CHEM1011 WORKSHEET A

SAMPLE EXAM QUESTIONS

EQUILIBRIUM

Question 1.

(a) Consider the following reaction:

$Na_2CO_3(s) + 2HOAc(aq) == 2Na^+(aq) + 2OAc^-(aq) + H_2O(l) + CO_2(g)$

(i) Write an equilibrium constant expression.

$$K = \frac{[Na^+]^2[OAc^-]^2[CO_2]}{[HOAc]^2}$$

(ii) What would be the effect on the position of equilibrium and the equilibrium constant of the following changes:

- adding extra acetic acid (HOAc)

adding more reactant, so reaction to more product, K_c unchanged

- increasing the carbon dioxide pressure

adding more product, so reaction to more reactants, K_c unchanged

- doubling the volume of the solution by adding water at constant carbon dioxide pressure

K_c and p_{CO_2} remain unchanged;
doubling the volume halves all solution concentrations, therefore:

$$Q_c = \frac{[Na^+]^2[OAc^-]^2[CO_2]}{[HOAc]^2} \cdot (½)^2\cancel{(½)^2}(1)/\cancel{(½)^2} = 0.25\ K_c\ .$$

Q* < *K, *so reaction occurs to give* ***more products***

K_c (or K_p) change only if the temperature changes

Question 2.

(a) Suggest four ways that the equilibrium concentration of ammonia can be increased in a reaction vessel in which the equilibrium is described by:

$N_2(g) + 3\ H_2(g) == 2\ NH_3(g) \qquad \Delta_r H = -92\ \text{kJ mol}^{-1}$

• *adding N_2* • *adding H_2* • *decreasing vessel size*
• *removing heat (since exothermic; heat given out by the reaction)*

(b) Given the equilibrium constant for the system

$H_2\,(g) + I_2\,(g) == 2\ HI\,(g) \qquad K_c\ (490\ °C) = 45.9$

(i) Calculate K_c' for the following reaction:

$HI\,(g) == ½\ H_2(g) + ½\ I_2(g)$

Reverse equilibrium reaction: $K' = 1/K$
Multiply equilibrium reaction by n: $K'' = K^n$

$K_c' = (K_c)^{-1/2} = (45.9)^{-1/2} = 0.148$

(ii) Write an expression for K_p' for the reaction given immediately (above in part (i)).

$$K_p = \frac{p(H_2)^{1/2}\, p(I_2)^{1/2}}{p\,(HI)} = K_c' \left(\frac{RT}{p^o}\right)^{\Delta n_{\text{gas}}}$$

(iii) Calculate K_p' for this reaction.

$\Delta n^{gas} = (½ + ½) - 1 = 0,$
so $K_p' = K_c' = 0.148$

(iv) 2.00 mol of HI is introduced into a 1.0 L flask at 490 °C. What will be the final concentrations in the flask at equilibrium?

	$HI\,(g)$	==	$½\ H_2(g)$	+	$½ I_2(g)$
Final conc's	$2.00 - x$		$½x$		$½x$

so $K_c' = (½x)^{1/2}(½x)^{1/2}/(2.00 - x) = 0.148$
rearrange $(½)x = 0.148 \times (2.00 - x)$
$x = (2.00 \times 0.148)/(0.5 + 0.148) = 0.457\ \text{mol L}^{-1}$

so $[H_2] = [I_2] = ½x = 0.23\ \text{mol L}^{-1}$
and $[HI] = 2.00 - 0.46 = 1.54\ \text{mol L}^{-1}$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$