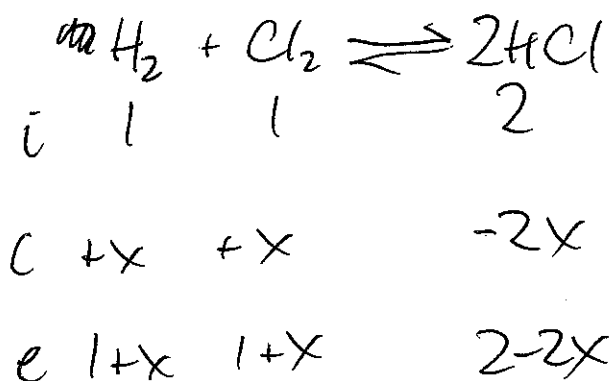


established.

- a) [1 mark] In which direction will the reaction shift in order to establish equilibrium? How do you know?

$$Q = \frac{2^2}{1 \cdot 1} = 4 \quad Q \text{ too big, so rxn } \leftarrow \text{ to reduce products.}$$

- b) [3 marks] What were the equilibrium concentrations of all species once equilibrium had been attained?



$$\frac{(2-2x)^2}{(1+x)^2} = 0.16$$

$\sqrt{\quad}$  both sides

$$\frac{2-2x}{1+x} = 0.4$$

$$\begin{aligned} 2-2x &= 0.4 + 0.4x \\ 2.4x &= 1.6 \\ x &= \frac{2}{3} \end{aligned}$$

$$[\text{H}_2]_e = [\text{Cl}_2]_e = 1+x = \frac{5}{3} \text{ M} \quad (1.\bar{6})$$

$$[\text{HCl}]_e = 2-2x = \frac{2}{3} \text{ M} \quad (0.\bar{6})$$