CHM 1046 - JOENS WORKSHEET #5 Due date: M,W,F Wednesday, October 5
T,R Thursday, October 6

* While I prefer you turn in a hard copy of the worksheet, I will accept scanned copies sent to my email address, joensj@fiu.edu

Section: (circle one) M,W,F Tu,Tr

For problems involving calculations you must show your work for credit.

1) For each of the following reactions, write the appropriate expression for K_C, K_p, and K, or say an expression cannot be written.

a) $H_2(g) + I_2(s) \leftrightarrows 2\ HI(g)$ (1.1)

$K_C = \dfrac{[HI]^2}{[H_2]}$ $K_p = \dfrac{(p_{HI})^2}{(p_{H2})}$ $K = \dfrac{(p_{HI})^2}{(p_{H2})}$

Note that in this case $K = K_p$. Also note $I_2(s)$ does not appear in any of the expressions, as it is a solid.

b) $2\ Na(s) + 2\ H_2O(\ell) \leftrightarrows 2\ Na^+(aq) + 2\ OH^-(aq) + H_2(g)$ (1.2)

$K_C = [Na^+]^2\ [OH^-]^2\ [H_2]$ K_p = n/a $K = [Na^+]^2\ [OH^-]^2\ (p_{H2})$

K_p is not defined because some of the concentrations appearing in K_C are not gases. In this example, K is not equal to either K_C or K_p. Na(s) and $H_2O(\ell)$ do not appear in any of the expressions, as they are a solid (Na(s)) and a liquid ($H_2O(\ell)$).

2) For the reaction

$CO(g) + 2\ H_2(g) \leftrightarrows CH_3OH(g)$ (2.1)

the equilibrium concentrations observed in a system at T = 1000. K were $[H_2]$ = 0.250 mol/L, [CO] = 0.110 mol/L, and $[CH_3OH]$ = 0.00260 mol/L. What are the numerical values for K_C and K_p at this temperature?

$K_C = \dfrac{[CH_3OH]}{[CO]\ [H_2]^2}$ $K_p = \dfrac{(p_{CH3OH})}{(p_{CO})\ (p_{H2})^2}$

So $K_C = \dfrac{(0.00260)}{(0.110)\ (0.250)^2} = 0.378$

For K_p, we use $K_p = K_C\ (RT)^{\Delta ng}$, where Δn_g = (moles of gas as products) – (moles of gas as reactants)

= 1 – 3 = – 2

So $K_p = (0.378)\ [\ (0.08206)\ (1000)\ ]^{-2} = (0.378)\ (1.49 \times 10^{-4}) = 5.6 \times 10^{-5}$

Note that when we use the above expression, we must use R = 0.08206, and T = the temperature (in Kelvin), so 1000. Note these are used without units as equilibrium constants do not have units.

3) The equilibrium constant for the reaction

$$H_2(g) + I_2(g) \leftrightarrows 2\ HI(g) \qquad (3.1)$$

is $K_C = 49.5$ at T = 440. °C.

In a particular system at T = 440. °C the initial concentrations present are $[H_2] = 0.140$ M, $[I_2] = 0.000$ M, and [HI] = 0.080 M. What concentrations will be present when equilibrium is reached?

$$K_C = \frac{[HI]^2}{[H_2]\,[I_2]} = 49.5$$

This is an equilibrium problem, and so we need to set up an ICE table.

	Initial	Change	Equilibrium
HI	0.080	– 2x	0.080 – 2x
H_2	0.140	x	0.140 + x
I_2	0.000	x	x

Substituting into the expression for K_C gives

$$\frac{(0.080 - 2x)^2}{(0.140 + x)\,(x)} = 49.5$$

We can try assuming x << 0.080. Then

$$\frac{(0.080)^2}{(0.140)\,(x)} = 49.5 \qquad x = \frac{(0.080)^2}{(0.140)\,(49.5)} = 9.2 \times 10^{-4}$$

So it turns out that assuming x is small is a good assumption.

[HI] = 0.0782 M $\qquad [H_2] = 0.141$ M $\qquad [I_2] = 9.2 \times 10^{-4}$ M

If we do the problem exactly, using the quadratic equation, then

$$(0.080 - 2x)^2 = 49.5\,(0.140 + x)\,(x)$$

$$0.0064 - 0.32\,x + 4x^2 = 6.93\,x + 49.5\,x^2$$

So $\qquad 45.5\,x^2 + 7.25\,x - 0.0064 = 0$

$$x = \frac{-7.25 \pm [\,(7.25)^2 - 4\,(45.5)\,(-0.0064)\,]^{1/2}}{2\,(45.5)} = \underline{8.8 \times 10^{-4}}\,,\ -0.160$$

The underlined root is the one that gives all positive concentrations. So at equilibrium

[HI] = 0.0782 M $\qquad [H_2] = 0.141$ M $\qquad [I_2] = 8.8 \times 10^{-4}$ M

Notice that the exact (quadratic) method gives a slightly lower value for $[I_2]$ than the approximate method, but the same concentrations (to three significant figures) for [HI] and $[H_2]$.

4) Which of the following will cause a change in the numerical value for the equilibrium constant for a reaction?

a) A change in pressure
b) A change in temperature
B c) A change in volume
d) Both a and c
e) Both a and b and c