Honors Chemistry Notes

Name: Key Date: ______ Mods: ______

Unit 14: Equilibrium Calculations Using the "ICE" Tables

Type 1: Calculating K_{eq} from Initial and Equilibrium Concentrations

These kinds of problems will give at least one equilibrium concentration however, the value of the equilibrium constant, K_{eq}, will NOT be given. Make sure you have a BALANCED chemical reaction to begin!

1) **Row 1 - INITIAL:** Any given *initial concentrations* (or partial pressures) of REACTANTS will transferred into the "I" row of the ice table. **Note that the *initial* concentrations of all PRODUCTS in a reaction will always be ZERO!!!**
2) **Row 3 - EQUILIBRIUM:** If any *equilibrium concentrations* are given, transfer these values into the "E" row.
3) **Row 2 - CHANGE:** Recall that initially (before equilibrium is established), the reactants are USED UP as the reaction progresses and the products are FORMED, therefore the change in **REACTANTS in always NEGATIVE**, whereas the change in **PRODUCTS is always POSITIVE**!!!!!
   - For reactants/products in which **BOTH the *equilibrium* and *initial* concentrations are both known**, the *change* in concentration can be calculated: **[*change* in concentration = equilibrium concentration – *initial* concentration]**
   - For reactants/products in which **ONLY the *initial* concentrations are known**...
     - Begin with the *change in concentrations* which is KNOWN and **use molar ratios** from the balanced equation to determine what the ***change in concentration* is equal to for all the missing reactants/products**
4) Now that the *change* in concentration is determined for every reactant/product, **calculate all the *equilibrium* concentrations in the ice table. [*equilibrium* concentration = *initial* +/– *change*]**
5) **Write the equilibrium expression** from the balanced reaction and **solve for the equilibrium constant, K_{eq}** by plugging in all the *equilibrium* concentrations.
   - Note that no pure solids or liquids will appear in ICE tables because they do not appear equilibrium expressions.

Example #1: A closed system initially containing 1.12 atm NH_3 at 448°C is allowed to reach equilibrium. Analysis of the equilibrium mixture shows that the pressure of H_2 is 0.86 atm. Calculate K_{eq} at 448°C for this reaction:

$\underline{2}\ NH_3(g) \leftrightharpoons \underline{1}\ N_2(g) + \underline{3}\ H_2(g)$

	$2\ NH_3$ $\rightleftharpoons$	$1\ N_2$ +	$3\ H_2$
I	1.12 atm	0	0
C	−0.573 atm	+ 0.287 atm	+ 0.86 atm
E	0.547 atm	0.287 atm	0.86 atm

$$K_{eq} = \frac{[N_2][H_2]^3}{[NH_3]^2}$$

$$= \frac{[0.287\text{ atm}][0.86\text{ atm}]^3}{[0.547\text{ atm}]^2}$$

$$= 0.61$$

Changes:

★ $\frac{0.86\text{ atm } H_2 \,|\, 2\text{ mols } NH_3}{\,|\, 3\text{ mols } H_2}$

C: 0.573 atm NH_3

★ $\frac{0.86\text{ atm } H_2 \,|\, 1\text{ mol } N_2}{\,|\, 3\text{ mols } H_2}$

C: 0.287 atm

Type 2: Calculating Equilibrium Concentrations from Initial Concentrations and K_{eq}

These kinds of problems will give the value of the equilibrium constant, K_{eq}, in the problem, but NONE of the equilibrium concentrations are given. Make sure you have a BALANCED chemical reaction to begin!

1) **Row 1 - INITIAL:** Any given *initial concentrations* (or partial pressures) of REACTANTS will transferred into the "I" row of the ice table. **Note that the *initial* concentrations of all PRODUCTS in a reaction will always be ZERO!!!**
2) **Row 2 - CHANGE:** Since no *equilibrium* concentrations are known, the *change* in concentrations are also unknown.
   - Therefore, in the "C" row of the ice table, *changes in concentration* will be given in terms of the variable "x". The value of "x" is based on the molar ratios from the balanced equation so the multiple of "x" (1x, 2x, 3x, etc) is based on the balanced coefficient in front of each reactant and product.
   - Recall that initially (before equilibrium is established), the reactants are USED UP as the reaction progresses and the products are FORMED, therefore the "x" change in **REACTANTS in always NEGATIVE**, whereas the "x" change in **PRODUCTS is always POSITIVE**!!!!!!
3) **Row 3 - EQUILIBRIUM:** For the "E" row in the ice table, the *equilibrium* concentration is equal to the *initial* concentration plus/minus (depending if it is a product or reactant) the "x" *change* in concentration.
4) **Write the equilibrium expression** from the balanced reaction and **set it equal to the K_{eq} value** given in the problem. Plug the *equilibrium* concentrations (given in terms of "x") into the equilibrium expression and solve the equation for the value of "x".
   - **Recall that if x^2 is in the expression, you must use the quadratic formula** to solve.
     - EXCEPTION: If $K_{eq} < 1 \times 10^{-5}$ then the *change* in concentration is deemed so small that it is negligible and the "x" values can be ignored. You will not see this for these problems, be we will in future classes.

$$ax^2+bx+c=0$$

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

5) Once the value of "x" is known, **solve for all the numerical values of the *equilibrium* concentrations.**
   - Note that if there are **two possible values for "x"** only ONE ANSWER will be meaningful (an answer for "x" should be ignored if the value itself is negative OR if gives calculated *equilibrium* concentrations which are negative)

Example 2: A sealed flask is filled with 0.924 atm H_2 and 0.771 atm I_2 at 320°C. The value of the equilibrium constant K_{eq} for the reaction at this temperature is 2.05. What are the equilibrium pressures of H_2, I_2, and HI?

$\underline{1}\ H_2(g) + \underline{1}\ I_2(g) \leftrightharpoons \underline{2}\ HI(g)$

$1\ H_2 + 1\ I_2 \rightleftharpoons 2\ HI$

I	0.924 atm	0.771 atm	0
C	− x	− x	+ 2x
E	0.924 − x	0.771 − x	2x

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

$$2.05 = \frac{[2x]^2}{[0.924-x][0.771-x]}$$

$$2.05 = \frac{4x^2}{0.712-1.695x+x^2}$$

$$2.05(0.712-1.695x+x^2) = 4x^2$$

$$1.46-3.475x+2.05x^2 = 4x^2$$

$$1.46-3.475x-1.95x^2 = 0$$

$a = -1.95$

$b = -3.475$

$c = 1.46$

you must SHOW this work!

$$x = \frac{-(-3.475) \pm \sqrt{(-3.475)^2-4(-1.95)(1.46)}}{2(-1.95)}$$

$$x = \frac{3.475+\sqrt{23.464}}{-3.9}$$

$x = -2.133$ NOT VALID

$$x = \frac{3.475-\sqrt{23.464}}{-3.9}$$

$x = 0.351$ ✱

Equilibrium Pressures: $H_2 = 0.573$ atm, $I_2 = 0.42$ atm, $HI = 0.702$ atm

https://www.youtube.com/watch?v=W2oL52t0NVo

↳ How to use graphing calculator to solve.