



## 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 <sub>4</sub> OH  | 1       | 35/1 = 35         |
| Magnesium hydroxide | Mg(OH) <sub>2</sub> | 2       | 58/2 = 29         |
| Calcium hydroxide   | Ca(OH) <sub>2</sub> | 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.

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

## 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 <sub>2</sub> CO <sub>3</sub> | 2                                  | 106/2 = 53        |
| Magnesium chloride | MgCl <sub>2</sub>               | 2                                  | 95/2 = 47.5       |
| Magnesium sulphate | MgSO <sub>4</sub>               | 2                                  | 120/2 = 60        |
| Calcium carbonate  | CaCO <sub>3</sub>               | 2                                  | 100/2 = 50        |
| Silver nitrate     | AgNO <sub>3</sub>               | 1                                  | 170/1 = 170       |
| Copper sulphate    | CuSO <sub>4</sub>               | 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 <sub>4</sub>                             | 5 (in acidic medium)    | 158/5 = 31.6      |
| Potassium permanganate | KMnO <sub>4</sub>                             | 3 (in neutral)          | 158/3 = 52.6      |
| Potassium dichromate   | K <sub>2</sub> Cr <sub>2</sub> O <sub>7</sub> | 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 <sub>4</sub> (NH <sub>4</sub> ) <sub>2</sub> SO <sub>4</sub> · 6H <sub>2</sub> O | 1                     | 392/1 = 392       |
| Hypo                 | Na <sub>2</sub> S <sub>2</sub> O <sub>3</sub>                                         | 1                     | 158/1 = 158       |
| Oxalic acid          | H <sub>2</sub> C <sub>2</sub> O <sub>4</sub> · 2H <sub>2</sub> O                      | 2                     | 126/2 = 63        |
| Ferrous sulphate     | FeSO <sub>4</sub> · 7H <sub>2</sub> O                                                 | 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 <sup>+2</sup> | 2       | $56/2 = 28$       |
| Ferric    | Fe <sup>+3</sup> | 3       | $56/3 = 18.6$     |
| Zinc      | Zn               | 2       | $65.4/2 = 32.7$   |
| Cupric    | Cu <sup>+2</sup> | 2       | $63.5/2 = 31.75$  |
| Potassium | K                | 1       | $39/1 = 39$       |

**Example**

grams of H<sub>2</sub>SO<sub>4</sub> 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<sub>2</sub>SO<sub>4</sub> = 98 g

Basicity of H<sub>2</sub>SO<sub>4</sub> = 2

GEW of H<sub>2</sub>SO<sub>4</sub> =  $98/2 = 49$  g

Normality (N) = ?

$$\begin{aligned}
 N &= \frac{W}{\text{GEW}} \times \frac{1}{V \text{ in litres}} \\
 &= \frac{9.8}{49} \times \frac{1}{2} \\
 &= \frac{1}{10} \quad \text{or} \quad 0.1 \text{ N}
 \end{aligned}$$

**Example**

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

*Solution:*

$$W = 10 \text{ g}$$

$$V = 500 \text{ ml}$$

$$\text{GMW of NaOH} = 40 \text{ g}$$

$$\text{Acidity of NaOH} = 1$$

$$\text{GEW of NaOH} = \frac{40}{1} = 40 \text{ g}$$

$$N = ?$$

$$N = \frac{W}{\text{GEW}} \times \frac{1000}{V \text{ in ml}}$$

$$= \frac{10}{40} \times \frac{1000}{500} = \frac{1}{2} \text{ or } 0.5 \text{ N}$$

**Example**

Calculate the weight of  $\text{Na}_2\text{CO}_3$  present in 100 ml of 0.5 N solution.

*Solution:*

$$W = ?$$

$$V = 100 \text{ ml}$$

$$N = 0.5$$

$$\text{GMW of Na}_2\text{CO}_3 = 106 \text{ g}$$

$$\text{GEW of Na}_2\text{CO}_3 = \frac{106}{2} = 53 \text{ g}$$

$$N = \frac{W}{\text{GEW}} \times \frac{1000}{V \text{ in ml}}$$

$$W = \frac{N \times \text{GEW} \times V}{1000} = \frac{0.5 \times 53 \times 100}{1000} = 2.65 \text{ g}$$

**Example**

Calculate the weight of oxalic acid ( $\text{H}_2\text{C}_2\text{O}_4 \cdot 2\text{H}_2\text{O}$ ) required to prepare 0.05 normal solution in 2 litres.

$$W = ?$$

$$N = 0.05$$

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

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

$$\text{Basicity} = 2$$

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

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

$$\begin{aligned} W &= N \times \text{GEW} \times V \\ &= 0.05 \times 63 \times 2 \\ &= 6.3 \text{ g} \end{aligned}$$

**H.W**

1- Calculate the mass of compound required to prepare 100.00 mL of a  $1.40 \times 10^{-2}$  M  $\text{CuCl}_2$  ?.

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  $\text{CaCl}_2$  solution ?.

7- Determine the molarity of these solutions:



- a) 4.67 moles of  $\text{Li}_2\text{SO}_3$  dissolved to make 2.04 liters of solution.  
 b) 0.629 moles of  $\text{Al}_2\text{O}_3$  to make 1.500 liters of solution.  
 c) 4.783 grams of  $\text{Na}_2\text{CO}_3$  to make 10.00 liters of solution.  
 d) 0.897 grams of  $(\text{NH}_4)_2\text{CO}_3$  to make 250 mL of solution.  
 e) 0.0348 grams of  $\text{PbCl}_2$  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  $\text{K}_2\text{SO}_4$  ? .
- 9-Determine the final volume of these solutions:  
 a) 4.67 moles of  $\text{Li}_2\text{SO}_3$  dissolved to make a 3.89 M solution.  
 b) 4.907 moles of  $\text{Al}_2\text{O}_3$  to make a 0.500 M solution.  
 c) 0.783 grams of  $\text{Na}_2\text{CO}_3$  to make a 0.348 M solution.  
 d) 8.97 grams of  $(\text{NH}_4)_2\text{CO}_3$  to make a 0.250-molar solution.  
 e) 48.00 grams of  $\text{PbCl}_2$  to form a 5.0-molar solution.
- 10-Calculate molarity and normality of a solution prepared by dissolving 10.6 g of  $\text{Na}_2\text{CO}_3$  in 2 litres of water ?.
- 11-Find the normality of a solution prepared by dissolving 1.58 g of  $\text{KMnO}_4$  in 200 ml of water. (GMW of  $\text{KMnO}_4$  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  $\text{HNO}_3$  per 100 g of solution .



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

**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.

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

**2:4:2:3. Weight/volume percent(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.

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

#### **Example 4**

A concentrated solution of aqueous ammonia is 28.0% w/w  $\text{NH}_3$  and has a density of 0.899 g/mL. What is the molar concentration of  $\text{NH}_3$  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:

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

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

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).

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

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

### Common Units for Expressing Trace Concentrations:

| Unit             | Abbreviation | wt/wt            | wt/vol           | vol/vol         |
|------------------|--------------|------------------|------------------|-----------------|
| Part per million | ppm          | mg/kg            | mg/L             | $\mu\text{L/L}$ |
|                  |              | $\mu\text{g/g}$  | $\mu\text{g/mL}$ | nL/mL           |
| Part per billion | ppb          | $\mu\text{g/kg}$ | $\mu\text{g/L}$  | nL/L            |
|                  |              | ng/g             | ng/mL            | pL/mL           |

$$pL = \text{picoliter} = 10^{-12} L, \quad \mu L = \text{microliter} = 10^{-6} L, \quad nL = \text{nanoliter} = 10^{-9} L$$



| <b>Table 2.4 Common Units for Reporting Concentration</b> |                                                           |               |
|-----------------------------------------------------------|-----------------------------------------------------------|---------------|
| <b>Name</b>                                               | <b>Units<sup>a</sup></b>                                  | <b>Symbol</b> |
| 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                                                  | $\frac{\text{moles solute}}{\text{kg solvent}}$           | <i>m</i>      |
| weight %                                                  | $\frac{\text{g solute}}{100 \text{ g solution}}$          | % w/w         |
| volume %                                                  | $\frac{\text{mL solute}}{100 \text{ mL solution}}$        | % v/v         |
| weight-to-volume %                                        | $\frac{\text{g solute}}{100 \text{ mL solution}}$         | % 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           |

<sup>a</sup>FW = formula weight; EW = equivalent weight.

**Example 3**

A 2.6 g sample of plant tissue was analyzed and found to contain 3.6  $\mu\text{g}$  zinc .What is the concentration of zinc in the plant in ppm ? In ppb?

**Solution**

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

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

**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

$$63.3 \text{ ppm } K_3Fe(CN)_6 = 63.3 \text{ mg } K_3Fe(CN)_6/1L$$

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

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

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

The maximum allowed concentration of chloride in a municipal drinking water supply is  $2.50 \times 10^2$  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**

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

**H.W**

1-A solution contains  $Cu^{2+}$  ions at a concentration of  $3 \times 10^{-4}$  M. What is the  $Cu^{2+}$  concentration 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  $\mu$ L serum sample was analyzed for glucose content and found to contain 26.0  $\mu$ g .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  $\text{Cu}^{2+}$  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  $\text{Cu}^{2+}$  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

$$\frac{4 \text{ mL CH}_3\text{COOH}}{100 \text{ mL}} \times 2000 \text{ mL} = 80 \text{ mL CH}_3\text{COOH}$$

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

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

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 ( $V_d$ ) of a 2- $\mu\text{g}/\text{mL}$  ( $C_d$ ) standard will be prepared from a 20- $\mu\text{g}/\text{mL}$  ( $C_o$ ) stock or intermediate standard, the volume of the stock or intermediate standard to be withdrawn can be calculated by using the above equation:

$$20 \mu\text{g}/\text{mL} (C_o) \times V_o = 2 \mu\text{g}/\text{mL} (C_d) \times 10 \text{ mL} (V_d)$$

$$V_o = (2 \mu\text{g}/\text{mL} \times 10 \text{ mL}) / 20 \mu\text{g}/\text{mL} = 1 \text{ mL}$$



To prepare the 2- $\mu\text{g}/\text{mL}$  solution, 1 mL of the 20- $\mu\text{g}/\text{mL}$  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  $\text{NH}_3$ . Describe how you would prepare this solution using a stock solution of concentrated  $\text{NH}_3$  (14.8 M).

### **Solution**

Substituting known volumes in equation

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

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

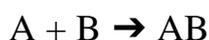
and solving for  $V_o$  gives  $1.69 \times 10^{-3} \text{ L}$ , or 1.7 mL. Since we are trying to make a solution that is approximately 0.10 M  $\text{NH}_3$ , we can measure the appropriate amount of concentrated  $\text{NH}_3$  using a graduated cylinder, transfer the  $\text{NH}_3$  to a beaker, 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).



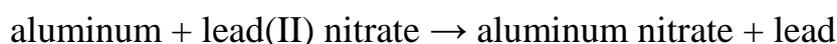
(The letters A and B represent elements.)



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



**(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.

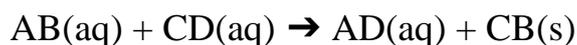




**(4) Double replacement:** reaction usually involves two ionic solutions that react to produce two other ionic compounds. One of the compounds forms a **precipitate**, which is an insoluble solid that forms from a solution. The precipitate floats in the solution, then settles and sinks to the bottom.

The other compound may also form a precipitate, or it may remain dissolved in solution.

ionic solution + ionic solution  $\rightarrow$  ionic solution + ionic solid



iron(II) chloride + lithium phosphate  $\rightarrow$  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  $\rightarrow$  salt + water



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

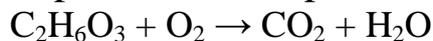
sulphuric acid + sodium hydroxide  $\rightarrow$  sodium sulphate + water

**(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  $\rightarrow$  carbon dioxide + water



(The subscripts X and Y represent integers.)





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<br>If A is a metal<br>If A is a non-metal | $A + BC \rightarrow B + AC$<br>$A + BC \rightarrow C + BA$ | • 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 $K_c$

Consider the reaction



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**  $K_c$  is *the value obtained for the equilibrium-constant expression when equilibrium concentrations are substituted*.

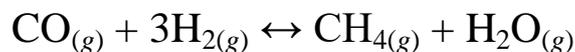
For the previous reaction, you have

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

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  $K_c$  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**

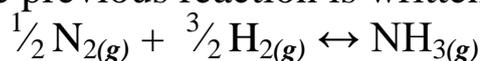
Write the equilibrium-constant expression  $K_c$  for catalytic methanation.

**Solution**

The expression for the equilibrium constant is:

$$K_c = \frac{[\text{CH}_4][\text{H}_2\text{O}]}{[\text{CO}][\text{H}_2]^3}$$

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  $K_c$  when the equation for the previous reaction is written

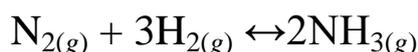


the equilibrium-constant expression becomes:

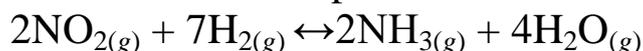
$$K_c = \frac{[\text{NH}_3]}{[\text{N}_2]^{1/2} \cdot [\text{H}_2]^{3/2}}$$

H.W

1-Write the equilibrium-constant expression  $K_c$  and what type interaction.



2- Write the equilibrium-constant expression  $K_c$  for the equation.



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  $K_c$ , 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  $K_c$ , 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  $K_c$  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,  $\text{CO} + 3\text{H}_2 \leftrightarrow \text{CH}_4 + \text{H}_2\text{O}$ , is an example of a change in moles of gas. When the reaction goes in the forward direction, four moles of reactant gas ( $\text{CO} + 3\text{H}_2$ ) become two moles of product gas ( $\text{CH}_4 + \text{H}_2\text{O}$ ). 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 ( $PV = \text{constant}$  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



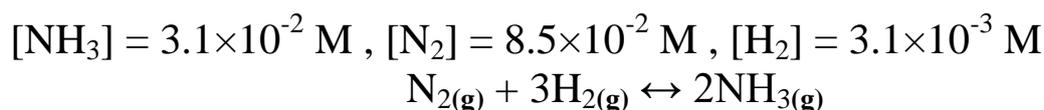
reaction mixture to increase the equilibrium amount of product can be shown by Le Chatelier's principle.

*For an endothermic reaction ( $\Delta H^\circ$  positive), the amounts of products are increased at equilibrium by an increase in temperature ( $K_c$  is larger at higher  $T$ ). For an exothermic reaction ( $\Delta H^\circ$  negative), the amounts of products are increased at equilibrium by a decrease in temperature ( $K_c$  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



What is the value of  $K$ ?

#### Solution

$$K = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(3.1 \times 10^{-2})^2}{(8.5 \times 10^{-1})(3.1 \times 10^{-3})^3} = 3.8 \times 10^4$$

#### Example

The equilibrium composition is 0.613 mol CO, 1.839 mol H<sub>2</sub>, 0.387 mol CH<sub>4</sub>, and 0.387 mol H<sub>2</sub>O. The volume of the reaction vessel is 10.00 L, so the equilibrium concentration of CO is:

#### Solution

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



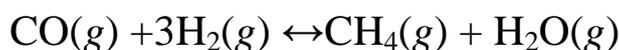
Similarly, the other equilibrium concentrations are  $[H_2] = 0.1839 M$ ,  $[CH_4] = 0.0387 M$ , and  $[H_2O] = 0.0387 M$ .

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.



What is the concentration of  $CH_4$  in this mixture?

### **Solution:**

The equilibrium-constant equation is:

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

**Substituting the known concentrations and the value of  $K_c$  gives:**

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

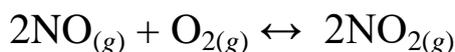
**You can now solve for  $[CH_4]$ .**

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

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

### **H.W.**

**1-**The equilibrium constant  $K_c$  for the reaction



equals  $4.0 \times 10^{-13}$  at  $25^\circ C$ . 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  $\text{PCl}_5$ ,  $\text{PCl}_3$ , and  $\text{Cl}_2$  when heated.



A 1.00-L vessel contains an unknown amount of  $\text{PCl}_5$  and 0.020 mol each of  $\text{PCl}_3$  and  $\text{Cl}_2$  at equilibrium at  $250^\circ\text{C}$ . How many moles of  $\text{PCl}_5$  are in the vessel if  $K_c$  for this reaction is 0.0415 at  $250^\circ\text{C}$  ? .