

Equilibrium Law II: ICE Tables



Important Note:

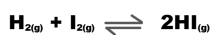
Heterogeneous vs. Homogenous Systems

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).

ex) For the reaction:



The initial $[\text{H}_2]$ and $[\text{I}_2]$ is 1.00 mmol/L. The equilibrium $[\text{I}_2]$ is 0.22 mmol/L. What is the equilibrium $[\text{H}_2]$ and of $[\text{HI}]$?

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

	$[\text{H}_2]$	$[\text{I}_2]$	$[\text{HI}]$
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.

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: $[\text{PCl}_{5(g)}] = [\text{Cl}_{2(g)}] = 0.014 \text{ mol/L}$
 $[\text{PCl}_{3(g)}] = 4.3 \times 10^{-4} \text{ mol/L}$

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

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]

	$[\text{H}_2]$	$[\text{I}_2]$	$[\text{HI}]$
I	1.00	1.00	0
C			
E		0.22	

We know the equilibrium $[\text{I}_2] = 0.22 \text{ mmol/L}$.

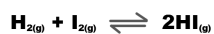
2. Calculate the change in [] of each species.

	$[\text{H}_2]$	$[\text{I}_2]$	$[\text{HI}]$
I	1.00	1.00	0
C		-0.78	
E		0.22	

$[\text{I}_2]$ 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	



Recall from the balanced equation that there is a 1:1:2 ratio. That means that for every 1 mmol of I₂, 1 mmol of H₂ 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 [].

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, CH₃Cl_(g) and some HCl_(g).

- What amounts of reactants and products are present at equilibrium?
- What is the K_{eq} 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 H₂O_(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.

$$n = 2.00 \text{ mol} / 0.5 \text{ L} = 4.00 \text{ mol/L}$$

	CO	H ₂ O	CO ₂	H ₂
I	4.00	4.00	0	0
C				
E				

Hint: initial [products] is always 0.

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

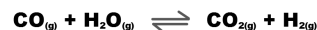
The equilibrium [H₂] = 0.22 mmol/L and [HI] = 1.56 mmol/L.

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.

- Calculate the amounts of all species present at equilibrium.
- Find the K_c for this reaction.

I			
C			
E			

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	H ₂ O	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 ₂ O	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.

$$4.20 = \frac{[\text{CO}_2][\text{H}_2]}{[\text{CO}][\text{H}_2\text{O}]}$$

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

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

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

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

$$(2.05)(4.00 - x) = x$$

$$8.20 - 2.05x = x$$

$$8.20 = 3.05x$$

$$\underline{2.69 = x}$$

5. Fill in the ICE table.

	CO	H ₂ O	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] = [H₂O] = 1.31 mol/L
[CO₂] = [H₂] = 2.69 mol/L

ex) Consider the system:



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