Equivalent weight of some Bases

Name of the base	Formula of base	Acidity	Equivalent weight
Sodium hydroxide	$NaOH$	1	40/1 = 40
Potassium hydroxide	KOH	1	56/1 = 56
Ammonium hydroxide	NH_4OH	1	35/1= 35
Magnesium hydroxide	$Mg(OH)_2$	2	58/2 = 29
Calcium hydroxide	$Ca(OH)_2$	2	74/2 = 37

3. Equivalent weight of a salt: It is the ratio of molecular weight to the total valency of cations or anions.

$$E_{Salt} = \frac{\text{Molecular weight}}{\text{Total valency of cations or anions}}$$

Equivalent weight of some Salts

Name of the salt	Formula of salt	Total valency of cations or anions	Equivalent weight
Sodium chloride	$NaCl$	1	58.5/1 = 58.5
Sodium carbonate	Na_2CO_3	2	106/2 = 53
Magnesium chloride	$MgCl_2$	2	95/2 = 47.5
Magnesium sulphate	$MgSO_4$	2	120/2 = 60
Calcium carbonate	$CaCO_3$	2	100/2 = 50
Silver nitrate	$AgNO_3$	1	170/1 = 170
Copper sulphate	$CuSO_4$	2	159.5/2 = 79.75

4. Equivalent weight of an oxidizing agent: It is the ratio of molecular weight to the number of electrons gained

$$E_{OA} = \frac{\text{Molecular weight}}{\text{No. of electrons gained}}$$

Equivalent weight of some oxidising agents

Name of the compound	Formula of compound	No. of electrons gained	Equivalent weight
Potassium permanganate	$KMnO_4$	5 (in acidic medium)	158/5 = 31.6
Potassium permanganate	$KMnO_4$	3 (in neutral)	158/3= 52.6
Potassium dichromate	$K_2Cr_2O_7$	6	294/6 = 49

5. Equivalent weight of a reducing agent: It is the ratio of molecular weight to the number of electrons lost.

$$E_{RA} = \frac{\text{Molecular weight}}{\text{No. of electrons lost}}$$

Equivalent weight of some reducing agents

Name of the compound	Formula of compound	No. of electrons lost	Equivalent weight
Mohr's salt	$FeSO_4(NH_4)_2SO_4 \cdot 6H_2O$	1	392/1 = 392
Hypo	$Na_2S_2O_3$	1	158/1 = 158
Oxalic acid	$H_2C_2O_4 \cdot 2H_2O$	2	126/2 = 63
Ferrous sulphate	$FeSO_4 \cdot 7H_2O$	1	278/1 = 278

6. Equivalent weight of an element: It is the ratio of atomic weight to the valency .

$$E_{ele} = \frac{\text{Atomic weight}}{\text{Valency}}$$

Equivalent weight of some elements

Element	Symbol	Valency	Equivalent weight
Sodium	Na	1	23/1 = 23
Magnesium	Mg	2	24/2 = 12
Aluminium	Al	3	27/3 = 9
Silver	Ag	1	108/1 = 108
Ferrous	Fe^{+2}	2	56/2 = 28
Ferric	Fe^{+3}	3	56/3 = 18.6
Zinc	Zn	2	65.4/2 = 32.7
Cupric	Cu^{+2}	2	63.5/2 = 31.75
Potassium	K	1	39/1 = 39

Example

grams of H_2SO_4 dissolved in 2 litres of water calculate the normality of solution ?.

Solution:

Wt. of solute (W) = 9.8 g
Volume of solution (V) = 2 lit.
GMW of H_2SO_4 = 98 g
Basicity of H_2SO_4 = 2
GEW of H_2SO_4 =98/2 = 49 g
Normality (N) = ?

$$N = \frac{W}{GEW} \times \frac{1}{V \text{ in litres}}$$

$$= \frac{9.8}{49} \times \frac{1}{2}$$

$$= \frac{1}{10} \quad \text{or} \quad 0.1\text{ N}$$

Example

Find the normality of solution prepared by dissolving 10 grams of NaOH in 500 ml of water.

Solution:

$$W = 10\ g$$
$$V = 500\ ml$$
$$\text{GMW of NaOH} = 40\ g$$
$$\text{Acidity of NaOH} = 1$$
$$\text{GEW of NaOH} = \frac{40}{1} = 40\ g$$
$$N = ?$$
$$N = \frac{W}{GEW} \times \frac{1000}{V\ in\ ml}$$
$$= \frac{10}{40} \times \frac{1000}{500} = \frac{1}{2} \quad \text{or} \quad 0.5\ N$$

Example

Calculate the weight of Na_2CO_3 present in 100 ml of 0.5 N solution.

Solution:

$$W = ?$$
$$V = 100\ ml$$
$$N = 0.5$$
$$\text{GMW of } Na_2CO_3 = 106\ g$$
$$\text{GEW of } Na_2CO_3 = \frac{106}{2} = 53\ g$$
$$N = \frac{W}{GEW} \times \frac{1000}{V\ in\ ml}$$
$$W = \frac{N \times GEW \times V}{1000} = \frac{0.5 \times 53 \times 100}{1000} = 2.65\ g$$

Example

Calculate the weight of oxalic acid ($H_2C_2O_4.2H_2O$) required to prepare 0.05 normal solution in 2 litres.

$$W = ?$$

$$N = 0.05$$

$$V = 2 \text{ litres}$$

$$\text{GMW of oxalic acid} = 126\text{ g}$$

$$\text{Basicity} = 2$$

$$\text{GEW of oxalic acid} = 126/2 = 63\text{ g}$$

$$N = \frac{W}{GEW} \times \frac{1}{V \text{ in litres}}$$

$$W = N \times GEW \times V$$

$$= 0.05 \times 63 \times 2$$

$$= 6.3\text{ g}$$

H.W

1-Calculate the mass of compound required to prepare 100.00 mL of a 1.40×10^{-2} M $CuCl_2$?.

2-If 2.50g of NaOH are dissolved in 100.00 mL of water, what is the molarity of the solution ?.

3-If 10.00mL of the solution in question 2 are diluted to 100.00 mL, what will be the molarity of the dilute solution ?.

4-If 10.00 mL of the solution in question 2 are combined with to 25.00 mL of the solution from question 3 in a 100 mL volumetric flask and diluted to 100.00 mL, what will be the molarity of the final solution ?.

5-Calculate the molarity of 29.25 grams of NaCl in 2.0 liters of solution ?.

6-How many grams and moles of solute are there in 250. mL of a 0.10M $CaCl_2$ solution ?.

7-Determine the molarity of these solutions:

a) 4.67 moles of Li_2SO_3 dissolved to make 2.04 liters of solution.
b) 0.629 moles of Al_2O_3 to make 1.500 liters of solution.
c) 4.783 grams of Na_2CO_3 to make 10.00 liters of solution.
d) 0.897 grams of $(NH_4)_2CO_3$ to make 250 mL of solution.
e) 0.0348 grams of $PbCl_2$ to form 45.0 mL of solution.

8- Calculate the concentration of potassium ion in grams per liter after mixing 100 mL of 0.250 M KCl and 200 mL of 0.100 M K_2SO_4 ? .

9-Determine the final volume of these solutions:
a) 4.67 moles of Li_2SO_3 dissolved to make a 3.89 M solution.
b) 4.907 moles of Al_2O_3 to make a 0.500 M solution.
c) 0.783 grams of Na_2CO_3 to make a 0.348 M solution.
d) 8.97 grams of $(NH_4)_2CO_3$ to make a 0.250-molar solution.
e) 48.00 grams of $PbCl_2$ to form a 5.0-molar solution.

10-Calculate molarity and normality of a solution prepared by dissolving 10.6 g of Na_2CO_3 in 2 litres of water ?.

11-Find the normality of a solution prepared by dissolving 1.58 g of $KMnO_4$ is 200 ml of water. (GMW of $KMnO_4$ is 158 g) ?.

2:4:2. Percent Concentration

Chemists frequently express concentrations in terms of percent (parts per hundred).Unfortunately, this practice can be a source of ambiguity because percent composition of a solution can be expressed in several ways. Three common methods are Note that the denominator in each of these expressions refers to the *solution* rather than to the solvent. Note also that the first two expressions do not depend on the units employed (provided, of course, that there is consistency between numerator and denominator).In the third expression, units must be defined because the numerator and denominator have different units that do not cancel. Of the three expressions, only weight percent has the virtue of being temperature independent.

***2:4:2:1.Weight percent*(w/w):** is frequently employed to express the concentration of commercial aqueous reagents. For example, nitric acid is sold as a 70% solution, which means that the reagent contains 70 g of HNO_3 per 100 g of solution .

$$\textbf{weight percent (w/w)} = \frac{\textbf{weight solute}}{\textbf{weight solution}} \times 100\%$$

2:4:2:2.*Volume percent*(v/v): is commonly used to specify the concentration of a solution prepared by diluting a pure liquid compound with another liquid.

For example, a 5% aqueous solution of methanol *usually* describes a solution prepared by diluting 5.0 mL of pure methanol with enough water to give 100 mL.

$$\textbf{volume percent (v/v)} = \frac{\textbf{volume solute}}{\textbf{volume solution}} \times 100\%$$

2:4:2:3.*Weight/volumepercent*(w/v): is often employed to indicate the composition of dilute aqueous solutions of solid reagents. For example, 5% aqueous silver nitrate *often* refers to a solution prepared by dissolving 5 g of silver nitrate in sufficient water to give 100 mL of solution.

$$\textbf{weight/volume percent (w/v)} = \frac{\textit{weight solute, g}}{\textit{volume solution,ml}} \times 100\%$$

Example 4

A concentrated solution of aqueous ammonia is 28.0% w/w NH_3 and has a density of 0.899 g/mL. What is the molar concentration of NH_3 in this solution?

Solution

$$\frac{28.0 \text{ g } NH_3}{100 \text{ g solution}} \times \frac{0.899 \text{ g solution}}{\text{mL solution}} \times \frac{1 \text{ mole } NH_3}{17.04 \text{ g } NH_3} \times \frac{1000 \text{ mL}}{\text{liter}} = 14.8 \text{ M}$$

2:5:3.Parts per Million (ppm) and Parts per Billion (ppb).

For very dilute solutions, parts per million (ppm) is a convenient way to express concentration:

$$C_{ppm}(\text{wt/vol}) = \frac{\text{mass solute (mg)}}{\text{volume solution (L)}}$$

$$C_{ppm}(\text{wt/wt}) = \frac{\text{mass solute (g)}}{\text{mass sample (g)}} \times 10^6 \quad \text{or} \quad C_{ppm}(\text{wt/wt}) = \frac{\text{mass solute (mg)}}{\text{mass sample (kg)}}$$

where c_{ppm} is the concentration in parts per million. Obviously, the units of mass in the numerator and denominator must agree.For even more dilute solutions, 10^9 ppb rather than 10^6 ppm is employed in the foregoing equation to give the results in parts per billion (ppb).

$$C_{ppb}(\text{wt/vol}) = \frac{\text{mass solute (µg)}}{\text{volume solution (L)}}$$

$$C_{ppb}(\text{wt/wt}) = \frac{\text{mass solute (g)}}{\text{mass sample (g)}} \times 10^9 \quad \text{or} \quad C_{ppb}(\text{wt/wt}) = \frac{\text{mass solute (µg)}}{\text{mass sample (kg)}}$$

Common Units for Expressing Trace Concentrations:

Unit	Abbreviation	wt/wt	wt/vol	vol/vol
Part per million	ppm	mg/kg µg/g	mg/L µg/mL	µL/L nL/mL
Part per billion	ppb	µg/kg ng/g	µg/L ng/mL	nL/L pL/mL

pL=picoliter =10^{-12} L , µL =microliter =10^{-6} L , nL = ninoliter =10^{-9} L

Table 2.4 Common Units for Reporting Concentration

Name	Units[a]	Symbol
molarity	$\frac{\text{moles solute}}{\text{liters solution}}$	M
formality	$\frac{\text{number FWs solute}}{\text{liters solution}}$	F
normality	$\frac{\text{number EWs solute}}{\text{liters solution}}$	N
molality	$\frac{\text{moles solute}}{\text{kg solvent}}$	*m*
weight %	$\frac{\text{g solute}}{\text{100 g solution}}$	% w/w
volume %	$\frac{\text{mL solute}}{\text{100 mL solution}}$	% v/v
weight-to-volume %	$\frac{\text{g solute}}{\text{100 mL solution}}$	% w/v
parts per million	$\frac{\text{g solute}}{10^6\ \text{g solution}}$	ppm
parts per billion	$\frac{\text{g solute}}{10^9\ \text{g solution}}$	ppb

[a]FW = formula weight; EW = equivalent weight.

Example 3

A 2.6 g sample of plant tissue was analyzed and found to contain 3.6 µg zinc .What is the concentration of zinc in the plant in ppm ? In ppb?

Solution

$$C_{ppm}\ (\text{wt/wt}) = \frac{\text{mass solute } (\mu g)}{\text{mass sample } (g)} = \frac{3.6\ \mu g}{2.6\ g} = \boxed{1.384\ \text{ppm}}$$

$$C_{ppb}\ (\text{wt/wt}) = \frac{\text{mass solute } (ng)}{\text{mass sample } (g)} = \frac{3.6\text{x}10^3}{2.6} = \boxed{1384.6\ \ \text{ppb}}$$

Example 4

What is the molarity of K^+ in a solution that contains 63.3 ppm of $K_3Fe(CN)_6$ (329.3 g/mol)?

Solution

Because the solution is so dilute, it is reasonable to assume that its density is 1.00 g/mL. Therefore, according to Equation

63.3 ppm $K_3Fe(CN)_6$ = 63.3 mg $K_3Fe(CN)_6$/1L

$$\text{Mol} = \frac{\text{Wt(g)}}{\text{MW}} = \frac{63.3 \text{ mg} \times 10^{-3}}{329.3} = 1.922 \times 10^{-4} \text{ mol}$$

$$\text{M} = \frac{\text{No. of mol}}{\text{Vol. (L)}} = \frac{1.922 \times 10^{-4} \text{ mol}}{1 \text{ L}} = 1.922 \times 10^{-4} \text{ M}$$

$$[K^+] = \frac{(1.922 \times 10^{-4}) \times 3}{1} = \boxed{5.77 \times 10^{-4} \text{ M}}$$

The maximum allowed concentration of chloride in a municipal drinking water supply is 2.50 × 102 ppm Cl^-. When the supply of water exceeds this limit, it often has a distinctive salty taste. What is this concentration in moles Cl^-/liter?

Solution

$$\frac{2.50 \times 10^2 \text{ mg Cl}^-}{\text{L}} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times \frac{1 \text{ mole Cl}^-}{35.453 \text{ g Cl}^-} = 7.05 \times 10^{-3} \text{ M}$$

H.W

1-A solution contains Cu^{2+} ions at a concentration of 3×10^{-4} M. What is the Cu^{2+}concentratrion in ppm?. the solution is in water and the density of water = 1 g/mL (A.W of Cu^{2+}=63.55 g/mol) ?.

2-A 25.0 µL serum sample was analyzed for glucose content and found to contain 26.0 µg .Calculate the concentration of glucose in ppm ?.

2:6.Standard solution.

Is a reagent of known concentration that is used to carry out a titrimetric analysis. A **titration** is performed by slowly adding a standard solution from a burette or other liquid-dispensing device to a solution of the analyte until the reaction between the two is judged complete. The volume or mass of reagent needed to complete the titration is determined from the difference between the initial and final readings.

2:6:1.utionsPreparing Sol .

Preparing a solution of known concentration is perhaps the most common activity in any analytical lab. The method for measuring out the solute and solvent depend on the desired concentration units, and how exact the solution's concentration needs to be known. Pipets and volumetric flasks are used when a solution's concentration must be exact; graduated cylinders, beakers, and reagent bottles suffice when concentrations need only be approximate. Two methods for preparing solutions are described in this section.

Example 5

Describe how you would prepare the following three solutions:

(a) 500 mL of approximately 0.20 M NaOH using solid NaOH;

(b) 1 L of 150.0 ppm Cu^{2+} using Cu metal; and

(c) 2 L of 4% v/v acetic acid using concentrated glacial acetic acid.

Solution

(a) Since the concentration only needs to be known to two significant figures, the mass of NaOH and volume of solution do not need to be measured exactly. The desired mass of NaOH is:

$$\frac{0.20 \text{ mol}}{\text{L}} \times \frac{40.0 \text{ g}}{\text{mol}} \times 0.50 \text{ L} = 4.0 \text{ g}$$

To prepare the solution we place 4.0 g of NaOH, weighed to the nearest tenth of a gram, in a bottle or beaker and add approximately 500 mL of water.

(b) Since the concentration of Cu^{2+} needs to be exact, the mass of Cu metal and the final solution volume must be measured exactly. The desired mass of Cu metal is:

$$\frac{150.0\ \text{mg}}{\text{L}} \times 1.000\ \text{L} = 150.0\ \text{mg} = 0.1500\ \text{g}$$

(c) The concentration of this solution is only approximate, so volumes do not need to be measured exactly. The necessary volume of glacial acetic acid is

$$\frac{4\ \text{mL}\ CH_3COOH}{100\ \text{mL}} \times 2000\ \text{mL} = 80\ \text{mL}\ CH_3COOH$$

To prepare the solution we use a graduated cylinder to transfer 80 mL of glacial acetic acid to a container that holds approximately 2 L, and we then add sufficient water to bring the solution to the desired volume.

Preparing Solutions by Dilution .

Solutions with small concentrations are often prepared by diluting a more concentrated stock solution. A known volume of the stock solution is transferred to a new container and brought to a new volume. Since the total amount of solute is the same before and after **dilution,** we know that

$$\boxed{C\text{o} \times V\text{o} = C\text{d} \times V\text{d}}$$

where *C*o is the concentration of the stock solution, *V*o is the volume of the stock solution being diluted, *C*d is the concentration of the dilute solution, and *V*d is the volume of the dilute solution. Again, the type of glassware used to measure *V*o and *V*d depends on how exact the solution's concentration must be known.

Example 6

If 10 mL (Vd) of a 2-µg/mL (Cd) standard will be prepared from a 20-µg/mL (Co) stock or intermediate standard, the volume of the stock or intermediate standard to be withdrawn can be calculated by using the above equation:

20 µg/mL (C*o*) x V*o* = 2 µg/mL (Cd) x 10 mL (Vd)

V*o* = (2 µg/mL x 10 mL)/20 µg/mL = 1 mL

To prepare the 2-µg/mL solution, 1 mL of the 20-µg/mL standard will be added into a 10-mL volumetric flask which contains approximately one-third of the proper solvent or reagent. The proper solvent or reagent will be added to the mark and mixed well.

Example7

A laboratory procedure calls for 250 mL of an approximately 0.10 M solution of NH_3. Describe how you would prepare this solution using a stock solution of concentrated NH_3 (14.8 M).

Solution

Substituting known volumes in equation

$$C_o \times V_o = C_d \times V_d$$

$$14.8\ \mathrm{M} \times V_o = 0.10\ \mathrm{M} \times 0.25\ \mathrm{L}$$

and solving for V_o gives 1.69×10^{-3} L, or 1.7 mL. Since we are trying to make a solution that is approximately 0.10 M NH_3, we can measure the appropriate amount of concentrated NH_3 using a graduated cylinder, transfer the NH_3 to a beaker, and add sufficient water to bring the total solution volume to approximately 250 mL.

Chemical reactions.

Chemical reactions are the heart of chemistry. ***A chemical reaction is a process that leads to the transformation of one set of chemical substance to another .*** Chemical reactions can be classified as one of six main types: synthesis, decomposition, single replacement, double replacement, neutralization (acid-base), or combustion. You can identify each type of reaction by examining the reactants. This makes it possible to classify a reaction and then predict the identity of the products.

(1)Synthesis: (combination) reaction, two or more reactants (A & B) combine to produce a single product (AB).

element + element ➔ compound

A + B ➔ AB

(The letters A and B represent elements.)

hydrogen + oxygen → water

(2)Decomposition: reaction a compound is broken down into smaller compounds or separate elements. A decomposition reaction is the reverse of a synthesis reaction.

compound ➔ element + element

AB ➔ A + B

calcium chlorate → calcium chloride + oxygen

(3)Single replacement: reaction, a reactive element (a metal or a non-metal) and a compound react to produce another element and another compound. In other words, one of the elements in the compound is replaced by another element. The element that is replaced could be a metal or a non-metal.

element + compound ➔ element + compound

A + BC ➔ B + AC where A is a metal OR

A + BC ➔ C + BA where A is a non-metal

aluminum + lead(II) nitrate → aluminum nitrate + lead

(4)Double replacement: reaction usually involves two ionic solutions that react to produce two other ionic compounds. One of the compounds forms a **precipitate**, which is an insoluble solid that forms from a solution. The precipitate floats in the solution, then settles and sinks to the bottom.
The other compound may also form a precipitate, or it may remain dissolved in solution.

ionic solution + ionic solution ➔ ionic solution + ionic solid

$$AB(aq) + CD(aq) \rightarrow AD(aq) + CB(s)$$

iron(II) chloride + lithium phosphate → iron(II) phosphate + lithium chloride

(5)Neutralization (acid-base): When an acid and a base are combined, they will neutralize each other. In a neutralization (acid-base) reaction, an acid and a base react to form a salt and water.

acid + base ➔ salt + water

$$HX + MOH \rightarrow MX + H_2O$$

(X represents a negative ion. M represents a positive ion.)

sulphuric acid + sodium hydroxide → sodium sulphate + water

(6)Combustion: is the rapid reaction of a compound or element with oxygen to form an oxide and produce heat. For example, organic compounds, such as methane, combust with oxygen to form carbon dioxide (the oxide of carbon) and water (the oxide of hydrogen).

hydrocarbon + oxygen ➔ carbon dioxide + water

$$C_XH_Y + O_2 \rightarrow CO_2 + H_2O$$

(The subscripts X and Y represent integers.)

$$C_2H_6O_3 + O_2 \rightarrow CO_2 + H_2O$$

The summary chart below compares the six types of chemical reactions.

Reaction Type	Reactants and Products	Notes on the Reactants
Synthesis (combination)	$A + B \rightarrow AB$	• Two elements combine
Decomposition	$AB \rightarrow A + B$	• One reactant only
Single replacement If A is a metal If A is a non-metal	 $A + BC \rightarrow B + AC$ $A + BC \rightarrow C + BA$	• One element and one compound
Double replacement	$AB + CD \rightarrow AD + CB$	• Two compounds react.
Neutralization (acid-base)	$HX + MOH \rightarrow MX + H_2O$	• Acid plus base
Combustion	$C_xH_y + O_2 \rightarrow CO_2 + H_2O$	• Organic compound with oxygen

Chemical Equilibrium.4:2

Is the state reached by a reaction mixture when the rates of forward and reverse reactions have become equal. If you observe the reaction mixture, you see no net change, although the forward and reverse reactions are continuing. The continuing forward and reverse reactions make the equilibrium a *dynamic* process.

Reaction rate.4:2:1

The rate of a reaction is the amount of product formed or the amount of reactant used up per unit of time. So that a rate calculation does not depend on the total quantity of reaction mixture used, you express the rate for a unit volume of the mixture.

Therefore, the **reaction rate** **is *the increase in molar concentration of product of a reaction per unit time or the decrease in molar concentration of reactant per unit time.*** The usual unit of reaction rate is moles per liter per second, mol/(L.s).Which is dependent on such factor as the temperature and the presence of catalysts .

4:2:2.Definition of the Equilibrium Constant Kc

Consider the reaction

$aA + bB \rightarrow cC + dD$

where A, B, C, and D denote reactants and products, and *a*, *b*, *c*, and *d* are coefficients in the balanced chemical equation.

The **equilibrium-constant expression** for a reaction is *an expression obtained by multiplying the concentrations of products, dividing by the concentrations of reactants, and raising each concentration term to a power equal to the coefficient in the chemical equation.* The **equilibrium constant** ***Kc*** is *the value obtained for the equilibrium-constant expression when equilibrium concentrations are substituted.* For the previous reaction, you have

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Here you denote the molar concentration of a substance by writing its formula in square brackets. The subscript *c* on the equilibrium constant means that it is defined in terms of molar concentrations. The **law of mass action** **is *a relation that states that the values of the equilibrium-constant expression Kc are constant for a particular reaction at a given temperature, whatever equilibrium concentrations are substituted*.** As the following example illustrates, the equilibrium-constant expression is defined in terms of the balanced chemical equation. If the equation is rewritten with different coefficients, the equilibrium-constant expression will be changed.

Example.1

Write the equilibrium-constant expression *Kc* for catalytic methanation.

$$CO_{(g)} + 3H_{2(g)} \leftrightarrow CH_{4(g)} + H_2O_{(g)}$$

Solution

The expression for the equilibrium constant is:

$$Kc = \frac{[CH_4][H_2O]}{[CO][H_2]^3}$$

Note that concentrations of products are on the top and concentrations of reactants are on the bottom. Also note that each concentration term is raised to a power equal to the coefficient of the substance in the chemical equation.Write the equilibrium-constant expression *Kc* when the equation for the previous reaction is written

$$\tfrac{1}{2} N_{2(g)} + \tfrac{3}{2} H_{2(g)} \leftrightarrow NH_{3(g)}$$

the equilibrium-constant expression becomes:

$$Kc = \frac{[NH_3]}{[N_2]^{1/2} . [H_2]^{3/2}}$$

H.W

1-Write the equilibrium-constant expression *Kc* and what type interaction.

$$N_{2(g)} + 3H_{2(g)} \leftrightarrow 2NH_{3(g)}$$

2- Write the equilibrium-constant expression *Kc* for the equation.

$$2NO_{2(g)} + 7H_{2(g)} \leftrightarrow 2NH_{3(g)} + 4H_2O_{(g)}$$

In the preceding sections, we described how a chemical reaction reaches equilibrium and how you can characterize this equilibrium by the equilibrium constant. Now we want to see the ways in which an equilibrium constant can be used to answer important questions. We will look at the following uses:

(1).Qualitatively interpreting the equilibrium constant.

By merely looking at the magnitude of *Kc*, you can tell whether a particular equilibrium favors products or reactants.

(2).Predicting the direction of reaction.

Consider a reaction mixture that is not at equilibrium. By substituting the concentrations of substances that exist in a reaction mixture into an expression similar to the equilibrium constant and comparing with *Kc*, you can predict whether the reaction will proceed toward products or toward reactants (as defined by the way you write the chemical equation).

(3).Calculating equilibrium concentrations.

Once you know the value of *Kc* for a reaction, you can determine the composition at equilibrium for any set of starting concentrations.

4:3. Le Châtelier's principle

Which states that ***when a system in chemical equilibrium is disturbed by a change of temperature, pressure, or a concentration, the system shifts in equilibrium composition in a way that tends to counteract this change of variable***. Suppose you remove a substance from or add a substance to the equilibrium mixture in order to alter the concentration of the substance. Chemical reaction then occurs to partially restore the initial concentration of the removed or added substance. (However, if the concentration of substance cannot be changed, as in the case of a pure solid or liquid reactant or product, changes in amount will have no effect on the equilibrium composition).

4:3:1.Effect of Concentrations on Equilibria.

When more reactant is added to, or some product is removed from, an equilibrium mixture, there by changing the concentration of reactant or product, net reaction occurs left to right (that is, in the forward direction) to give a new equilibrium, and more products are produced.

When more product is added to, or some reactant is removed from, an equilibrium mixture, thereby changing the concentration of reactant or product, net reaction occurs right to left (that is, in the reverse direction) to give a new equilibrium, and more reactants are produced.

4:3:2. Pressure Effect on Equilibria.

A pressure change *obtained by changing the volume* can affect the yield of product in a gaseous reaction if the reaction involves a change in total moles of gas. The methanation reaction,
$CO + 3H_2 \leftrightarrow CH_4 + H_2O$, is an example of a change in moles of gas. When the reaction goes in the forward direction, four moles of reactant gas ($CO + 3H_2$) become two moles of product gas ($CH_4 + H_2O$). To see the effect of such a pressure change, consider what happens when an equilibrium mixture from the methanation reaction is compressed to one-half of its original volume at a fixed temperature The total pressure is doubled (PV= constant at a fixed temperature, according to Boyle's law, so halving V requires that P double).

Because the partial pressures and therefore the concentrations of reactants and products have changed, the mixture is no longer at equilibrium. The direction in which the reaction goes to re-establish equilibrium can be predicted by applying Le Chatelier's principle. Reaction should go in the forward direction, because then the moles of gas decrease, and the pressure (which is proportional to moles of gas) decreases. In this way, the initial pressure increase is partially reduced.

If the pressure is increased by decreasing the volume of a reaction mixture, the reaction shifts in the direction of fewer moles of gas.

4:3:3. Temperature Effect on Equilibria.

Temperature has a profound effect on most reactions. In the first place, reaction rates usually increase with an increase in temperature, meaning that equilibrium is reached sooner. Many gaseous reactions are sluggish or have imperceptible rates at room temperature but speed up enough at higher temperature to become commercially feasible processes. Second, equilibrium constants vary with temperature. Whether you should raise or lower the temperature of a reaction mixture to increase the equilibrium amount of product can be shown by Le Chatelier's principle.

For an endothermic reaction (ΔH^o positive), the amounts of products are increased at equilibrium by an increase in temperature (Kc is larger at higher T). For an exothermic reaction (ΔH^o negative), the amounts of products are increased at equilibrium by a decrease in temperature (Kc is larger at lower T).

4:3:3. Catalysts Effect on Equilibria.

Catalysts either speed up or retard the rate at which an equilibrium is attained by affecting the rates of both the forward and the backward reactions .But catalysts affect both rates to same extent and thus have no effect on the value of an equilibrium constant .

Example

$[NH_3] = 3.1\times10^{-2}$ M , $[N_2] = 8.5\times10^{-2}$ M , $[H_2] = 3.1\times10^{-3}$ M

$$N_{2(g)} + 3H_{2(g)} \leftrightarrow 2NH_{3(g)}$$

What is the value of K? .

Solution

$$K = \frac{[NH_3]^2}{[N_2]\,[H_2]^3} = \frac{(3.1 \text{ X } 10^{-2})^2}{(8.5 \text{ X } 10^{-1})\,(3.1 \text{ X } 10^{-3})^3} = 3.8 \text{ X} 10^4$$

Example

The equilibrium composition is 0.613 mol CO, 1.839 mol H2, 0.387 mol CH_4, and 0.387 mol H_2O. The volume of the reaction vessel is 10.00 L, so the equilibrium concentration of CO is:

Solution

$$M = \frac{Mol}{V} \qquad [CO] = \frac{0.613 \text{ mol}}{10.00 \text{ L}} = 0.0613\ M$$

Similarly, the other equilibrium concentrations are $[H_2] = 0.1839\ M$, $[CH_4] = 0.0387\ M$, and $[H_2O] = 0.0387\ M$.
Substitution of concentrations into the equilibrium-constant expression gives:

$$K_c = \frac{[CH_4][H_2O]}{[CO][H_2]^3} = \frac{(0.0387)(0.0387)}{(0.0613)(0.1839)^3} = 3.93$$

Example

A gaseous mixture contains 0.30 mol CO, 0.10 mol H_2, and 0.020 mol H_2O, plus an unknown amount of CH_4, in each liter. This mixture is at equilibrium at 1200 K.

$$CO(g) + 3H_2(g) \leftrightarrow CH_4(g) + H_2O(g)$$

What is the concentration of CH_4 in this mixture?

Solution:

The equilibrium-constant equation is:

$$Kc = \frac{[CH_4][H_2O]}{[CO][H_2]^3}$$

Substituting the known concentrations and the value of Kc gives:

$$3.92 = \frac{[CH_4](0.020)}{(0.30)(0.10)^3}$$

You can now solve for $[CH_4]$.

$$[CH_4] = \frac{(3.92)(0.30)(0.10)^3}{(0.020)} = 0.059$$

The concentration of CH_4 in the mixture is 0.059 mol/L.

H.W.

1-The equilibrium constant *Kc* for the reaction

$$2NO_{(g)} + O_{2(g)} \leftrightarrow 2NO_{2(g)}$$

equals 4.0×10^{-13} at 25°C. Does the equilibrium mixture contain predominantly reactants or products? If $[NO] = [O_2] = 2.0 \times 10^{-6}\ M$ at equilibrium, what is the equilibrium concentration of NO_2?

2- Phosphorus pentachloride gives an equilibrium mixture of PCl_5, PCl_3, and Cl_2 when heated.

$$PCl_{5(g)} \leftrightarrow PCl_{3(g)} + Cl_{2(g)}$$

A 1.00-L vessel contains an unknown amount of PCl_5 and 0.020 mol each of PCl_3 and Cl_2 at equilibrium at 250°C. How many moles of PCl_5 are in the vessel if *Kc* for this reaction is 0.0415 at 250°C ? .