VHEMBE TEACHER DEVELOPMENT WORKSHOP
ONLINE LEARNING: 07 – 09 FEB 2022

Chemical equilibrium

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Chemical Equilibrium

- Define Chemical Dynamic Equilibrium
- Differentiate between $\rightarrow and \leftrightarrows$
- State Le Chatelier's’ Principle
- Info extracted from a given graph
- Use Le Chatelier's Principle to explain (P, T and C)
- Calculate Kc value
- Kc at a certain temperature
- Determine if the reaction is exothermic or endothermic
- Initial mol/conc./mass Reactants or products (RICE)
- Amount in the change (RICE)
- Equilibrium conc. of reactants or products (RICE)

Chemical Equilibrium

The dissociation of iodine molecules to iodine atoms (I) is a reversible reaction taking place in a sealed container at 727 °C. The balanced equation for the reaction is:

$$I_2(g) \rightleftharpoons 2I(g)$$

K_c for the reaction at 727 °C is 3,76 x 10^{-3}.

6.1 Write down the meaning of the term *reversible reaction*. (1)

Solution:

- Products can be converted to reactants.
- Both forward and reverse reactions can take place
- A reaction can take place in both directions

$$I_2(g) \rightleftharpoons 2I(g)$$

6.2 At equilibrium the pressure of the system is increased by decreasing the volume of the container at constant temperature.

How will EACH of the following be affected? Choose from INCREASES, DECREASES or REMAINS THE SAME.

6.2.1 The value of the equilibrium constant (1)

6.2.2 The number of I_2 molecules (1)

solutions:

6.2.1 Remain the same
(Kc changes with increasing or decreasing Temperature). In this case the Temperature is unchanged therefore Kc remain unchanged.

6.2.2 Increase
Pressure increases, the volume decreases. On the RHS there's 1 mole of I_2 and the LHS 2 mole of I. if the volume is small only 1 mole can be accommodates in the that space. In this case I_2 can be seen. Hence an increase in $[I_2]$ or rather the equilibrium shifts to the left and favours the formation of $[I_2]$.

$$I_2(g) \rightleftharpoons 2I(g)$$

6.3 Explain the answer to QUESTION 6.2.2 by referring to Le Chatelier's principle. (2)

solutions:

6.3 When pressure is decreases, the volume increases and vice-versa.

- In this case the pressure was increases.
- Volume decreases letting the side with smaller molecules to be accommodated or favoured.
- Therefore the reverse reaction will be favoured.

$$I_2(g) \rightleftharpoons 2I(g)$$

6.4 At 227 °C, the K_C value for the reaction above is 5,6 x 10^{-12}.

Is the forward reaction ENDOTHERMIC or EXOTHERMIC?
Fully explain the answer. (4)

solutions:

6.4 Endothermic.

- At **227°C** Kc = **5.6 x 10^{-12}** and **727°C** Kc = **3.76 x 10^{-3}**
- Kc decrease with a decrease in Temperature and likewise.
- But reverse reaction is favoured.
- From **227°C** to **727°C** there's an increase in temperature.

(An increase in Temperature results in an Endothermic reaction)

6.5 A certain mass of iodine molecules (I_2) is sealed in a 12,3 dm^3 flask at a temperature of 727 °C (K_c = 3,76 x 10^{-3}).

When equilibrium is reached, the concentration of the iodine atoms is found to be 4,79 x 10^{-3} $mol \cdot dm^{-3}$. Calculate the INITIAL MASS of the iodine molecules in the flask. (9)

The **RICE** solution problem

Tips:

- *RICE table can only work with moles and/or concentration.*
- *Remember to convert the moles to concentration at equilibrium.*
- *Kc only works with concentrations not moles.* $Kc = \frac{[products]}{[reactants]}$

What is **RICE?**
R = reaction
I = initial
C = change
E = equilibrium

R	$I_2(g)$ ⇋	$2I(g)$
I	[Y]	0
C	−x	+2x
E	[Y] – x	2x

Map/Route

R	$I_2(g)$ ⇋	$2I(g)$
I	[Y]	0
C	–x	+2x
E	[Y] – x	2x

R	$I_2(g)$ ⇋	$2I(g)$
I	[Y]	0
C	–x	+2x
E	[Y] – x	**$4.79x10^{-3}$**

R	$I_2(g)$ ⇋	$2I(g)$
I	[Y]	0
C	–x	**$4.79x10^{-3}$**
E	[Y] – x	$4.79x10^{-3}$

R	$I_2(g)$ ⇋	$2I(g)$
I	[Y]	0
C	**$2.40x10^{-3}$**	$4.79x10^{-3}$
E	[Y] – x	$4.79x10^{-3}$

R	$I_2(g)$ ⇋	$2I(g)$
I	[Y]	0
C	$2.40x10^{-3}$	$4.79x10^{-3}$
E	**[unknown]**	$4.79x10^{-3}$

$$Kc = \frac{[products]}{[reactants]}$$

$$Kc = \frac{[I]^2}{[I_2]}$$

$$[3.76x10^{-3}] = \frac{[4.79x10^{-3}]^2}{[I_2]}$$

$[I_2] = 6.10x10^{-3}$ Mol.dm^{-3}

R	$I_2(g)$ ⇋	$2I(g)$
I	[Y]	0
C	$2.40x10^{-3}$	$4.79x10^{-3}$
E	**$6.10x10^{-3}$**	$4.79x10^{-3}$

Find [Y]:

[Y] - $2.40x10^{-3}$ = $6.10x10^{-3}$

[Y] - $2.40x10^{-3}$ = $6.10x10^{-3}$

[Y] = $8.50x10^{-3}$ mol.dm^{-3}

[Y] = initial conc. of I_2

$$n = CV$$

n = 8.50x10–3x 12.3

$$n = 0.1046 \text{ mol}$$

$$m = nM$$

$$m = 0.1046\ x\ 127$$

$$\boldsymbol{m = 26.56\ g}$$

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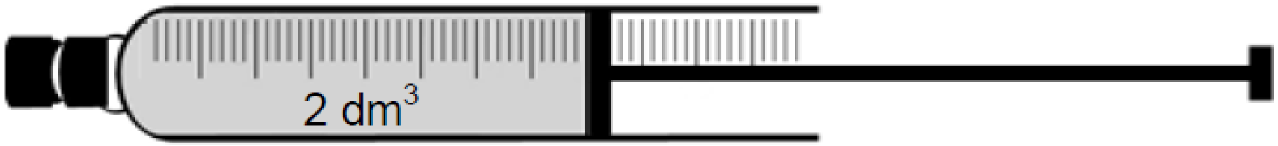
Chemical Equilibrium

Dinitrogen tetraoxide, $N_2O_4(g)$, decomposes to nitrogen dioxide, $NO_2(g)$, in a sealed syringe of volume 2 dm^3.



The mixture reaches equilibrium at 325 °C according to the following balanced equation:

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

colourless   brown

When equilibrium is reached, it is observed that the colour of the gas in the syringe is brown.

6.1 State Le Chatelier's principle. (2)

Solution:

When the equilibrium in a closed system is disturbed, the system will re-instate a (new) equilibrium ✓ by favouring the reaction that will cancel/oppose the disturbance. ✓

6.2 The syringe is now dipped into a beaker of ice water. After a while the brown colour disappears.

Is the forward reaction EXOTHERMIC or ENDOTHERMIC? Explain the answer using Le Chatelier's principle. (3)

Solution:

Endothermic.

- Decrease in temperature favours exothermic reaction
- The reverse reaction is favoured
- Number of moles/amount of NO2/brown gas decreases **or** number of moles/amount of N2O4/colourless gas increases

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

colourless   brown

6.3 The volume of the syringe is now decreased while the temperature is kept constant.

How will EACH of the following be affected? Choose from: INCREASES, DECREASES or REMAINS THE SAME.

6.3.1 The number of moles of $N_2O_4(g)$ (1)

6.3.2 The value of the equilibrium constant (1)

6.3.3 The rate of the forward and reverse reactions (1)

Solution:

6.3.1 increase

6.3.2 Remains the same

6.3.3 Increase

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

colourless   brown

6.4 Initially **X** moles of $N_2O_4(g)$ were placed in the syringe of volume 2 dm^3. When equilibrium was reached, it was found that 20% of the $N_2O_4(g)$ had decomposed.

If the equilibrium constant, K_c, for the reaction is 0,16 at 325 °C, calculate the value of **X**. (8)

Solution:

Can either use moles or concentration on the Rice table

Solution:

Using moles on the Rice table

R	$N_2O_4(g) \leftrightharpoons$	$2NO_2(g)$
I	[Y]	0
C	−x	+2x
E	[Y] − x	2x (mol)

R	$N_2O_4(g) \leftrightharpoons$	$2NO_2(g)$
I	**X**	0
C	−x	+2x
E	X − x	2x

R	$N_2O_4(g) \leftrightharpoons$	$2NO_2(g)$
I	X	0
C	−**0.2**x	+2x
E	0.8x	2x

R	$N_2O_4(g) \leftrightharpoons$	$2NO_2(g)$
I	X	0
C	−0.2x	+2x
E	0.8x	2(0.2)=**0.4**x

$$C = \frac{n}{V}$$

R	$N_2O_4(g) \leftrightharpoons$	$2NO_2(g)$
I	X	0
C	−0.2x	+2x
E	0.8x	2(0.2)=**0.4**x (mol)
E	**0.4x**	**0.2x** (mol.dm⁻³)

$$Kc = \frac{[P]}{[R]} = \frac{[NO_2]^2}{N_2O_4}$$

$$0.16 = \frac{[0.2x]^2}{[0.4x]}$$

$$0.16 = \frac{[0.4x]^2}{0.4x}$$

$$x = 0.16\ mol$$

Solution:

Using concentration on the Rice table

$$n = CV$$

$$C = \frac{n}{V}$$

$C = \frac{0.8}{2} = 0.4\ mol.dm^{-3}$ at equilibrium

$C = \frac{0.2}{2} = 0.1\ mol.dm^{-3}$ at the change

R	$N_2O_4(g) \leftrightharpoons$	$2NO_2(g)$
I	[Y]	0
C	–x	+2x
E	[Y] – x	2x ($mol.dm^{-3}$)

R	$N_2O_4(g) \leftrightharpoons$	$2NO_2(g)$
I	0.5**X**	0
C	–0.1x	+ 0.2x
E	0.5X - 0.1x	0.2x

$$Kc = \frac{[P]}{[R]} = \frac{[NO_2]^2}{N_2O_4}$$

$$0.16 = \frac{[0.2x]^2}{[0.5x - 0.1x]}$$

$$0.16 = \frac{[0.4x^2]}{0.4x}$$

$$0.16(0.4x) = 0.4x^2$$

$$0.064x = 0.04x^2$$

$$0.064 = 0.4x$$

$$x = 0.16\ mol$$

Thank you