EQUILIBRIUM CALCULATIONS

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Two types of questions:

1. calculating K_{eq} from known concentration values

2. calculating concentration values when K_{eq} is given

EQUILIBRIUM CALCULATIONS

Calculating K_{eq}

Example #1

$$N_2O_{4(g)} <===> 2\ NO_{2(g)}$$

At 25°C, the equilibrium concentrations are:

$[N_2O_4] = 0.0292$ mol / L

$[NO_2] = 0.0116$ mol / L

Calculate K_{eq} at 25°C.

$$K_{eq} = \frac{[NO_2]^2}{[N_2O_4]}$$

$$= \frac{[0.0116]^2}{[0.0292]}$$

$$= 4.61 \times 10^{-3}$$

$$\therefore k_{eq} = 4.61 \times 10^{-3}$$

EQUILIBRIUM CALCULATIONS

Calculating K_{eq}

Example #2

$$N_{2(g)} + 3\ H_{2(g)} <====> 2\ NH_{3(g)}$$

At 200°C, the concentrations at equilibrium are:

$[N_2] = 2.12$, $[H_2] = 1.75$, and $[NH_3] = 84.3$

Calculate K_{eq} at 200°C.

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

$$= \frac{[84.3]^2}{[2.12][1.75]^3}$$

$$= 625$$

.: the k_{eq} is 625

EQUILIBRIUM CALCULATIONS

Calculating K_{eq}

Example #3

$$H_{2(g)} + I_{2(g)} <====> 2\ HI_{(g)}$$

Initial moles in a 2.00 L flask:
$[H_2]$ = 0.200 mol and $[I_2]$ = 0.200 mol

At equilibrium, $[I_2]$ = 0.020 mol / L.

a)What is K_{eq} at steady conditions?
b)What percent of iodine vapour reacted?

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ICE tables

Whenever questions involve initial conditions changing to reach equilibrium, ICE tables are a good method to organize your information.

I = initial concentrations

C = change in concentrations

E = equilibrium concentrations

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Calculating K_{eq} - Example #3

Initial moles in a 2.00 L flask: $[H_2] = 0.200$ mol and $[I_2] = 0.200$ mol
At equilibrium, $[I_2] = 0.020$ mol / L.

	$H_{2(g)}$	+	$I_{2(g)}$	<=>	$2\ HI_{(g)}$
I	$^{0.2/2\ =}$ 0.1M		0.1M		0.0M
C	**-x**		**-x**		**2x**
E	0.02M		0.02M		0.16M

a) What is K_{eq} at steady conditions?
b) What percent of iodine vapour reacted?

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

$= \frac{[0.16]^2}{[0.02][0.02]}$

$= 64$

$\therefore K_{eq} = 64$

b) $\%I_2\ _{\text{reacted}} = \frac{0.08}{0.1} \times 100\%$

$= 80\%$

$\therefore \%I_2$ reacted $= 80\%$

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% Reaction

$$\% \text{ reaction} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

- Actual yield: measured @ equilibrium
- Theoretical yield:
  - Maximum possible yield
  - Calculated with stoichiometry
  - Based on assumptions of forward reaction

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Calculating K_{eq}

Example #4

2.00 mol of HI in 2.00 L flask at 425°C react to produce H_2 and I_2. At equilibrium, $[H_2]$ and $[I_2]$ = 0.214 mol / L.

What is K_{eq} for this reaction?

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Calculating K_{eq} - Example #4

2.00 mol of HI in 2.00 L flask at 425°C react to produce H_2 and I_2. At equilibrium, $[H_2]$ and $[I_2]$ = 0.214 mol / L.

What is K_{eq} for this reaction?

	$2\ HI_{(g)}$	<=>	$H_{2(g)}$	+	$I_{2(g)}$
I	1M		0M		0M
C	**-2x**		**+x**		**+x**
E	**1-2x**		0.214M		0.214M

1-2x

1-2(0.214) = 0.572

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

$$= \frac{[0.214][0.214]}{[0.572]^2}$$

$$= 0.140$$

.: the K_{eq} is 0.140

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Calculating K_{eq}

Example #5

$$N_{2(g)} + 3\ H_{2(g)} <====> 2\ NH_{3(g)}$$

Initial concentrations:

$[N_2] = 0.32$ M and $[H_2] = 0.66$ M

What is K_{eq} when equilibrium $[H_2]$ is 0.30 M?

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Calculating K_{eq} - Example #5

Initial concentrations: $[N_2] = 0.32$ M and $[H_2] = 0.66$ M

What is K_{eq} when equilibrium $[H_2]$ is 0.30 M?

	$N_{2(g)}$ +	$3\ H_{2(g)}$ <=>	$2\ NH_{3(g)}$
I	0.32M	0.66M	0
C	-x	-3x	+2x
E	0.20	0.30M	0.24

$0.66-3x = 0.30$

$0.66-0.30 = 3x$

$x = 0.12$

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

$$= \frac{[0.24]^2}{[0.20][0.30]^3}$$

$$= 11$$

$\therefore k_{eq} = 11$

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Calculating K_{eq} - Example #6

Calculate K_{eq}.

$$2\ C_{(s)} + O_{2(g)} \text{ <====> } 2\ CO_{(g)}$$

I	10	8	20
C	N/A	+7	-14
E		15	6

$8-x = 15$
$x = 8-15$
$x = -7$

$$K_{eq} = \frac{[CO]^2}{[O_2]}$$
$$= \frac{[6]^2}{[15]}$$
$$= 2.4$$

$\therefore k_{eq} = 2$