## EXERCISE

**Q1 Select the most suitable answer from the given ones in each question.**

1. Isotopes difer in
   1. properties which depend upon mass
   2. arrangement of electrons in orbitals
   3. chemical properties
   4. the extent to which they may be afected in electromagnetic ield.
2. Select the most suitable answer from the given ones in each question.
   1. Isotopes with even atomic masses are comparatively abundant.
   2. Isotopes with odd atomic masses are comparatively abundant.
   3. Isotopes with even atomic masses and even atomic numbers are comparatively abundant.(d) Isotopes with even atomic masses and odd atomic numbers are comparatively abundant.
3. Many elements have fractional atomic masses. This is because (a) the mass of the atom is itself fractional.
   1. atomic masses are average masses of isobars.
   2. atomic masses are average masses of isotopes.
   3. atomic masses are average masses of isotopes proportional to their relative abundance. (iv) The mass of one mole of electrons is

(a) 1.008 mg (b) 0.55 mg (c) 0.184 mg (d)1.673mg

(v) 27 g of A1 will react completely with how much mass of O2 to produce Al2O3. (a) 8 g of oxygen (b) 16 g of oxygen (c) 32 g of oxygen (d) 24 g of oxygen (vi) The number of moles of CO2 which contain 8.0 g of oxygen.

(a) 0.25 (b) 0.50 (c) 1.0 (d)1.50

1. The largest number of molecules are present in

(a) 3.6g of H2O (b) 4.8g of C2H5OH (c) 2.8g of CO (d) 5.4g of N2O5

1. One mole of SO2 contains

(a) 6.02x1023 atoms of oxygen (b) 18.1 x 1023 molecules of SO2

(c) 6.02x1023 atoms of sulphur (d) 4 gram atoms of SO2

1. The volume occupied by 1.4 g of N2 at S.T.P is

(a) 2.24 dm3 (b) 22.4 dm3 (c) 1.12 dm3 (d) 112 cm3

1. A limiting reactant is the one which
   1. is taken in lesser quantity in grams as compared to other reactants.
   2. is taken in lesser quantity in volume as compared to the other reactants.
   3. gives the maximum amount of the product which is required.
   4. gives the minimum amount of the product under consideration.

Q 2. Fill in the blanks

1. The unit of relative atomic mass is\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.
2. The exact masses of isotopes can be determined by\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ spectrograph.
3. The phenomenon of isotopy was irst discovered by .\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
4. Empirical formula can be determined by combustion analysis for those compounds which have \_\_\_\_\_\_ and \_\_\_\_\_\_\_\_\_\_in them.
5. A limiting reagent is that which controls the quantities of\_\_\_\_\_\_\_\_\_\_\_\_
6. 1 mole of glucose has\_\_\_\_\_\_\_\_\_ atoms of carbon, \_\_\_\_\_\_\_\_\_\_ of oxygen and \_\_\_\_\_\_\_of hydrgen.
7. 4g of CH4 at 0°C and 1 atm pressure has \_\_\_\_\_\_\_\_\_\_\_\_\_molecules of CH4
8. Stoichiometric calculations can be performed only when \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ is obeyed.

Q3. Indicate true or false as the case may be:

* + 1. Neon has three isotopes and the fourth one with atomic mass 20.18 amu.
    2. Empirical formula gives the information about the total number of atoms present in the molecule.
    3. During combustion analysis Mg(ClO4)2 is employed to absorb water vapours. (iv) Molecular formula is the integral multiple of empirical formula and the integral mutiple can never be unity.
    4. The number of atoms in 1.79 g of gold and 0.023 g of sodium are equal.
    5. The number of electrons in the molecules of CO and N2 are 14 each, so 1 g of each gas will have same number of electrons.
    6. Avogadro’s hypothesis is applicable to all types of gases i.e. ideal and non-ideal. (viii) Actual yield of a chemical reaction may be greater than the theoretical yield. Q.4 What are ions? Under what conditions are they produced?

Q.5 (a) What are isotopes? How do you deduce the fractional atomic masses of elements from the relative isotopic abundance? Give two examples in support of your answer.

(b) How does a mass spectrograph show the relative abundance of isotopes of an elment? (c) What is the justiication of two strong peaks in the mass spectrum for bromine; while for iodine only one peak at 127 amu is indicated?

* 1. Silver has atomic number 47 and has 16 known isotopes but two occur naturally i.e. Ag-107 and Ag-109. Given the following mass spectrometric data, calculate the average atomic mass of silver.

|  |  |  |
| --- | --- | --- |
| **Isotopes** | **Mass (amu)** | **Percentage abundance** |
| 107Ag  109Ag | 106.90509  108.90476 | 51.84  48.16 |

* 1. Boron with atomic number 5 has two naturally occurring isotopes. Calculate the percentage abun dance of 10B and 11B from the following informations.

Average atomic mass of boron = 10.81 amu

Isotopic mass of 10B = 10.0129 amu

Isotopic mass of 11B =11.0093amu (Ans: 20.002%, 79.992) Q.8 Deine the following terms and give three examples of each.

|  |  |
| --- | --- |
| 1. Gram atom 2. Gram molecular **mass** 3. Gram formula 4. Gram ion | 1. Molar volume 2. Avogadro’s number 3. Stoichiometry 4. Percentage yield |

Q.9 Justify the following statement!:

* + 1. 23 g of sodium and 238 g of uranium have equal number of atoms in them.
    2. Mg atom is twice heavier than that of carbon atom.
    3. 180 g of glucose and 342 g of sucrose have the same number of molecules but diferent number of atoms present in them.
    4. 4.9 g of H2SO4 when completely ionized in water, have equal number of positive and negative charges but the number of positively charged ions are twice the number of negatively charged ions.
    5. One mg of K2CrO4 has thrice the number of ions than the number of formula units when ionized in water.
    6. Two grams of H2,16g of CH4 and 44 g of CO2 occupy separately the volumes of 22.414 dm3, although the sizes and masses of molecules of three gases are very diferent from each other.
  1. Calculate each of the following quantities.

|  |  |  |
| --- | --- | --- |
| 1. Mass in grams of 2.74 moles of KMnO4. 2. Moles of O atoms in 9.00g of Mg (N03)2. 3. Number of O atoms in 10.037 g of CUSO4.5H2O. 4. Mass in kilograms of 2.6 x 1020 molecules of SO2. 5. Moles of Cl atoms in 0.822 g C2H4Cl2. 6. Mass in grams of 5.136 moles of Ag2CO3. 7. Mass in grams of 2.78 x 1021 molecules of CrO2Cl2. 8. Number of moles and formula units in 100g of KClO3. 9. Number of K+ ions, CIO-3 ions, Cl atoms, and O atoms in (h). | * •   •  •   * • •   •  • | (Ans: 432.92g)  (Ans: 0.36 mole)  (Ans: 2.18 x 1023 atoms)  (Ans: 2.70x10-5 kg)  (Ans: 0.0178 moles)  (Ans: 1416.2 g)  (Ans: 0.7158 g)  (Ans: 0.816 moles, 4.91 x 1023 formula units)  (Ans: 4.91 x 1023 K+, 4.91x 1023 CIO3-1, 4.91x 1023  Cl-1,1.47x 1024 O atoms) |

* 1. Aspartame, the artiicial sweetner, has a molecular formula of C14H18N2O5.

1. What is the mass of one mole of aspartame? (Ans: 294 g mol-1)
2. How many moles are present in 52 g of aspartame? (Ans: 0.177 mole)
3. What is the mass in grams of 10.122 moles of aspartame? (Ans: 2975.87
4. How many hydrogen atoms are present in 2.43 g of aspartame?(Ans: 8.96 x 1022 atoms of H)

Q.12 A sample of 0.600 moles of a metal M reacts completely with excess of luorine to form 46.8 g of MF2.

1. How many moles of F are present in the sample of MF2 that forms? (Ans: 1.2 moles)
2. Which element is represented by the symbol M? (Ans: calcium)

Q.13 In each pair, choose the larger of the indicated quantity, or state if the samples are equal.

1. Individual particles: 0.4 mole of oxygen molecules or 0.4 mole of oxygen atoms.

(Ans: both are equal)

1. Mass: 0.4 mole of ozone molecules or 0.4 mole of oxygen atoms. (Ans: ozone)
2. Mass: 0.6 mole of C2H4 or 0.6 mole of I2. (Ans: I2)
3. Individual particles: 4.0 g N2O4 or 3.3 g SO2. (Ans: SO2)
4. Total ions: 2.3 moles of NaCIO3 or 2.0 moles of MgCl2. (Ans: MgCl2)
5. Molecules: 11.0 g H2O or 11.0 g H2O2. (Ans:H2O)
6. Na+ ion: 0.500 moles of NaBr or 0.0145 kg of NaCl. (Ans: NaBr)
7. Mass: 6.02 x 1023 atoms of 235U or 6.02 x 1023 atoms of 238U. (Ans: U238) Q.14 a) Calculate the percentage of nitrogen in the four important fertilizers i.e.,

(i) NH3 (ii) NH2CONH2(urea) (iii) (NH4)2SO4 (iv) NH4NO3.

(Ans: 82.35%, 46.67%, 21.21%, 35%)

b) Calculate the percentage of nitrogen and phosphorus in each of the following:

(i) NH4H2PO4 (ii) (NH4)2HPO4 (iii) (NH4)3PO4

(Ans: (i)N=12.17%,P=26.96% (ii)N=21.21%,P=23.48% (iii)N=28.18%,P=20.81%) Q.15 Glucose C6H12 O6 is the most important nutrient in the cell for generating chemical potential energy. Calculate the mass % of each element in glucose and determine the number of C, H and O atoms in 10.5 g of the sample.

(Ans: C=40%, H=6.66%, 0 =53.33%, C=2.107x1023, H=4.214x1023, O=2.107x 1023)Q.16 Ethylene glycol is used as automobile antifreeze. It has 38.7% carbon, 9.7 % hydrogen and 51.6% oxygen. Its molar mass is 62.1 grams mol-1. Determine its empirical formula.?

(Ans: CH3O)

Q.17 Serotenin (Molar mass = 176g mol-1) is a compound that conducts nerve impulses in brain and muscles. It contains 68.2 % C.6.86 % H, 15.09 % N, and 9.08 % O. What is its molecular formula.

(Ans: C10H12N2O) Q.18 An unknown metal M reacts with S to form a compound with a formula M2S3. If 3.12 g of M reacts with exactly 2.88 g of sulphur, what are the names of metal M and the-compowad M2S3?

(Ans: Cr; Cr2S3)

Q.19 The octane present in gasoline burns according to the following equation.

2C H8 18 (l) + 25O (g) 2 → 16CO (g) + 18H O(l)2 2

1. How many moles of O2 are needed to react fully with 4 moles of octane?

(Ans: 50 moles)

1. How many moles of CO2 can be produced from one mole of octane?

(Ans: 8 moles)

1. How many moles of water are produced by the combustion of 6 moles of octane?

(Ans: 54 moles)

1. If this reaction is to be used to synthesize 8 moles of CO2 how many grams of oxygen are needed? How many grams of octane will be used?

(Ans: 400 g: 114 g)

Q.20 Calculate the number of grams of Al2S3 which can be prepared by the reaction of 20 g of Al and 30 g of sulphur. How much the non-limiting reactant is in excess?

(Ans: 46.87g; 3.125g) Q.21 A mixture of two liquids, hydrazine N2H4 and N2O4 are used in rockets. They produce N2 and

water vpours. How many grams of N2 gas will be formed by reacting 100 g of N2H4 and 200g of

N2O4. (Ans: 131.04g)

2N H + N O 2 4 2 4 → 3N + 4H O2 2

* 1. Silicon carbide (SiC) is an important ceramic material. It is produced by allowing sand (SiO2 ) to react with carbon at high temperature.

SiO + 3C 2 → SiC + 2CO

When 100 kg sand is reacted with excess of carbon, 51.4 kg of SiC is produced. What is the

pecentage yield of SiC? (Ans: 77%)

* 1. a. What is stoichiometry? Give its assumptions? Mention two important laws, which help to perform the stoichiometric calculations?

b. What is a limiting reactant? How does it control the quantity of the product formed? Explain with three examples?

Q.24 a. Deine yield. How do we calculate the percentage yield of a chemical reaction?

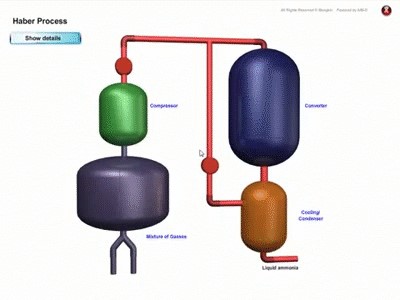
b. What are the factors which are mostly responsible for the low yield of the products in chemical reactions?

Q.25 Explain the following with reasons.

* + 1. Law of conservation of mass has to be obeyed during stoichiometric calculations.
    2. Many chemical reactions taking place in our surrounding involve the limiting reactants.iii) No individual neon atom in the sample of the element has a mass of 20.18 amu. iv) One mole of H2SO4 should completely react with two moles of NaOH. How does Avogadro’s number help to explain it.
    3. One mole of H2O has two moles of bonds, three moles of atoms, ten moles of electrons and twenty eight moles of the total fundamental particles present in it.
    4. N2 and CO have the same number of electrons, protons and neutrons.

4.85A pm)

## CHAPTER 8 CHEMICAL EQUILIBRIUM



Animation 8.1: Haber’s Process

Source & Credit: makeagif

### 8.1.0 REVERSIBLE AND IRREVERSIBLE REACTIONS

A chemical reaction can take place in both directions, i.e. forward and reverse, but in some cases the tendency of reverse reaction is very small and is negligible. For example, sodium reacts with water to form sodium hydroxide and hydrogen gas.

2Na(s)+2H O( ) 2 → 2NaOH(aq)+H (g)2

The tendency for hydrogen to react with sodium hydroxide to form sodium and water is negligible at normal temperature. This is an example of irreversible reaction.

Let us take another example of the reaction between two parts of hydrogen and one part of oxygen by means of an electric spark at normal temperature and pressure. The reaction occurs stoichiometrically according to the following chemical equation.

2H (g) + O (g) 2 2 → 2H O( )2 

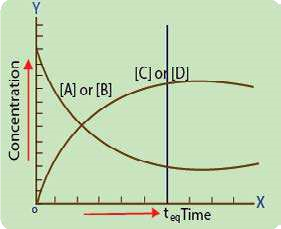
If hydrogen and oxygen are present in correct proportion, there will be no residual gases i.e. hydrogen and oxygen. If the product is heated to a temperature of 15000C, a noticeable quantity of H2O decomposes, producing hydrogen and oxygen. It means that reverse reaction does occur, but only at higher temperature. It is very likely that the reverse reaction occurs at low temperature, but it is too small to be noticeable. The reaction between stoichiometric amounts of hydrogen and oxygen proceeds to completion in the presence of electric spark. Such reactions are called irreversible reactions and they take place in one direction only.

Now, consider a reaction between nitrogen and hydrogen at 4500C under high pressure in the presence of iron as a catalyst.

N (g) + 3H (g) 2 2 ฀฀฀฀฀฀฀฀฀฀฀฀฀฀high pressureFe/450 Co  2HN (g)3

There action mixture, after some time, will contain all the three species i.e. nitrogen, hydrogen and ammonia. No matter, how long the reaction is allowed to continue, the percentage composition of species present remains constant. The conditions are favourable for the forward as well as for a reverse reaction to occur to a measurable extent. This type of reaction is described as a reversible reaction.

**8.1.1 State of Chemical Equilibrium**

 If a reversible reaction is allowed to continue for a considerable long time, without changing the conditions, there is no further change in composition of the reaction mixture. The reaction is said to have attained a state of chemical equilibrium. Once this equilibrium has been established, it will last forever if undisturbed.

To illustrate an example of the attainment of equilibrium, let us consider a general reaction in which A reacts with B to produce C and D.

A(g) + B(g) ฀ C(g) + D(g)

Suppose that all the substances are in gaseous state.

Let the initial concentrations of A and B be equal .

As time goes on, concentrations of A and B decrease, at irst quite rapidly but later slowly. Eventually, the concentrations of A and B level of and become

constant. Fig (8.1) Reversible reaction and state of equilibrium The graph is plotted between time and

concentrations for reactants and products, Fig(8.1). The initial concentrations of C and D are zero. As the time passes the products C and D are formed. Their concentrations increase rapidly at irst and then level of. At the time of equilibrium concentrations become constant. This is how the chemical equilibrium is attained and state of equilibrium is reached.

Now, let us consider the example of a reversible reaction between hydrogen gas and iodine vapours to form hydrogen iodide at 4250 C. At equilibrium three components will be present in deinite proportions in the reaction m ixture Fig (8.2). The equilibrium is established when the rising curve of product HI and the falling curve of reactants [H2] and [I2] become parallel to time axis.

H (g) + l (g)2 2 ฀฀฀฀฀฀฀฀฀฀4250C 2Hl(g)

The same equilibrium mixture is obtained irrespective whether the reaction starts by mixing hyd rogen and iodine or by decomposition of hydrogen iodide. The situation suggests two possibilities of the state of reaction at equilibrium

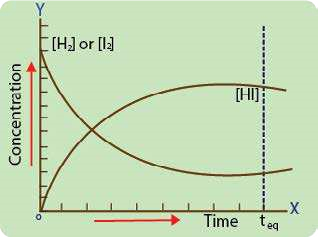
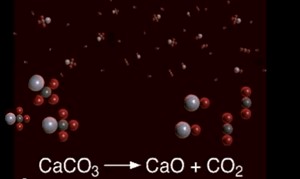


Fig (8.2) State of dynamic equilibrium

1. All reactions cease at equilibrium so that the system becomes stationary.
2. The forward and reverse reactions are taking place simultaneously at exactly the same rate. It is now universally accepted that the later conditions prevail in a reversible reaction at equilibrium stage of reaction. It is known as the state of dynamic equilibrium.



*Animation 8.2: Chemical Equilibrium*

*Source & Credit :*

[*dynamicscienc*](http://www.dynamicscience.com.au/tester/solutions1/chemistry/equilibrium/equilibrium2.htm)

[*e*](http://www.dynamicscience.com.au/tester/solutions1/chemistry/equilibrium/equilibrium2.htm)

**8.1.2 Law of Mass Action**

A state of dynamic equilibrium helps to determine the composition of reacting substances and the products at equilibrium. We use the relationship which was derived by C.M. Guldberg and R Waage in 1864. **It is known as the law of mass action. It states that the rate at which the reaction proceeds is directly proportional to the product of the active masses of the reactants.**

The term active mass represents the concentration in mole dm-3 of the reactants and products for a dilute solution.

Now, consider a general reaction in which A and B are the reactants and C and D are the products. The reaction is represented by the following chemical equation.

A + B ฀฀฀฀฀฀฀฀kkf*r*  C + D

The equilibrium concentrations of A, B, C and D are represented in square brackets like [A], [B], [C] and [D] respectively and they are expressed in moles dm-3. According to the law of mass action, the rate of the forward reaction, is proportional to the product of molar concentrations of A and B.

Rate of forward reaction (R ) f ∝ [A][B]

or R = k [A] [B]f f

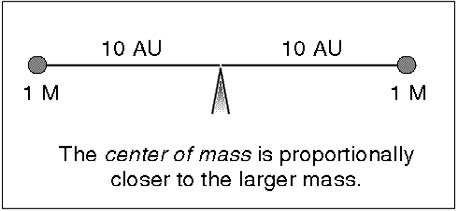
kf is the proportionality constant and is called rate constant for forward reaction and Rf is the rate of forward reaction. Similarly, the rate of reverse reaction (Rr) is given by

Rate of reverse reaction (R ) r ∝ [C][D]

R = k [C] [D]r r

Where kr is the proportionality constant and is called the rate constant for backward reaction.

Remember that C and D are the reactants for backward step.



*Animation 8.3: Mass Action*

*Source & Credit :*

[*astronomynote*](http://www.astronomynotes.com/starprop/s10.htm)

[*s*](http://www.astronomynotes.com/starprop/s10.htm)

At equilibrium,

R = Rf  r

or k [A] [B] = k [C] [D]f r

On rearranging, we get

kf [C][D] =

kr [A][B]

Let kf = Kc

k r

So, Kc = [C][D]

[A][B]

The constant Kc is called the equilibrium constant of the reaction. Kc is the ratio of two rate

constants.

Conventionally, while writing equilibrium constant, the products are written as numerator and reactants as denominator.

K = c [products] or Kc = rate constant for forward step

[reactants] rate constant for reverse step

For a more general reaction aA + bB ฀฀฀฀฀฀฀฀kkfr cC + dD

Where a, b, c and d are the coeicients of balanced chemical equation. They are number of moles of A, B, C and D, respectively in the balanced equation. The equilibrium constant is given by

[C] [Dc ]d

K = c a  [A] [B]b

Hence, the coeicients in the equation appear as exponents of the terms of concentrations in the equilibrium constant expression.

**Units of Equilibrium Constants**

Equilibrium constant is the ratio of the products of the concentrations of the products to the product of concentrations of the reactants. If the reaction has equal number of moles on the reactant and product sides, then equilibrium constant has no units. When the number of moles is unequal then it has units related to the concentration or pressure. But it is a usual practice that we don’t write the units with Kp or Kc values.

Following are some important reversible reactions. Their units of Kc are expressed as

1. CH COOH(aq) + C H OH(aq) 3 2 5 ฀ CH COOC H (aq) + H O(3 2 5 2 )

[CH COOC H ][H O3 2 5 2 ] [moles dm ][moles dm-3 -3]

Kc === -3 -3 no units

[CH COOH][C H OH3 2 5 ] [moles dm ][moles dm ]

1. N (g) + 3H (g) 2 2 ฀ 2NH (g3 )

[NH3]2 [moles dm-3 2] -2

Kc == 3 = -3 -3 3 moles dm+6

[N ][H2 2 ] [moles dm ][moles dm ]

In the expression of Kc, we have ignored the physical states for the sake of convenience.

**Example 1:**

The following reaction was allowed to reach the state of equilibrium.

2A(aq) + B(aq) ฀ C(aq)

The initial amounts of the reactants present in one dm3 of solution were 0.50 mole of A and 0.60 mole of B. At equilibrium, the amounts were 0.20 moles of A and 0.45 mole of B and 0.15 mole of C. Calculate the equilibrium constant Kc.

**Solution**

Equation: 2A(aq) + B(aq) ฀฀฀฀฀฀฀฀kkfr C(aq)

Kc for the reaction is given by

[C]

K = c 2

[A] [B]

2A(aq) + B(aq) ฀C(aq)

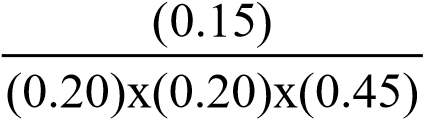
Initial concentrations 0.50 mol 0.60 mol 0.00 mol

Equilibrium concentrations 0.20 mol 0.45 mol 0.15 mol

Since K = c [C2 ]

[A] [B]

Putting values of concentrations, which are present at equilibrium stage

So, K = c 

K = c 1 == 1 8.3 Answer

0.20x0.20x3 0.12

The units have been ignored for the sake of convenience.

**8.1.3 Equilibrium Constant Expressions for Some Important Reactions**

1. **Formation of Ester from an Organic Acid and Alcohol (aqueous phase reaction)**  This is a well known reversible reaction in the solution state.

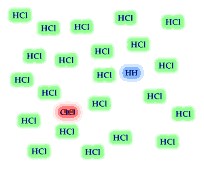
CH COOH(aq) + C H OH(aq) ฀฀฀฀฀฀฀฀฀฀H O3 + CH COOC H (aq) + H O()

3 3 2 3 2 5 2 acid alcohol ester water

Let us suppose that ‘a’ moles of CH3COOH and ‘b’ moles of C2H5OH are initially taken in a vessel in the presence of small amount of a mineral acid as a catalyst.

The progress of the reaction can be studied by inding out the concentrations of acetic acid after regular intervals.

A very small portion of the reaction mixture is withdrawn and the concentration of acetic acid is determined by titrating it against a standard solution of sodium hydroxide. The concentration of acetic acid will decrease until the attainment of state of equilibrium, when it will become constant. At equilibrium stage, x moles of ester and ‘x’ moles of H2O are produced. The number of moles of acid and alcohol left behind are ‘a-x’ moles and ‘b-x’ moles respectively. If the volume of reaction mixture at equilibrium stage is ‘V’ dm3, then



*Animation 8.4:*

[*Equilibrium Constan*](http://www.800mainstreet.com/7/0007-000-ans-selftest.html)

[*t*](http://www.800mainstreet.com/7/0007-000-ans-selftest.html)

*Source & Credit :*

[*mainstree*](http://www.800mainstreet.com/7/0007-000-ans-selftest.html)

[*t*](http://www.800mainstreet.com/7/0007-000-ans-selftest.html)

CH COOH(aq) + C H OH(aq) 3 2 5 ฀ CH COOC H (aq) + H O(3 2 5 2 )

'a' moles 'b' moles ฀ '0' moles '0' moles t=0sec

(a-x) moles (b-x) moles ฀ 'x' moles 'x' moles t=teq

When number of moles are divided by total volume of the reaction mixture, we get concentration of each species at equilibrium stage in moles dm-3.

 a-xV moles dm +-3  b-xV moles dm-3 ฀   Vx moles dm-3 +   Vx moles dm-3

Kc = [CH COOC H ][H O3 2 5 2 ]

Since [CH COOH][C H OH3 2 5 ]

Brackets [ ] denote the concentrations in moles dm-3.

Putting concentrations at equilibrium

X X

.

K = c V V

(a-x) (b-x)

.

V V

Simplifying the right hand side, we get

X2

K =c

(a-x)(b-x)

In this expression of Kc, the factor of volume is cancelled out. So, the change of volume at equilibrium stage does not afect the Kc value or equilibrium position of reaction.

1. **Dissociation of PCl5 (gaseous phase reaction)**

The dissociation of PCl5 into PCI3 and Cl2, is a well known homogeneous gaseous phase reaction. This reaction has unequal number of moles of reactants and products.

PCl (g)5 ฀ PCl (g) + Cl (g)3 2

Let ‘a’ moles of PCl5 present initially are decomposed by ‘x’ moles. So, at equilibrium stage, ‘ax’ moles of PCl5 are left behind while ‘x‘ moles of PCI3 and ‘x’ moles of Cl2 are produced. If the volume of equilibrium mixture is ‘V’ dm3, then

PCl (g)5 ฀ PCl (g) 3 + Cl (g)2

‘a’moles ‘O’moles ‘O’moles t = 0sec

(a - x) moles ‘x‘ moles ‘x‘ moles t = teq

Dividing the number of moles by total volume of reactants and products at equilibrium.

 a-x moles dm-3 ฀    Vx moles dm-3 +   Vx moles dm-3

 V 

Since K = c [PCl ][Cl3 2 ]

PCl5

Putting the concentrations at equilibrium

X X

.

K = c V V

(a-x)

V

Simplifying the right hand side, we get

x2

K =c

V(a-x)

The inal expression is not independent of the factor of volume. So, the change of volume at equilibrium stage disturbs the equilibrium position of the reaction. We will discuss this reaction in Le-Chatelier’s principle with reference to efect of volume change and its efect on change of equilibrium position.

1. **Decomposition of N2O4 (gaseous phase reaction)**

Similarly, for decomposition of N2O4 (g). the expression of Kc involves the factor of volume.

N O (g) 2 4 ฀ 2NO (g2 )

4x2 K =c

(a-x)V

‘a’ is the initial number of moles of N2O4 ‘x’ is number of moles of N2O4 decomposed and ‘V’ is total volume of N2O4 , and NO2 at equilibrium stage. **iv. Synthesis of NH3 ( gaseous phase reaction)**

For the synthesis of ammonia,

N (g) + 3H (g) 2 2 ฀2NH (g3 ) the expression of Kc is

4x V2 2 K =c

(a-x)(b-3x)3

Where ’a’ and ’b’ are the initial number of moles of N2 and H2 and ’x’ is number of moles of N2, decomposed at equilibrium stage. ‘V’ is the total volume of N2, H2 and NH3 at equilibrium. The inal expression involves V2 in the numerator:

Hence, it depends upon the coeicients of balanced equation that whether the factor of volume will appear in numerator or denomenator.

**8.1.4 Relationship Between Equilibrium Constants**

The expressions of equilibrium constants depend upon the concentration units used. Mostly the concentrations are expressed in mole dm-3. Let us consider the following reversible reaction.

aA + bB ฀ cC + dD

[C] [Dc ]d C DCc Dd

K =c a b or K = c a b

[A] [B] AA BB

The square brackets represent the concentration of species in moles dm-3. Anyhow, the capital C is also used for molar concentrations.

If the reactants A, B, and the products C, D of the reaction under consideration are ideal gases, then molar concentration of each gas is proportional to its partial pressure . When the concentrations are expressed in terms of partial pressures, the expression of Kp is,

p pCc d

K = p a D

pA p b

B

Here PA, PB, PC and PD are partial pressures of A, B, C, D respectively at equilibrium position. As long as the number of moles of products and reactants, which are in the gaseous state, are equal, the values of KC and KP remain the same. Otherwise, the following relationship between KP and KC can be derived by using Dalton’s law of partial pressures.

∆n K =K (RT)p c

Where ‘∆n’ is the diference between number of moles of the gaseous products and the number of moles of gaseous reactants.

∆n = no. of moles of products - no. of moles of reactants

‘R’ is the general gas constant and ‘T’ is absolute temperature at which the reaction is being carried out

Where, ∆n = 0, then all the equilibrium constants have the same values.

**Example 2:**

N2 (g) and H2 (g) combine to give NH3 (g). The value of Kc in this reaction at 500 °C is 6.0 x 10-2. Calculate the value of Kp for this reaction.

**Solution:**

The reaction for the synthesis of NH3 is

N (g) + 3H (g) 2 2 ฀ 2NH (g3 )

This reaction takes place with decrease in the number of moles. The relationship of Kpand Kc is

∆n K =K (RT)p c

Now Kp = 6.0 x 10-2

Temperature = 500 + 273=773 K

∆n = no of moles of products - no of moles of reactants

∆n = 2 - 4 = -2

R = 0.0821 dm3 atm K-1 mol-1

Substituting these values in the expression

Kp = 6.0 x 10-2 (773 x 0.0821)-2 = 6.0 x 10-2 (63.5)-2

6.0x10-2

K =p

(63.5)2

|  |
| --- |
| -5 |

K =1.5x10pAnswer

In this case the value of Kp is smaller than Kc . Those reactions, which take place with the increase in the number of moles mostly have greater Kp than Kc.

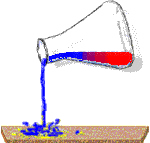
### 8.1.5 APPLICATIONS OF EQUILIBRIUM CONSTANT

The value of equilibrium constant is speciic and remains constant at a particular temperature.

The study of equilibrium constant provides us the following informations:

(i) Direction of reaction (ii) Extent of reaction

(iii) Efect of various factors on equilibrium constant and equilibrium position.



*Animation 8.5: APPLICATIONS OF*

*EQUILIBRIUM CONSTANT*

*Source & Credit :*

[*oocitie*](http://www.oocities.org/rjwarren_stm/OAC_Part_Two.html)

[*s*](http://www.oocities.org/rjwarren_stm/OAC_Part_Two.html)

**(i) Direction of Reaction**

we know that, K = c [Products] for any reaction.

[Reactants]

The direction of a chemical reaction at any particular time can be predicted by means of [products] / [reactants] ratio, calculated before the reaction attains equilibrium. The value of [product] / [reactants] ratio leads to one of the following three possibilities.

1. The ratio is less than Kc. This implies that more of the product is required to attain the equilibrium, therefore, the reaction will proceed in the forward direction.
2. The ratio is greater than Kc. It means that the reverse reaction will occur to attain the equilibrium. (c) When the ratio is equal to Kc, then the reaction is at equilibrium.

**Example 3:**

Esteriication reaction between ethanol and acetic acid was carried out by mixing deinite amounts of ethanol and acetic acid alongwith some mineral acid as a catalyst. Samples were drawn out of the reaction mixturq to check the progress of the esteriication reaction. In one of the samples drawn after time t, the concentrations of the species were found to be [CH3COOH] = 0.025 mol dm-3, [C2H5OH]= 0.032 mol dm-3, [CH3COOC2H5] = 0.05 mol dm-3, and [H20] = 0.04 mol dm-3. Find out the direction of the reaction if Kc for the reaction at 25°C is 4.

**Solution:**

Esteriication reaction is represented by the following stoichiometric equation. 3 2 5 ฀฀฀฀฀฀฀฀H+ CH COOC H + H O3 2 5 2

CH COOH + C H OH

All the substances are present in the same volume of solution, therefore KC is given by

Kc = [CH COOC H ][H O3 2 5 2 ]

[CH COOH][C H OH3 2 5 ]

The various values of concentrations, at time t are substituted to get the ratio

0.05x0.04

K =c  = 2.50 Answer

0.025x0.032

The given value of KC for this reaction is 4 and 2.5 is less than KC. Therefore, the reaction will proceed in the forward direction to attain the equilibrium.

**(ii) Extent of Reaction**

1. If the equilibrium constant is very large, this indicates that the reaction is almost complete.
2. If the value of Kc is small, it relects that the reaction does not proceed appreciably in the forward direction.
3. If the value of Kc is very small, this shows a very little forward reaction.

**Examples:**

Equilibrium constant for the decomposition of ozone to oxygen is 1055 at 25°C.

i.e., 2O3 ฀3O , K =10 at 25 C2 c 55 o

It infers that at room temperature 03 is unstable and decomposes very rapidly to 02. This reaction is almost complete.

On the other hand the value of equilibrium constant for the decomposition of HF at 2000°C is 10-13.

2HF(g) ฀H (g) + F (g) K =10-13 at 2000 Co  2 2 c

It indicates high stability and slow decomposition of HF, even at 2000°C.

**(iii) The Effect of Conditions on the Position of Equilibrium**

Equilibrium constant and position of equilibrium are two diferent entities. Kc is equilibrium constant and has constant value at a particular temperature whereas the ratio of products to reactants in equilibrium mixture is described as the position of equilibrium and it can change if the external conditions e.g. temperature, pressure and concentrations are altered. If Kc is large the position of equilibrium lies on the right and if it is small, the position of the equilibrium lies on the left, for a reversible reaction.

Chemists are interested in inding the best conditions to obtain maximum yield of the products in reversible reaction, by favourably shifting the position of equilibrium of a reaction. For this purpose, we have to discuss an important principle in this respect i.e.Le-Chatelier’s principle.

**8.1.6 The Le-Chatelier's Principle**

Le-Chatelier studied the efects of concentration, pressure and temperature on equilibria. **This principle states that if a stress is applied to a system at equilibrium, the system acts in such a way so as to nullify, as far as possible, the efect of that stress.**

The system cannot completely cancel the efect of change, but will minimize it. The LeChatelier’s principle has wide range of applications for ascertaining the position and composition of the physical and chemical equilibria.



*Animation 8.6: The Le-Chatelier’s Principle*

*Source & Credit :*

[*Anotherequilibriumsit*](http://anotherequilibriumsite.weebly.com/le-chacircteliers-principle.html)

[*e*](http://anotherequilibriumsite.weebly.com/le-chacircteliers-principle.html)

**(a) Effect of Change in Concentration**

In order to understand the efect of change in concentration on the reversible reaction, consider the reaction in which BiCl3 reacts with water to give a white insoluble compound BiOCl.

BiCl + H O3 2 ฀BiOCl + 2HCl

The equilibrium constant expression for above reaction can be written as

[BiOCl][HCl]2

Kc =

[BiCl ][H O3 2 ]

Aqueous solution of BiCl3 is cloudy, because of hydrolysis and formation of BiOCl. If a small amount of HCl is added to this solution, it will disturb the equilibrium and force the system to move in such a way so that efect of addition of HCl is minimized. The reaction will move in the backward direction to restore the equilibrium again and a clear solution will be obtained. However, if water is added to the above solution the system will move in the forward direction and the solution will again become cloudy. The shifting of reaction to forward and backward direction by disturbing the concentration is just according to Le-Chatelier’s principle.

So, in general, we conclude that addition of a substance among the reactants, or the removal of a substance among the products at equilibrium stage disturbs the equilibrium position and reaction is shifted to forward direction. Similarly, the addition of a substance among the products or the removal of a substance among the reactants will derive the equilibrium towards the backward direction. Removing one of the products formed can therefore increase the yield of a reversible reaction. The value of K however remains constant. This concept is extensively applied in common ion efect and follows the Le-Chatelier’s principle.

1. **Effect of Change in Pressure or Volume**

The change in pressure or volume are important only for the reversible gaseous reactions where the number of moles of reactants and products are not equal. Le-Chatelier’s principle plays an important role, to predict the position and direction of the reaction. Take the example of formation of SO3 gas from SO2 gas and O2 gas.

2SO (g) + O (g2 2 ) ฀2SO (g3 )

This gas phase reaction proceeds with the decrease in the number of moles and hence decreases in volume at equilibrium stage. When the reaction approaches the equilibrium stage, the volume of the equilibrium mixture is less than the volume of reactants taken initially. If one decreases the volume further at equilibrium stage, the reaction is disturbed. It will move to the forward direction to minimize the efect of disturbance. It establishes a new equilibrium position while Kc remains constant. The reverse happens when the volume is increased or pressure is decreased at equilibrium stage.

1. **Quantitative Effect of Volume on Equilibrium Position**

The quantitative efect of change of volume or pressure can be inferred from the mathematical expression of Kc for SO3 (g) synthesis.

2SO (g) + O (gas2 2 ) ฀2SO (g3 )

4x V K =

(a-2x) (b-x)

Where ‘V’ is the volume of reaction mixture at equilibrium stage, ‘a’ and ‘b’ are the number of moles of SO2 and O2 present initially and Y are the number of moles of oxygen which has reacted at equilibrium. According to the above equation, when volume is increased, then ‘x’ has to be decreased to keep Kc constant. The decrease of x means that reaction is pushed to the backward direction. From the amount of the increase in volume, we can calculate the amount of x which has to be decreased to keep Kc constant

Similarly,increasing the pressure on the above reaction at equilibrium, will decrease tlie volume and hence the value of Kc will increase. In order to keep the value of Kc constant, the reaction will move in the forward direction.

In the same way, we can explain the efect of change of pressure on the equilibrium positions for the dissociation of PCl5 and N2O4 reactions. These reactions are homogenous gaseous phase reactions.

PCI5 dissolves to give PCI3 and Cl2

PCl (g) 5 ฀PCl (g) + Cl (g)3 2

Kc for this reaction is as follows:

x2

K =c

V(a-x)

The dissociation of N2O4 gives NO2 gas

N O (g) 2 4 ฀ 2NO (g2 )

The Kc for this reaction is as follows

4x2

K =c

V(a-x)

Both these reactions have the factor of volume present in the denominator. The reason is that numbers of moles of products are greater than those of reactants. So, increase in pressure will decrease x to keep the value of Kc constant and the reaction will be pushed to the backward direction. The equilibrium position is disturbed but not the Kc value.

Remember that, those gaseous reactions in which number of moles of reactants and products are same, are not afected by change in pressure or volume. Same is the case for reactions in which the participating substances are either liquids or solids.

**(d) Effect of Change in Temperature**

Most of the reversible chemical reactions are disturbed by change in temperature. If we consider heat as a component of equilibrium system, a rise in temperature adds heat to the system and a drop in temperature removes heat from the system. According to Le-Chatelier’s principle, therefore, a temperature increase favours the endothermic reactions and a temperature decrease favours the exothermic reactions.

The equilibrium constant changes by the change of temperature, because the equilibrium position shifts without any substance being added or removed. Consider the following exothermic reaction in gas phase at equilibrium taking place at a known temperature.

CO(g) + H O(g) 2 ฀ CO (g) + H (g) H= -41.84 kJ mole2 2 ∆ -1

At equilibrium stage, if we take out heat and keep the system at this new lower temperature, the system will readjust itself, so as to compensate the loss of heat energy. Thus, more of CO and H2O molecules will react to form CO2 and H2 molecules, thereby, liberating heat because reaction is exothermic in the forward direction. It means by decreasing temperature, we shift the initial equilibrium position to the right until a new equilibrium position is established. On the contrary, heating the reaction at equilibrium will shift the reaction to the backward direction because the backward reaction is endothermic.

An interesting feature of Le-Chatelier’s principle is the efect of temperature on the solubility. Consider a salt such as KI. It dissolves in water and absorbs heat.

Kl(s) ฀ Kl(aq) H=21.4kJ mol∆ −1

Let us have a saturated solution of KI in water at a given temperature. It has attained equilibrium at this temperature. A rise in temperature at equilibrium favours more dissolution of the salt.

Equilibrium is shifted to the forward direction. On the other hand, cooling will favour crystallization of salt. Hence the solubility of KI in water must increase with increase in temperature. For some salts the heat of solution is close to zero (heat is neither evolved or absorbed). The solubility of these salts in water is not afected by the change in temperature. Formation of aqueous solution of NaCl is an example of such a salt.

Those substances, whose heats of solutions are negative (exothermic), decrease their solubilities by increasing temperature, as LiCl and Li2CO3 etc.

**(e) Effect of Catalyst on Equilibrium Constant**

In most of the reversible reactions the equilibrium is not always reached within a suitable short time. So, an appropriate catalyst is added. A catalyst does not afect the equilibrium position of the reaction. It increases the rates of both forward and backward reactions and this reduces the time to attain the state of equilibrium.

Actually, a catalyst lowers the energy of activation of both forward and reverse steps by giving new path to the reaction.

### 8.2 APPLICATIONS OF CHEMICAL EQUILIBRIUM IN INDUSTRY

Concept of chemical equilibrium is widely applicable for preparation of certain materials on industrial scale. Let us discuss the manufacture of NH3 and SO3 gases on industrial scale.

**8.2.1 Synthesis of Ammonia by Haber’s Process**

The process of ammonia synthesis was developed by German chemist F. Haber and irst used in 1933. This process provides an excellent setting in which to apply equilibrium principle and see the compromises needed to make an industrial process economically worth while. The chemical equation is as follows.

N (g) + 3H (g) 2 2 ฀ 2NH (g) H=-92.46kJ3 ∆

When we look at the balanced chemical equation it is inferred, from Le-Chatelier’s principle that one can have three ways to maximize the yield of ammonia.

(i) By continual withdrawl of ammonia after intervals, the equilibrium will shift to forward direction in accordance with Le-Chatelier’s principle. To understand it look at the efect of change of concentration in Le-Chatelier’s principle.

ii) In crease the pressure to decrease the volume of the reaction vessel. Four moles of the reactants combine to give two moles of the products. High pressure will shift the equilibrium position to right to give more and more ammonia.

(iii) Decreasing the temperature will shift it to the forward direction according to Le-Chatelier’s principle.

**Table (8.2) Effect of temperature on**

#### Kc for ammonia synthesis

|  |  |
| --- | --- |
| T(K) | Kc |
| 200  300  400  500  600  700  800 | 7.7x1015  2.69x108 3.94x101 1.72x102  4.53x100  2.96x10-1  3.96x10-2 |

So high pressure, low temperature and continual removal of ammonia will give the maximum yield of ammonia. Table (8.2) shows the efect of the rise in temperature on the value of Kc and the Fig. (8.3) shows the optimum conditions to get maximum yield of ammonia. Fig (8.3) shows percent yield of ammonia vs. temperature (0C) at ive diferent operating pressures. At very high pressure and low temperature (top left), the yield of NH3 is high but the rate of formation is low. Industrial conditions denoted by circle are between 200 and 300 atmospheres at about 4000C.

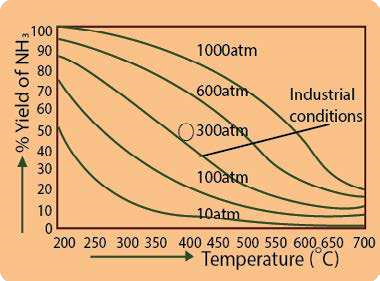
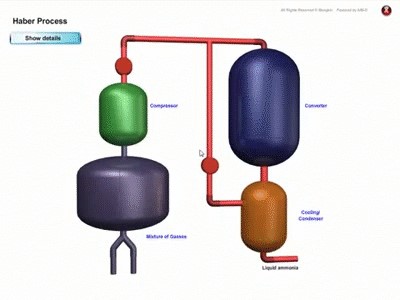


Fig (8.3). Graphical representation of temperature and pressure for NH3 synthesis.

No doubt, the yield of NH3 is favoured at low temperature, but the rate of its formation does not remain favourable. The rate becomes so slow and the process is rendered uneconomical. One needs a compromise to optimize the yield and the rate. The temperature is raised to a moderate level and a catalyst is employed to increase the rate. If one wants to achieve the same rate without a catalyst, then it requires much higher temperature, which lowers the yield. Hence the optimum conditions are the pressure of 200-300 atm and temperature around 673 K (4000C). The catalyst is the pieces of iron crystals embedded in a fused mixture of MgO, Al2O3 and SiO2.

The equilibrium mixture has 35% by volume of ammonia. The mixture is cooled by refrigeration coils until ammonia condenses (B.P = -33.40C) and is removed. Since, boiling points of nitrogen and hydrogen are very low, they remain in the gaseous state and are recycled by pumps back into the reaction chamber.

Nearly 13% of all nitrogen ixation on earth is accomplished industrially through Haber’s process. This process synthesizes approximately 110 million tons of ammonia in the world. About 80% of this is used for the production of fertilizers and some is used in manufacture of explosives or the production of nylon and other polymers.



*Animation 8.7: Haber’s Process*

*Source & Credit :*

[*makeagi*](http://makeagif.com/SW_z6l)

[*f*](http://makeagif.com/SW_z6l)

**8.2.2 Preparation of Sulphur Trioxide**

In the contact process for manufacture of H2SO4, the conversion of SO2 to SO3 is achieved in a reversible reaction.

2SO (g) + O (g) 2 2 ฀ 2SO (g) H=-194kJ/mol3 ∆

The temperature and pressure are the most essential factors for controlling the rate of this reaction. The principles involved here are the same as those discussed previously for Haber’s process. At low temperature, the equilibrium constant for formation of SO3 is large but equilibrium is reached very slowly. As the temperature is raised the rate increases but the yield of SO3 drops of according to Le-Chatelier’s principle. High pressure tends to increase yield of SO3. However, instead of using high pressure, the concentration of O2 (air) is increased to increase the yield of SO3. Table (8.3) helps to understand the efect of diferent conditions on the yield of SO3. During the process pressure is kept at one atmosphere.

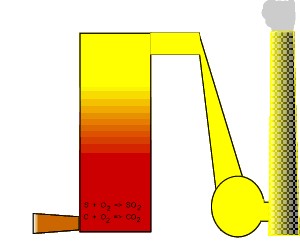
To have the best possible yield of SO3 within a Table (8.3) Efect of temperature reasonable time, a mixture of SO2 and O2 (air) at 1 atm pressure is passed over a solid catalyst at 6500C.

**Table (8.3) Effect of temperature on the yield of SO3**

|  |  |  |
| --- | --- | --- |
| Temp. 0C | Kc | Mole% of SO3 |
| 200  300  400  500  600  700 | 5500  690  160  55  25  13 | 98  91  75  61  46  31 |

The equilibrium mixture is then recycled at lower temperature, 400 to 5000C, to increase the yield of SO3. The most efective catalysts are V2O5 and inely divided platinum. SO3 is dissolved in H2SO4 to get oleum, which is diluted to get H2SO4.

H2SO4 is the king of chemicals. A country’s industrial progress is measured by the amount of H2SO4 manufactured each year.



*Animation 8.8: Preparation of Sulphur Trioxide*

*Source & Credit :*

[*dynamicscienc*](http://www.dynamicscience.com.au/tester/solutions1/chemistry/greenhouse/ghegases.htm)

[*e*](http://www.dynamicscience.com.au/tester/solutions1/chemistry/greenhouse/ghegases.htm)

#### 8.3.0 IONIC PRODUCT OF WATER

Pure water is a very poor conductor of electricity but its conductance is measurable. Water undergoes self ionization as follows and the reaction is reversible.

+ −

H O + H O 2 2 ฀H O + OH3 or H O 2 ฀ H + OH+ −

The equilibrium constant for this reaction can be written as follows.

[H ][OH+ - ] -16 -3

K =c  = 1.8 x 10 moles dm H O2

The concentration of H2O i.e.[H2O] in pure water may be calculated to be 1000gdm3 divided by 18gmol -1 giving 55.5 moles dm-3

Since, water is present in very large excess and very few of its molecules undergo ionization, so its concentration remains efectively constant. Constant concentration of water is taken on L.H.S. and multiplied with Kc to get another constant called Kw.

1.8 x 10-16 x 55.5 = 1.01x10-14 =[H ][OH+ - ]

This 1.01x10-14 is called Kw of water of 250C

K [H O]=[H ][OH ]c 2 + - **Table (8.4) K at various**

**w**

So, K =[H ][OH ] =10+ - -14at 25 C.o **temperatures.**

w

|  |  |
| --- | --- |
| Temp.  (0C) | Kw |
| ne0  10  25  40  100 | 0.11x10-14  0.30x10-14 1.0x10-14  3.00x10-14  7.5x10-14 |

Kw is called ionic product of water or dissociation constant of water. The value of Kw increases almost 75 times when temperature is increased from 0°C to 100°C. Anyhow, the increase in K is not regular. The efect of temperature on K. is shown in Table (8.4).

When ever some quantity of acid or base is added in water, then Kw remains the same, but [H+] and [OH-]are no more equal. Anyhow, in neutral water

[H ] = [OH+ - ]

or [H ][H ] = + + 10−14

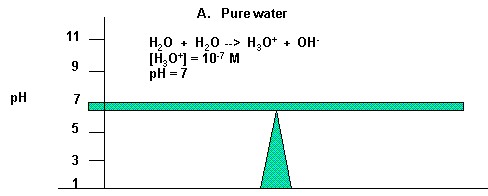
[H ] = + 2 10−14

[H ] = 10 moles dm+ -7 -3

and [OH ] = 10 moles dm− -7 -3

This means that out of 55.5 moles of pure water in one dm3 of it, only 10-7 moles of it have

dissociated into ions. This shows that water is a very weak electrolyte. At 400C, the [H+] = [OH-] but the values are more than 10-7 moles dm-3 and pure water is again neutral at 40°C. Similarly, pure water is neutral at 1000C. [H+] and [OH-] are greater than those at 40°C.



*Animation 8.9: IONIC PRODUCT OF WATER*

*Source & Credit :*

[*employees.csbsj*](http://employees.csbsju.edu/hjakubowski/classes/ch111/olsg-ch111/acidbase/acidsummary.htm)

[*u*](http://employees.csbsju.edu/hjakubowski/classes/ch111/olsg-ch111/acidbase/acidsummary.htm)

In case of addition of small amount of an acid

[H ] > [OH+ - ]

While in the case of addition of few drops of a base

[OH ] > [H− +]

During both of these additions, the value of Kw will remain the same i.e. 10-14 at 250C. **pH and pOH**

Actually, in all the aqueous solutions, the concentration of H+ and OH- are too low to be conveniently expressed and used in calculations. In 1909, Sorenson, a Danish biochemist, introduced the term pH and pOH. So, the scales of pH and pOH have been developed. pH and pOH are abbreviations of negative log of hydrogen ion concentration and negative log of hydroxide ion concentration, respectively.

pH = -log[H ]+

and

pOH = -log[OH ]−

For neutral water, pH = -log10 = 7-7

pOH = -log10 = 7-7

pH = 7, → solution is neutral

when pH < 7, → solution is acidic

pH > 7, → solution is basic

If we take the negative log of Kw, then it is called pKw.

pK = -logKw w

= -log10−14

pK = 14log10w  Since (log 10=1)

pK = 14x1 = 14(at 25 C)w 

The value of pKw is less than 14 at higher tem peratures i.e. at 400Cand 1000C.

The value of pH normally varies between 0 →14 at 25°C. Solutions of negative pH and having values more than 14 are also known. Table (8.5) shows the relationship among [H+],[OH-], pH and pOH of various solutions.

##### Table (8.5) Relationship of [H30+], [OH-], pH and pOH

|  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- | --- |
| More acidic    More basic | |  |  |  |  |  | | --- | --- | --- | --- | --- | |  | [H3O+] | pH | [OH-] | pOH | | Basic | 1x10-14  1x10-13  1x10-12  1x10-11  1x10-10  1x10-9  1x10-8 | 14.0 13.0 12.0 11.0  10.0  9.0  8.0 | 1x10  1x10-1  1x10-2  1x10-3  1x10-4  1x10-5  1x10-6 | 0.0 1.0 2.0 3.0 4.0 5.0  6.0 | | Neutral | 1x10-7 | 7.0 | 1x10-7 | 7.0 | | Acidic | 1x10-6  1x10-5  1x10-4  1x10-3  1x10-2  1x10-1  1x10-0 | 6.0 5.0 4.0 3.0 2.0 1.0  0.0 | 1x10-8  1x10-9  1x10-10  1x10-11  1x10-12  1x10-13  1x10-14 | 8.0  9.0  10.0 11.0 12.0 13.0  14.0 | |

The pH values of some familiar aqueous solutions are shown inTable (8.6). This table can help you to understand the acidic or basic nature of commonly used solutions.

**Table (8.6) Approximate pH and pOH of some common materials at 25°C**

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| Material | pH | pOH | Material | pH | pOH |
| 1.0 M HCl  0.1 M HCl  0.1 M CH3COOH  gastric juice  lemons vinegar soft drinks apples grapefruit oranges tomatoes  cherries bananas | 0.1 1.1 2.9 2.0 2.3 2.8 3.0 3.1 3.1 3.5 4.2 3.6  4.6 | 13.9  12.9  11.10  12.00  11.7  11.2  11.00  10.9 10.9  10.5  9.8  10.4  9.4 | bread potatoes rainwater milk  saliva pure water eggs  0.1 M NaHCO3 seawater milk of magnesia 0.1 M NH3  0.05 M Na2CO3  0.1 M NaOH | 5.5 5.8 6.2  6.5  6.5-6.9  7.0 7.8 8.4  8.5  10.5 11.1 11.6  13.0 | 8.5 8.2 7.8  7.5  7.5-7.1  7.00  6.2 5.6 5.5 3.5 2.9  2.4  1.00 |