

## Chemistry 1105 R11 Fall 2011 Test 3

Friday, November 25, 2011

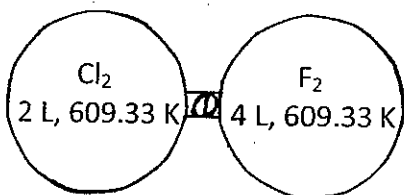
Time: 1 hour 50 minutes

Name: ANSWERS

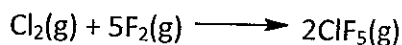
Student Number: \_\_\_\_\_

This test consists of **seven** pages of questions, a periodic table, and a page of useful constants. Please ensure you have a complete paper and, if you do not, obtain one from me **immediately**. There are **31** marks available. Good luck!

1) [3 marks] The following apparatus was assembled:



When the valve was opened, the following reaction occurred:



After reaction, the partial pressures of  $\text{ClF}_5$ ,  $\text{Cl}_2$ , and  $\text{F}_2$ , were found to be 2 atm, 0.5 atm, and 0 atm, respectively. What were the partial pressures of all species present *before* reaction?

$$n_{\text{ClF}_5} = \frac{PV}{RT} = \frac{(2)(6)}{(0.0820575)(609.33)} = 0.24$$

$$\text{F}_2 \text{ is LR, so } n_{\text{F}_2} = 0.24 \text{ moles ClF}_5 \times \frac{5\text{F}_2}{2\text{ClF}_5} = 0.6$$

$$n_{\text{Cl}_2 \text{ reacted}} = 0.24 \text{ moles ClF}_5 \times \frac{1\text{Cl}_2}{2\text{ClF}_5} = 0.12$$

$$n_{\text{Cl}_2 \text{ excess}} = \frac{PV}{RT} = \frac{(0.5)(6)}{(0.0820575)(609.33)} = 0.06$$

$$\left. \begin{array}{l} n_{\text{Cl}_2 \text{ reacted}} \\ n_{\text{Cl}_2 \text{ excess}} \end{array} \right\} n_{\text{Cl}_2 \text{ TOT}} = 0.18 \text{ moles}$$

$$\text{So } P_{\text{F}_2} = \frac{nRT}{V} = \frac{(0.6)(0.0820575)(609.33)}{4} = 7.5 \text{ atm}$$

$$P_{\text{Cl}_2} = \frac{(0.18)(0.0820575)(609.33)}{2} = 4.5 \text{ atm}$$

- 2) [3 marks] A 500-gram piece of copper ( $S = 0.385 \frac{J}{g \cdot ^\circ C}$ ) at  $90.0^\circ C$  was immersed in 75 grams of water ( $S = 4.184 \frac{J}{g \cdot ^\circ C}$ ) at  $25.0^\circ C$ . The water was contained in a cup with  $C = 50.0 \frac{J}{^\circ C}$ . What was the final temperature of the water?

$$q_{Cu} + q_{H_2O} + q_{cup} = 0J$$

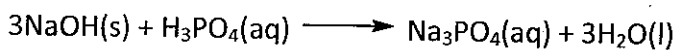
$$500g \times 0.385 \frac{J}{g \cdot ^\circ C} \times (T-90)^\circ C + 75g \times 4.184 \frac{J}{g \cdot ^\circ C} \times (T-25)^\circ C + 50 \frac{J}{^\circ C} \times (T-25)^\circ C = 0J$$

$$192.5(T-90) + 363.8(T-25) = 0$$

$$192.5T - 17325 + 363.8T - 9095 = 0$$

$$556.3T = 26420 \Rightarrow T = 47.49^\circ C$$

- 3) [3 marks] If 3.00 grams of NaOH (40.0 g/mol) are mixed with 100.0 mL of 0.500 M  $H_3PO_4$  ( $4.184 \frac{J}{g \cdot ^\circ C}$ , 1.00 g/mL), the temperature of the solution increases from  $21.605^\circ C$  before reaction to  $30.310^\circ C$  after. Use this information to calculate  $\Delta H^\circ$  for the reaction:



$$3g NaOH \times \frac{1 \text{ mole}}{40.0g} \times \frac{1 \text{ rxn}}{3NaOH} = 0.025 \text{ moles rxn} \leftarrow \text{NaOH is LR}$$

$$\frac{0.5 \text{ moles } H_3PO_4}{L} \times 0.1 L \times \frac{1 \text{ rxn}}{1 H_3PO_4} = 0.05 \text{ moles rxn}$$

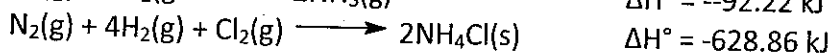
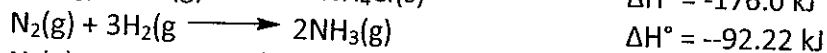
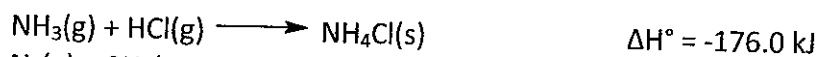
$$q_{rxn} = -3751 J$$

$$+ q_{soln} = (103g) \left( 4.184 \frac{J}{g \cdot ^\circ C} \right) (30.310 - 21.605)^\circ C = 3751 J$$

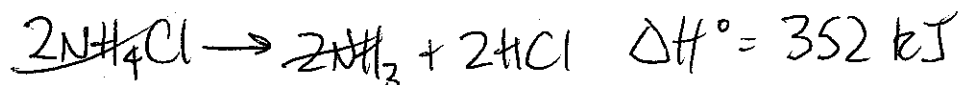
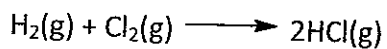
0

$$\Delta H^\circ = \frac{-3751 J}{0.025} = -150 \text{ kJ}$$

4) [3 marks total] Given the following:



a) [2 marks] Calculate  $\Delta H^\circ$  for the reaction

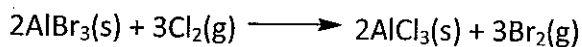


$$\Delta H^\circ = -184.64 \text{ kJ}$$

b) [1 mark] What is the molar enthalpy of formation of  $\text{NH}_4\text{Cl}(\text{s})$ ?

$$-314.43 \text{ kJ}$$

- 5) [5 marks total] Given that the enthalpy of formation for  $\text{AlCl}_3(\text{s})$  is  $-674.8 \text{ kJ/mol}$ , and that  $\Delta H^\circ$  for the reaction



is  $-327.04 \text{ kJ}$ :

- a) [2 marks] what is the molar enthalpy of formation of  $\text{AlBr}_3(\text{s})$ ?

$$-327.04 = 2(-674.8) + 3(0) - 3(0) - 2x$$

$$\Rightarrow x = \boxed{-511.28 \text{ kJ}}$$

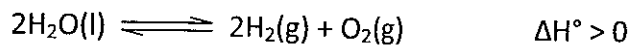
- b) [3 marks] How many kJ of heat will be liberated by the reaction of  $266.7 \text{ g}$  of  $\text{AlBr}_3$  ( $266.7 \text{ g/mol}$ ) with  $85.08 \text{ g}$  of  $\text{Cl}_2$  ( $70.9 \text{ g/mol}$ )?

$$266.7 \text{ g AlBr}_3 \times \frac{1 \text{ mol}}{266.7 \text{ g}} \times \frac{-327.04 \text{ kJ}}{2 \text{ mol AlBr}_3} = -163.52 \text{ kJ}$$

$$85.08 \text{ g Cl}_2 \times \frac{1 \text{ mol}}{70.9 \text{ g}} \times \frac{-327.04 \text{ kJ}}{3 \text{ mol Cl}_2} = -130.816 \text{ kJ}$$

So  $\boxed{130.816 \text{ kJ}}$  of heat will be liberated

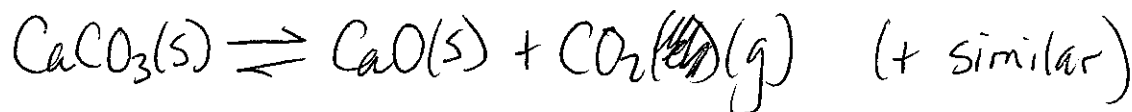
- 6) [4 marks] Predict the effect that each of the indicated changes would have on the equilibrium numbers of moles of  $\text{H}_2\text{O}(\text{l})$  and on the value of  $K_c$  for the reaction:



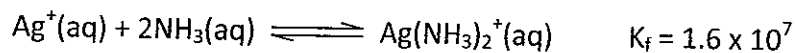
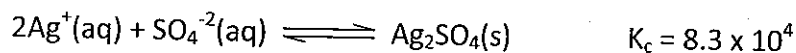
Your choices are Increase from the current value, Decrease from the current value, or Not Change the current value. Circle your choice. Unless explicitly stated otherwise, you may assume that all changes take place at constant temperature.

| Change:                                       | Effect on: |   |    |                                |   |    |
|---|------------|---|----|--------------------------------|---|----|
|   | $K_c$      |   |    | $\text{H}_2\text{O}(\text{l})$ |   |    |
| Add some $\text{H}_2(\text{g})$               | I          | D | NC | I                              | D | NC |
| Add some $\text{H}_2\text{O}(\text{l})$       | I          | D | NC | I                              | D | NC |
| Cool the reaction mixture                     | I          | D | NC | I                              | D | NC |
| Decrease the volume of the reaction container | I          | D | NC | I                              | D | NC |

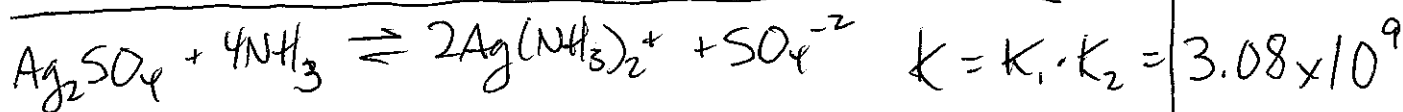
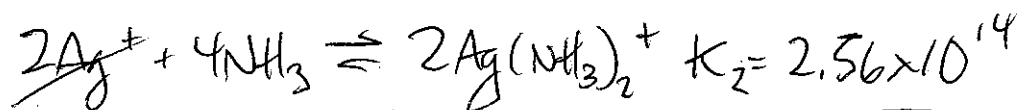
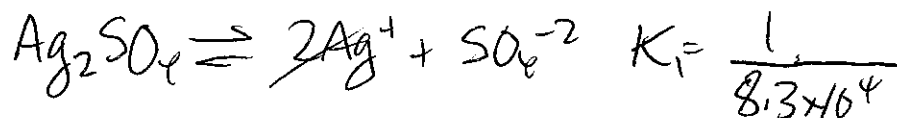
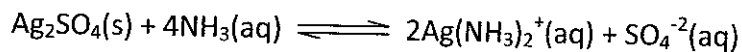
- 7) [1 mark] Write one possible reaction for which the equilibrium expression is  $[\text{CO}_2]_e$ .



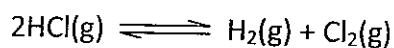
8) [3 marks] Given the two equilibria:



Calculate the value of  $K_c$  for the equilibrium:



9) [2 marks] 0.300 moles of  $\text{H}_2$  and 0.450 mole of  $\text{Cl}_2$  were placed in a 3.00 L flask and heated. At equilibrium, 60% of the  $\text{Cl}_2$  had reacted. Calculate  $K_c$  for



$$\begin{array}{cccc} \text{i} & 0 & 0.1 & 0.15 \end{array}$$

$$\text{c} + 1.2(0.15) \quad - 0.6(0.15) \quad 0.6(0.15)$$

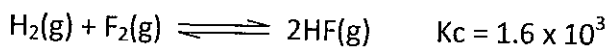
$$\text{e} \quad 0.18 \quad \cancel{0.10} \quad 0.06$$

$$0.01$$

$$K_c = \frac{(0.01)(0.06)}{(0.18)^2}$$

$$= 1.85 \times 10^{-2}$$

10) [4 marks total] For the system:



If 1.00 mole of HF and 0.500 moles each of  $\text{H}_2$  and  $\text{F}_2$  are placed in a 10.0 L container and allowed to come to equilibrium:

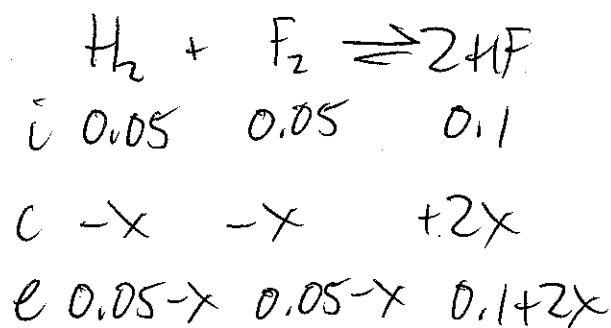
a) [1 mark] In which direction will the reaction shift to attain equilibrium? How do you know?

$$[\text{H}_2] = [\text{F}_2] = \frac{0.5}{10} = 0.05 \text{ M} \quad Q = \frac{(0.1)^2}{(0.05)^2} = 4$$

$$[\text{HF}] = \frac{1}{10} = 0.1 \text{ M}$$

$$Q < K, \text{ rxn} \rightarrow$$

b) [3 marks] What will be the concentrations of all the gases at equilibrium?



$$\frac{(0.1+2x)^2}{(0.05-x)^2} = 1.6 \times 10^3$$

$$\frac{(0.1+2x)}{(0.05-x)} = 40$$

$$0.1 + 2x = 2 - 40x$$

$$42x = 1.9$$

$$\Rightarrow x = 0.0452$$

$$\Rightarrow [\text{HF}]_e = 0.1 + 2x = 0.1905 \text{ M}$$

$$[\text{H}_2]_e = [\text{F}_2]_e = 0.05 - x = 0.004762 \text{ M}$$