Chemistry 30 Unit 4 - Acid Base / Equilibrium

Equilibrium Law II: ICE Tables

Review:

a) What is the value of the equilibrium constant for the decomposition of phosphorous pentachloride gas to phosphorous trichloride gas and chlorine gas, at 200°C?

Evidence: $[PCl_{3(g)}] = [Cl_{2(g)}] = 0.014$ mol/L
$[PCl_{5(g)}] = 4.3 \times 10^{-4}$ mol/L

b) Judging from the K_{eq}, are products or reactants favored in this equilibrium?

Important Note:

Heterogeneous vs. Homogenous Systems

You might run into these two terms while working problems or in the text...

Heterogeneous Systems: Are systems which are made up of two different states of matter (i.e. aqueous and gas).

Homogenous Systems: Are systems made up of only one state of matter (i.e. all gases, all aqueous, etc).

Extensions of the Equilibrium Law: ICE tables

Last day, we only looked at problems where the system was already at equilibrium.

We must also study problems where the initial conditions (i.e. initial concentrations before equilibrium) are given.

To help organize questions like this, it is useful to construct an ICE table:

I = [initial]
C = [change]
E = [equilibrium]

ex) For the reaction:

$$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$$

The initial $[H_2]$ and $[I_2]$ is 1.00 mmol/L. The equilibrium $[I_2]$ is 0.22 mmol/L. What is the equilibrium $[H_2]$ and of [HI]?

1. Make an ICE table, fill in what you know.

	$[H_2]$	$[I_2]$	[HI]
I	1.00	1.00	0
C			
E			

Initially, there is 1.00 mmol/L of the reactants. There are no products yet because the initial time is before the reaction begins.

	$[H_2]$	$[I_2]$	[HI]
I	1.00	1.00	0
C			
E		0.22	

We know the equilibrium $[I_2]$ = 0.22 mmol/L.

2. Calculate the change in [ ] of each species.

	$[H_2]$	$[I_2]$	[HI]
I	1.00	1.00	0
C		-0.78	
E		0.22	

$[I_2]$ changes by 1.00 - 0.22 = 0.78 mmol.

We write decreases in [ ] as -ive numbers and increases as +ive numbers.

	$[H_2]$	$[I_2]$	[HI]
I	1.00	1.00	0
C	-0.78	-0.78	+1.56
E		0.22	

$$H_{2(g)} + I_{2(g)} \rightleftharpoons 2HI_{(g)}$$

Recall from the balanced equation that there is a 1:1:2 ratio. That means that for every 1 mmol of I_2, 1 mmol of H_2 react. Twice as much HI reacts because there is a 2 in front of that species.

The change in [HI] is also +ive because it is an increase in [ ].

3. Do the math to fill in the bottom row.

	$[H_2]$	$[I_2]$	[HI]
I	1.00	1.00	0
C	-0.78	-0.78	+1.56
E	0.22	0.22	1.56

The equilibrium $[H_2]$ = 0.22 mmol/L and [HI] = 1.56 mmol/L.

ex. 2.00 mol of methane gas is initially added to 10.00 mol of chlorine gas. At equilibrium the system contains 1.40 mol of chloromethane, $CH_3Cl_{(g)}$ and some $HCl_{(g)}$.

a) What amounts of reactants and products are present at equilibrium?
b) What is the K_{eq} for this reaction?

I				
C				
E				

ex. At high temperature, 1.00 mol of $HBr_{(g)}$ is introduced into a 2.00 L container. After a decomposition reaction, equilibrium is reached and the [HBr] is found to be 0.100 mol/L.

a) Calculate the amounts of all species present at equilibrium.
b) Find the K_c for this reaction.

I			
C			
E			

More different examples: Combining equilibrium law and ICE tables:

ex) In a 500 mL container, 2.00 mol of $CO_{(g)}$ is combined with 2.00 mol $H_2O_{(g)}$ to produce $CO_{2(g)}$ and $H_{2(g)}$. The K_c for this reaction is 4.20 at 900°C. What [ ] of each substance will be present at equilibrium?

1. Build an ICE table and fill in as much as you can.

n = 2.00 mol / 0.5 L = 4.00 mol/L

	CO	H_2O	CO_2	H_2
I	4.00	4.00	0	0
C				
E				

Hint: initial [products] is always 0.

2. Because we don't have any equilibrium values and we don't know what the change is, just call the change "x" for now.

$$CO_{(g)} + H_2O_{(g)} \rightleftharpoons CO_{2(g)} + H_{2(g)}$$

	CO	H_2O	CO_2	H_2
I	4.00	4.00	0	0
C	-x	-x	+x	+x
E				

Each species increases or decrease by x because it is a 1:1:1:1 ratio.

3. Do the math to work out the equilibrium [ ].

	CO	H_2O	CO_2	H_2
I	4.00	4.00	0	0
C	-x	-x	+x	+x
E	4.00 - x	4.00 - x	+x	+x

4. Sub in the equilibrium [ ]'s into the Equilibrium Law expression and solve for x.

$$4.20 = \frac{[CO_2]\,[H_2]}{[CO]\,[H_2O]}$$

$$4.20 = \frac{(x)(x)}{(4.00 - x)(4.00 - x)}$$

$$4.20 = \frac{(x)^2}{(4.00 - x)^2}$$

$$\sqrt{4.20} = \frac{(x)}{(4.00 - x)}$$

$$2.05 = \frac{(x)}{(4.00 - x)}$$

$$(2.05)(4.00 - x) = x$$
$$8.20 - 2.05x = x$$
$$8.20 = 3.05x$$
$$\underline{\underline{2.69}} = x$$

5. Fill in the ICE table.

	CO	H_2O	CO_2	H_2
I	4.00	4.00	0	0
C	-2.69	-2.69	+2.69	+2.69
E	1.31	1.31	2.69	2.69

The equilibrium [CO] = [H_2O] = 1.31 mol/L
[CO_2] = [H_2] = 2.69 mol/L

ex) Consider the system:
$2HBr_{(g)} \rightleftharpoons H_{2(g)} + Br_{2(g)}$

Initially, 0.25 mol of hydrogen and 0.25 mol bromine gas are placed in a 500 mL flask. The K_c for the reaction at the temperature used is 0.020.

Find the concentration of each substance at equilibrium.

HW: UA