## **Chapter 9 - Lecture Worksheet 4**

- 1. Summarize the ICE Method
- 1. Write the *Balanced Chemical Equation*. (always start with this for any equilibrium calculation !)
- 2. Write the *Initial Concentrations* under each substance. (may be in units of M, or partial pressures for gases)
- 3. Write the *Change in Concentrations* in terms of x. (moles/L of reactant lost)
- 4. Write the *Equilibrium Concentrations* in terms of initial concs and x.
- 5. Put into *K* expression. (equation relating K to equilibrium concentrations)
- 6. Calulate K or Solve for x. (Depends on what you are looking for. Is your answer physically reasonable ?)
- 7. Check !! (Check approximation or exact calculation.)
- 2. The equilibrium constant,  $K_c$  for the following reaction is 5.9 x 10<sup>-3</sup> at 25<sup>o</sup>C. Suppose 0.34 moles of  $N_2O_4(g)$  are placed in a 1.00 L flask. What is the equilibrium concentration of  $NO_2(g)$ ? You must use the ICE method to solve this problem.

	$N_2O_4(g)$	1	2 NO <sub>2</sub> (g)
Initial	0.34 M		0
Change	<u>—X</u>		<u>+2x</u>
Equilibrium	0.34 - x		2x

$$K_c = \frac{[NO_2(g)]^2}{[N_2O_4(g)]} = \frac{(2x)^2}{(0.34 - x)} = 5.9 \times 10^{-3}$$

This is a quadratic equation. Put into form  $ax^2 + bx + c = 0$ 

$$4x^{2} + K_{c}x - K_{c}(0.34) = 0$$

$$x = \frac{-b \pm \sqrt{b^{2} - 4ac}}{2a} = \frac{-K_{c} \pm \sqrt{K_{c}^{2} - 4(4)(-K_{c}(0.34))}}{2(4)}$$

$$= \frac{-(5.9x10^{-3}) \pm \sqrt{(5.9x10^{-3})^{2} + (16)(5.9x10^{-3})(0.34)}}{8}$$

 $x = -2.12x10^{-2}$  (phycically unreasonable) or  $+2.17x10^{-2}$  M (correct answer) [NO<sub>2</sub>] = 2x = 0.0434 = 0.043 M

Check (use 3 sig figs):

 $\frac{(2x)^2}{(0.34-x)} = \frac{(0.0434)^2}{(0.34-0.0217)} = 0.0059$  to two significant figures. Check !

- **3.** The principle industrial source of hydrogen gas is from natural gas and water via a two step process:
- 1. Reforming Reaction:  $CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3H_2(g)$
- 2. Shift reactions:  $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$



The equilibrium constant,  $K_p$  for the first reaction is 1.8 x 10<sup>-7</sup> at 600 K. Suppose 1.40 atm of CH<sub>4</sub>(g) and 2.30 atm of H<sub>2</sub>O(g) are placed in a reaction chamber. What will the equilibrium partial pressure of H<sub>2</sub>(g) be after the first reaction ?

You must use the ICE method to solve this problem.

	CH <sub>4</sub> (g)	+ $H_2O(g)$	1	CO(g)	+ $3H_2(g)$
Initial	1.40 atm	2.30 atm		0	0
Change	<u>—X</u>	<u>—X</u>		<u>+x</u>	<u>+3x</u>
Equilibrium	1.40 – x	2.30-x		Х	3x

$$K_{p} = \frac{P_{CO}P_{H_{2}}^{3}}{P_{CH_{4}}P_{H_{2}O}} = \frac{x(3x)^{3}}{(1.40 - x)(2.30 - x)} = \frac{27x^{4}}{(1.40 - x)(2.30 - x)} = 1.8 \times 10^{-7}$$

This is a *QUARTIC* equation. Can't put into simple form to solve.

 $\frac{\text{APPROXIMATE}}{\text{K}_{p}} << 1 \text{ Reactant favored. Therefore x is SMALL}$ 

BIG NUMBER – SMALL NUMBER **≈** BIG NUMBER

Kp expression simplifies  

$$K_p = 1.8x10^{-7} \approx \frac{27x^4}{(1.40)(2.30)}$$

$$\mathbf{x} = \sqrt[4]{\frac{(1.40)(2.30)K_p}{27}} = \sqrt[4]{2.15x10^{-8}} = 1.21x10^{-2} = 1.2 \times 10^{-2} \text{ atm}$$

 $\frac{CHECK}{Substitute in x, see if get correct K_p}$ 

$$\frac{27(1.21x10^{-2})^4}{(1.40 - 1.21x10^{-2})(2.30 - 1.21x10^{-2})} = 1.8x10^{-7}$$

*Not Bad* for an approximate answer. Sure simplifies the mathematics !!! We will use this technique frequently for Acid/Base calculations.