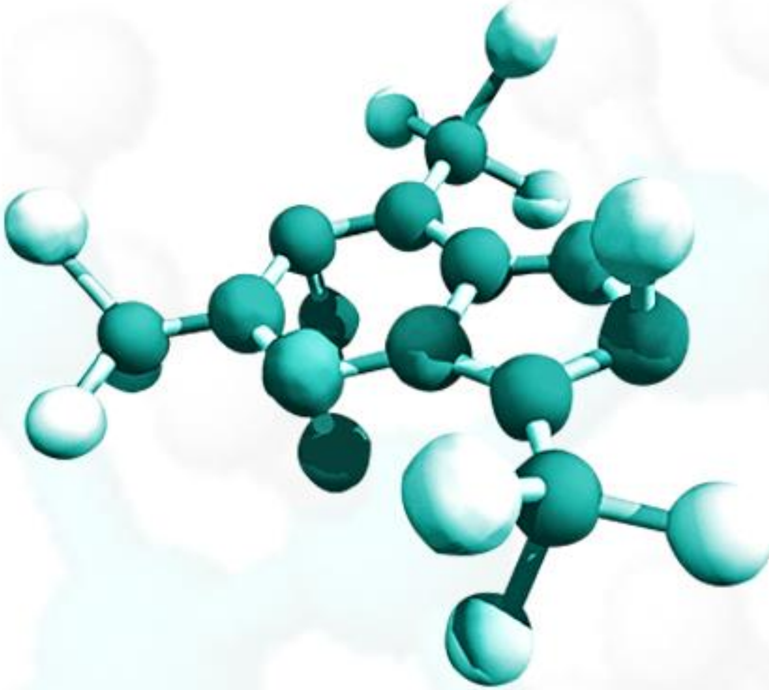
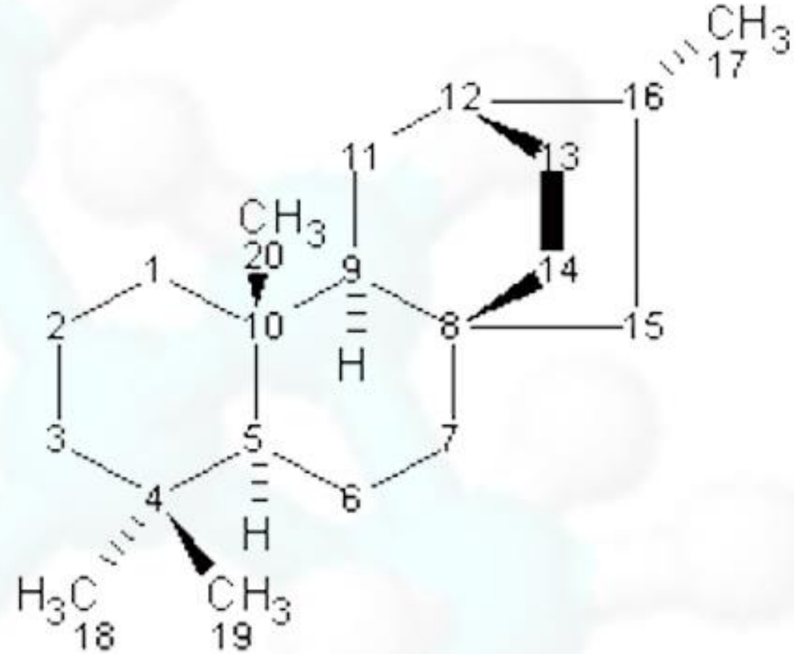




Chapter Fifteen

Acids and Bases

The Acid-Base Properties of Water

- acids is a substances that ionize in water to produce H^+ ions

$$HCl\ (aq) \rightarrow H^+(aq) + Cl^-(aq)$$

- bases is a substances that ionize in water to produce OH^- ions.

$$NaOH\ (aq) \rightarrow Na^+(aq) + OH^-(aq)$$

- An acid neutralizes a base

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(\ell)$$

The Acid-Base Properties of Water

- Water has the ability to act either as an acid or as a base.
- Water undergo ionization to a small extent this reaction is sometimes called the autoionization of water.

$$H_2O\ (l) \rightleftharpoons H^+\ (aq) + OH^-\ (aq)$$

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$$

$$K_c = [H_3O^+][OH^-] = [H^+][OH^-]$$

$$K_W = [H^+][OH^-]$$

- The ***ion-product constant*** (***K_w***) is the product of the molar concentrations of H^+ and OH^- ions **at a particular temperature.**
- **At 25 °C, K_w = 1.0 × 10^{-14}**

The Acid-Base Properties of Water

For water:

$$K_w = [H_3O^+][HO^-] = [H^+][HO^-] = 1x10^{-14}$$

Because water is neutral then

$$[H^+] = [HO^-] = \sqrt{1x10^{-14}} = 1x10^{-7}\ \mathrm{M}$$

	Solution Is
$[H^+] = [OH^-]$	neutral
$[H^+] > [OH^-]$	acidic
$[H^+] < [OH^-]$	basic

The Acid-Base Properties of Water

Example:

Calculate the $[H^+]$ ions in aqueous ammonia , $[OH^-]$ =0.0025 M?

$$K_W = [H^+][OH^-]$$

$$[H^+] = \frac{K_W}{[OH^-]}$$

$$[H^+] = \frac{1x10^{-14}}{0.0025} = 4x10^{-12} M$$

THUS $[H^+] < [OH^-]$

therefore the solution is basic

pH—A Measure of Acidity

- Because the concentrations of H^+ and OH^- ions in aqueous solutions are frequently very small numbers and therefore inconvenient to work with, Soren Sorensen in 1909 proposed a more practical measure called pH.
- The pH of a solution is defined as the negative logarithm of the hydrogen ion concentration (in mol/L).

$$pH = -\log[H^+] = -\log[H_3O^+]$$

$$[H^+] = 10^{-pH}$$

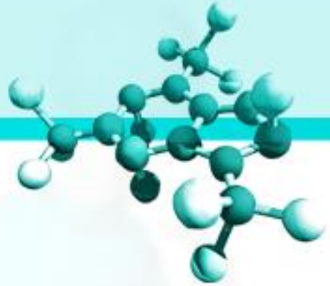
- For [OH]

$$pOH = -\log[OH^-]$$

$$[OH^-] = 10^{-pOH}$$

$$pK_w = -\log 1x10^{-14} = 14$$

$$pH + pOH = 14$$



pH—A Measure of Acidity

0 1 2 3 4 5 6 7 8 9 10 11 12 13 14

Acidic **Basic**

Increase the acidity

Increase the basisty

Solution Is		**At 25⁰C**	
neutral	$[H^+] = [OH^-]$	$[H^+] = 1 \times 10^{-7}$	pH = 7
acidic	$[H^+] > [OH^-]$	$[H^+] > 1 \times 10^{-7}$	pH < 7
basic	$[H^+] < [OH^-]$	$[H^+] < 1 \times 10^{-7}$	pH > 7

pH is inversely proportional to $[H^+]$

pH—A Measure of Acidity

Example

The concentration of H^+ ions in a bottle of vinegar was 3.2 x 10^{-4} M right after the cork was removed. Only half of the vinegar was consumed. The other half, after it had been standing open to the air for a month, was found to have a hydrogen ion concentration equal to 1.0 x 10^{-3} M. Calculate the pH of the vinegar on these two occasions.

$$pH = -\log[H^+] = -\log(3.2x10^{-4}) = 3.49$$

$$pH = -\log[H^+] = -\log(1.0x10^{-3}) = 3.00$$

pH—A Measure of Acidity

Example

The pH of rainwater collected in a certain region of Saudi Arabia on a particular day was 4.82. Calculate the H^+ ion concentration of the rainwater.

$$pH = -\log[H^+]$$

$$[H^+] = 10^{-pH}$$

$$[H^+] = 10^{-4.82} = 1.5x10^{-5} M$$

pH—A Measure of Acidity

Example

In a NaOH solution $[OH^-]$ is 2.9×10^{-4} M. Calculate the pH of the solution?

$$pOH = -\log[OH^-] = -\log(2.9x10^{-4}) = 3.54$$

$$pH + pOH = 14$$

$$pH = 14 - pOH = 14 - 3.54 = 10.46$$

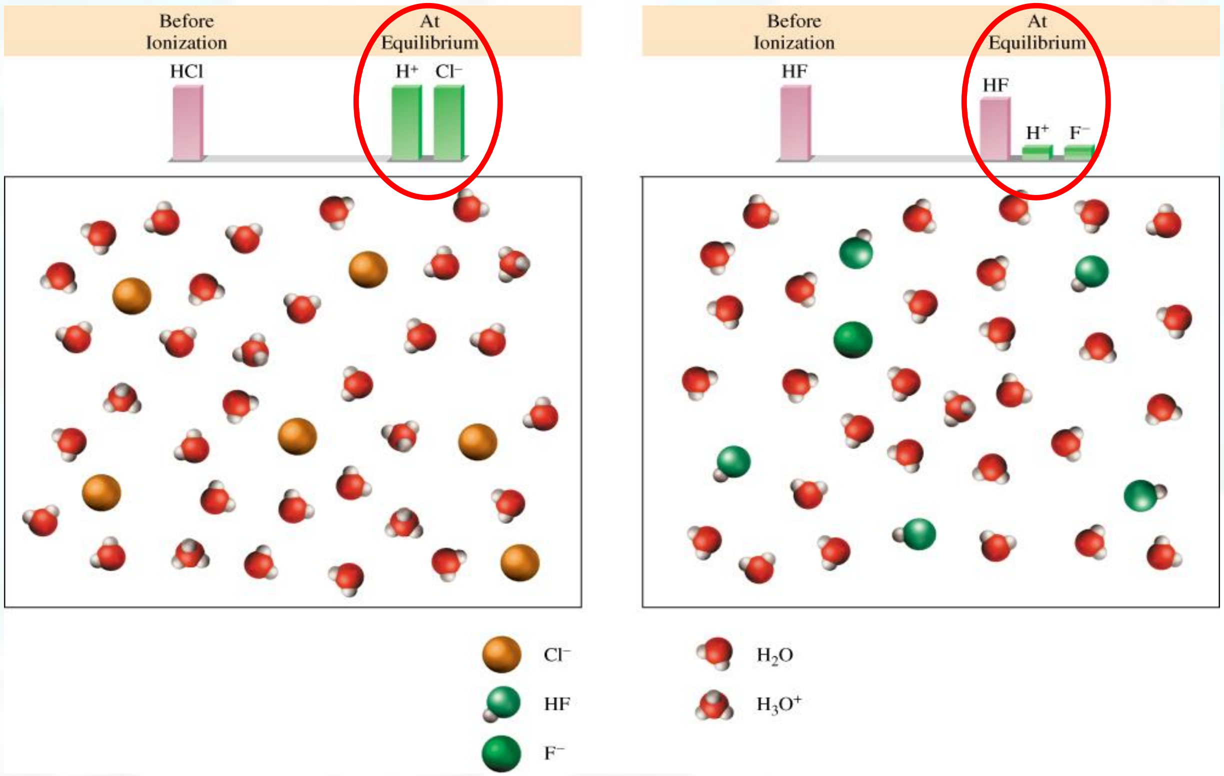
Weak Acids and Acid Ionization Constants

- Strong acid (or base) have 100 % dissociation.

$$HCl\ (s) \xrightarrow{H_2O} H^+\ (aq) + Cl^-\ (aq)$$

- Weak acid (or base) have incomplete dissociation.

$$CH_3COOH \rightleftharpoons CH_3COO^-\ (aq) + H^+\ (aq)$$



Weak Acids and Acid Ionization Constants

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

- The acid ionization constant (K_a), is the equilibrium constant for the ionization of an acid.

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

- At a given temperature, the strength of the acid HA is measured quantitatively by the magnitude of K_a. The larger K_a, the stronger the acid that is, the grater the concentration of H^+ ions at equilibrium due to its ionization.

$$[H^+] = \sqrt{K_a[acid]}$$

$$[OH^-] = \sqrt{K_b[base]}$$

Weak Acids and Acid Ionization Constants

Example

What is the pH of a 0.5 *M* HF solution (at 25^{0}C) if K_a= 7.1x10^{-4} ?

$$HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$$

$$[H^+] = \sqrt{K_a[acid]}$$

$$[H^+] = \sqrt{7.1x10^{-4}x0.5}$$

$$[H^+] = 0.019$$

pH = -log [H$^+$] = 1.72

Weak Acids and Acid Ionization Constants

Example

- What is the pH of a 0.122 *M* monoprotic acid whose K_a is 5.7 x 10^{-4}?

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

$$[H^+] = \sqrt{K_a[acid]}$$

$$[H^+] = \sqrt{5.7x10^{-4}x0.122}$$

$$[H^+] = 0.008$$

pH = -log [H^+] = 2.08

Weak Acids and Acid Ionization Constants

Example

The pH of a 0.10 M solution of formic acid (HCOOH) is 2.39. What is the K_a of the acid?

$$HCOOH\,(aq) \rightleftharpoons H^{+}\,(aq) + HCOO^{-}\,(aq)$$

$$pH = -\log[H^{+}]$$

$$[H^{+}] = 10^{-pH}$$

$$[H^{+}] = 10^{-2.39} = 4.1x10^{-3}\,M$$

$$[H^{+}] = \sqrt{K_a[acid]}$$

$$[H^{+}]^2 = K_a[acid]$$

$$K_a = \frac{[H^{+}]^2}{[acid]}$$

$$K_a = \frac{[4.1x10^{-3}]^2}{[0.1]}$$

$$K_a = 1.7x10^{-4}$$

Weak Acids and Acid Ionization Constants

- K_a indicates the strength of an acid. Another measure of the strength of an acid is percent ionization.

$$\textbf{\textit{percent ionization}} = \frac{\text{Ionized acid concentration at equilibrium}}{\text{Initial concentration of acid}} \times 100\%$$

$$\text{Percent ionization} = \frac{[H^+]}{[HA]_0} \times 100\%$$

$[HA]_0$ = initial concentration

- The stronger the acid, the greater the percent ionization.

Weak Acids and Acid Ionization Constants

Example

Calculate the percent ionization of hydrofluoric acid at the concentrations of 0.50 M if K_a= 7.1x10^{-4} ?

$$HF\ (aq) \rightleftharpoons H^+\ (aq) + F^-\ (aq)$$

$$[H^+] = \sqrt{K_a x[acid]}$$

$$[H^+] = \sqrt{7.1x10^{-4} x0.5}$$

$$[H^+] = 0.019M$$

$$\textbf{\textit{percent ionization}} = \frac{\text{Ionized acid concentration at equilibrium}}{\text{Initial concentration of acid}} \times 100\%$$

$$\textbf{\textit{percent ionization}} = \frac{0.019}{0.5} \times 100\% = 3.8\%$$

Weak Acids and Acid Ionization Constants

Example

A 0.040 M solution of a monoprotic acid is 3 percent ionized. Calculate the ionization constant of the acid.?

$$\textbf{\textit{percent ionization}} = \frac{\text{Ionized acid concentration at equilibrium}}{\text{Initial concentration of acid}} \text{ x } 100\%$$

$$3 = \frac{[H^+]}{0.04} X100$$

$$[H^+] = \frac{0.04x3}{100} = 0.0012M$$

$$[H^+] = \sqrt{K_a[acid]}$$

$$[H^+]^2 = K_a[acid]$$

$$K_a = \frac{[H^+]^2}{[acid]}$$

$$K_a = \frac{[0.0012]^2}{[0.04]}$$

$$K_a = 3.6x10^{-5}$$

Weak Bases

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

- The ionization of weak bases is treated in the same way as the ionization of weak acids.
- The base ionization constant (K_b), is the equilibrium constant for the ionization of a base.

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

- At a given temperature, the strength of the base BA is measured quantitatively by the magnitude of K_b. The larger K_b, the stronger the base—that is, the greater the concentration of OH^- ions at equilibrium due to its ionization
- In solving problems involving weak bases, we follow the same procedure we used for weak acids. The main difference is that we calculate $[OH^-]$ first, rather than $[H^+]$.

Weak Bases

- Example

What is the pH of a 0.40 M ammonia solution if K_b= 1.8x10^{-5}?

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

$$[OH^-] = \sqrt{K_b[base]}$$

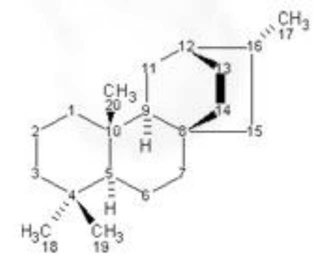
$$[OH^-] = \sqrt{1.8x10^{-5}x0.4}$$

$$[OH^-] = 0.0027$$

pOH = -log [OH$^-$] = 2.57

$$pH + pOH = 14$$

$$pH = 14 - pOH = 14 - 2.57 = 11.43$$


CH3
17
16
12
13
11
14
15
8
9
H
CH3
20
10
1
2
3
4
5
H
6
7
H3C
18
CH3
19

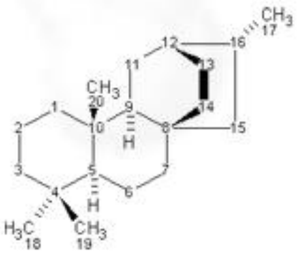
Problems of Chapter (15)

What is the concentration of H^+ in a 2.5 M HCl solution and?

HCl is a strong acid, it dissociates 100% in water according to the following formula:
$HCl + H_2O \rightarrow Cl^- + H^+$

Because HCl dissociates 100% and is a strong acid, [HCl] = [H+].

Thus, [H+] = 2.5



What is the OH^- ion concentration in a 5.2×10^{-4} M HNO_3 solution?

HNO3 dissociate completely, so the [H+]= 5.2×10^{-4}

$$K_w = [H^+]\,[OH^-]$$

$$[OH^-] = \frac{K_w}{[H^+]}$$

$$[OH^-] = \frac{1 \text{ X } 10^{-14}}{5.2 \times 10\text{-}4} = 1.92 \text{ X } 10^{-11} \text{ M}$$

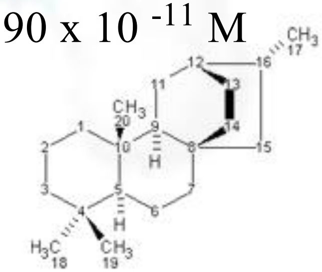
OR

$$pH = -\text{Log}\,[H^+] = -\log\,[\,5.2 \times 10^{-4}\,] = 3.2$$

$$pOH = 14 - pH = 14 - 3.2 = 10.72$$

$$[OH^-] = 10^{-pOH}$$

$$= \text{Shift log}\,(-10.72) = 1.90 \times 10^{-11} \text{ M}$$



Calculate the H^+ ion concentration in lemon juice having a pH of 2.4

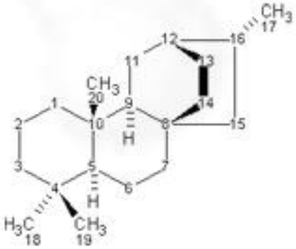
$$[\ H^+] = 10^{\ -\ pH}$$

$$= \text{Shift log}\ (\ -\ 2.4\ ) = 3.\ 98 \times 10^{\ -3}\ M$$

Calculate the pH of a 6.71×10^{-2} M NaOH solution ?.

$$pOH = -\ Log\ [OH^-] = -\ log\ [\ 6.71 \times 10^{\ -2}\ ] = 1.7$$

$$pH = 14 - pOH = 14\ -\ 1.17 = 12.82$$



What is the pH of 0.0200 M aqueous solution of HBr?

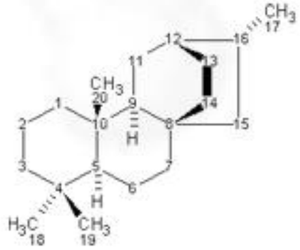
$$pH = -\text{Log}\,[H^+] = -\log\,[0.0200] = 1.69$$

The pOH of a solution of *NaOH is 11.30, what is the* $[H^+]$ *for this solution?*

$$pH = 14 - pOH = 14 - 11.30 = 2.7$$

$$[H^+] = 10^{-pH}$$

$$= \text{Shift log}\,(-2.7) = 1.99 \times 10^{-3}\ M$$



If the pH = 2 for an HNO_3 solution, what is the concentration of HNO_3?

HNO3 is a strong acid, it dissociates 100% in water according to the following formula:

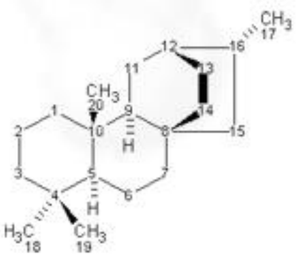
$$HNO3 \rightleftharpoons NO3^{-} + H^{+}$$

Because HNO3 dissociates 100% and is a strong acid, [HNO3] = [H+].

$$[\ H^{+}] = 10^{\ -\ pH}$$

$$= \text{Shift log}\ (-\ 2\ ) = 0.01\ M$$

$$[HNO_3] = [H^{+}] = 0.01\ M$$



A solution in which $[H^+] = 10^{-8}$ M has a Ph of ……..and is ……….

a) 8, acidic

b) 6, basic

c) -6, basic

d) 8, basic

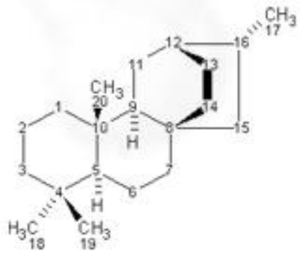
Which of the following solutions has the lowest pH at 25°C? (No calculations required.)

a) 0.2 M NaOH

b) 0.2 M NH_3

c) 0.2 M HCl

d) pure water



The equilibrium position of the reaction can be shifted in the forward direction by _____.

$$2SO_3\,(g) \rightleftharpoons 2SO_2\,(g) + O_2\,(g) \qquad \Delta H = 197\ \text{kJ}$$

a) increase pressure

b) decrease volume

c) decrease temperature

d) add SO_3

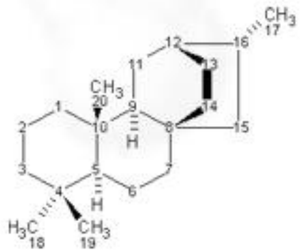
For the reaction $N_2\,(g) + 3Cl_2\,(g) \rightleftharpoons 2NCl_3\,(g)$
an analysis for the mixture at equilibrium is performed at a certain temperature. It is found that $[N_2]$=2.0 M, $[Cl_2]$=3.0 M, and $[NCl_3]$=5.0 M. Calculate K_c for the reaction at this temperature.

a) 1

b) 0.83

c) 0.46

d) 0.72



Consider the following: (1) ,

1- $2A(g) + 3B(g) \rightleftharpoons 3C(g) + D(g)$

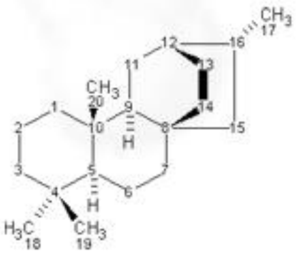
2- $2A(aq) + 3B(g) \rightleftharpoons 3C(g)$

3- $2A(aq) + 3B(g) \rightarrow 3C(s)$

4- $2A(aq) + 3B(aq) \rightarrow 3C(aq) + D(aq)$

______is an example of heterogeneous equilibrium and _____ is an example of homogenous equilibrium?

a) 2and 3 , 1 and 4



the value of K_c for the reaction

$$N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g)$$

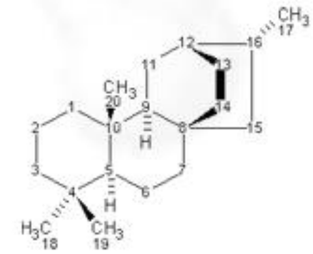
, is 1.2 . The reaction is started with $[H_2]_0 = 0.06$ M, $[N_2]_0 = 0.07$ M and $[NH_3]_0 = 0.1$ M. Which of the following is correct as the reaction comes to equilibrium?

a) The concentration of N_2 will increase

c) The concentration of NH_3 will increa

b) The concentration of H_2 will decrease

d) The reaction is at equilibrium



The equilibrium constant, K_p , for the reaction $2A_2\,(g) + 3B\,(s) \rightleftharpoons D\,(g)$, is 4.7×10^{-4} at 415°C. The value of K_p for the equilibrium

$$2D\,(g) \rightleftharpoons 4A_2\,(g) + 6B\,(s)$$

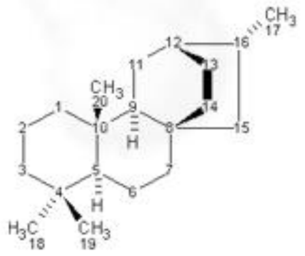
at the same temperature is _____.

a) $1.2\text{x}10^{-9}$

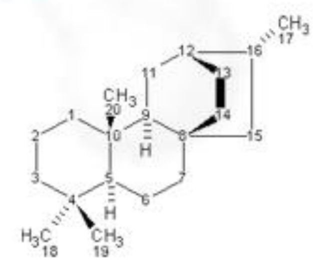
b) $4.53\text{x}10^{6}$

c) $4.7\text{x}10^{-4}$

d) $1.4\text{x}10^{10}$



For the reaction $H_2O(g) + CO(g) \leftrightarrow CO_2\ (g) + H_2(g)$, $K_c = 1.87$ at 700°C Calculate the concentration of H_2 present at equilibrium if a mixture of 0.300 moles of CO and 0.300 moles of H_2O is heated to 700°C in a 10.0 L container?



n mols of CO= 0.300 / 10 = 0.03
n mols of H2O = 0.300/ 10= 0.03

$$H_2O + CO \rightleftharpoons CO_2 + H_2$$

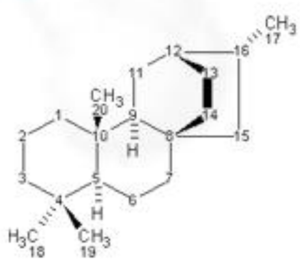
	H_2O	CO	CO_2	H_2
Initial (M)	0.03	0.03	0.0	0.0
Change	-x	-x	x	x
Equilibrium	0.03 - x	0.03 - x	x	x

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

At equilibrium $\quad Kc = \dfrac{[x][x]}{[0.03 - x][0.03 - x]}$

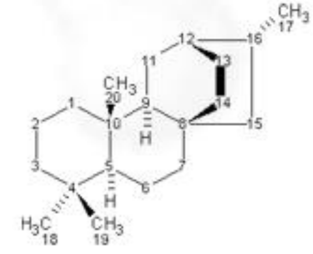
$$1.87 = \frac{[x]^2}{[0.03 - x]^2}$$

$$1.36 = \frac{[x]}{[0.03 - x]}$$



x = 0.017

[H2] = x= 0.017 M


CH3
H3C
CH3
CH3
H
H