Analytical Chemistry for Pharmacy Students

By

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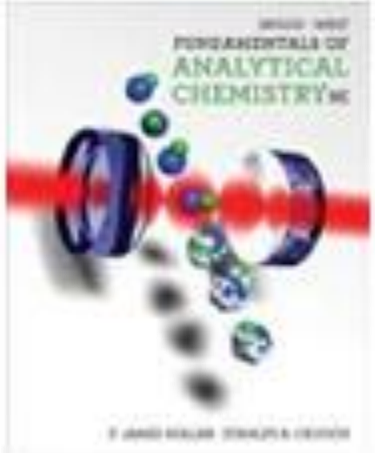
References



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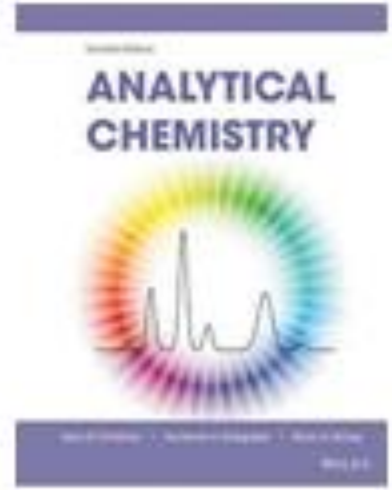

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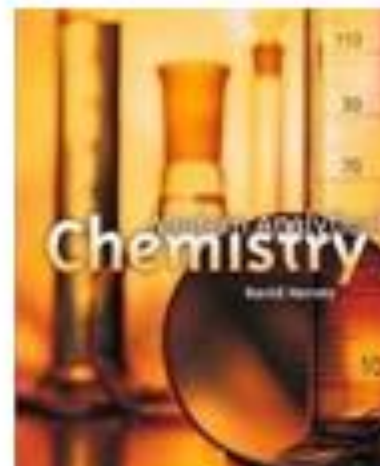

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Lecture 1-2

Introduction

- Chemistry play a very important role in all areas of science and technology :
- **A – health and medicine**
- B – energy and environment's
- C – materials and technology
- D – food and agriculture

Lecture 1-2

Introduction

- Chemistry play a very important role in all areas of science and technology :
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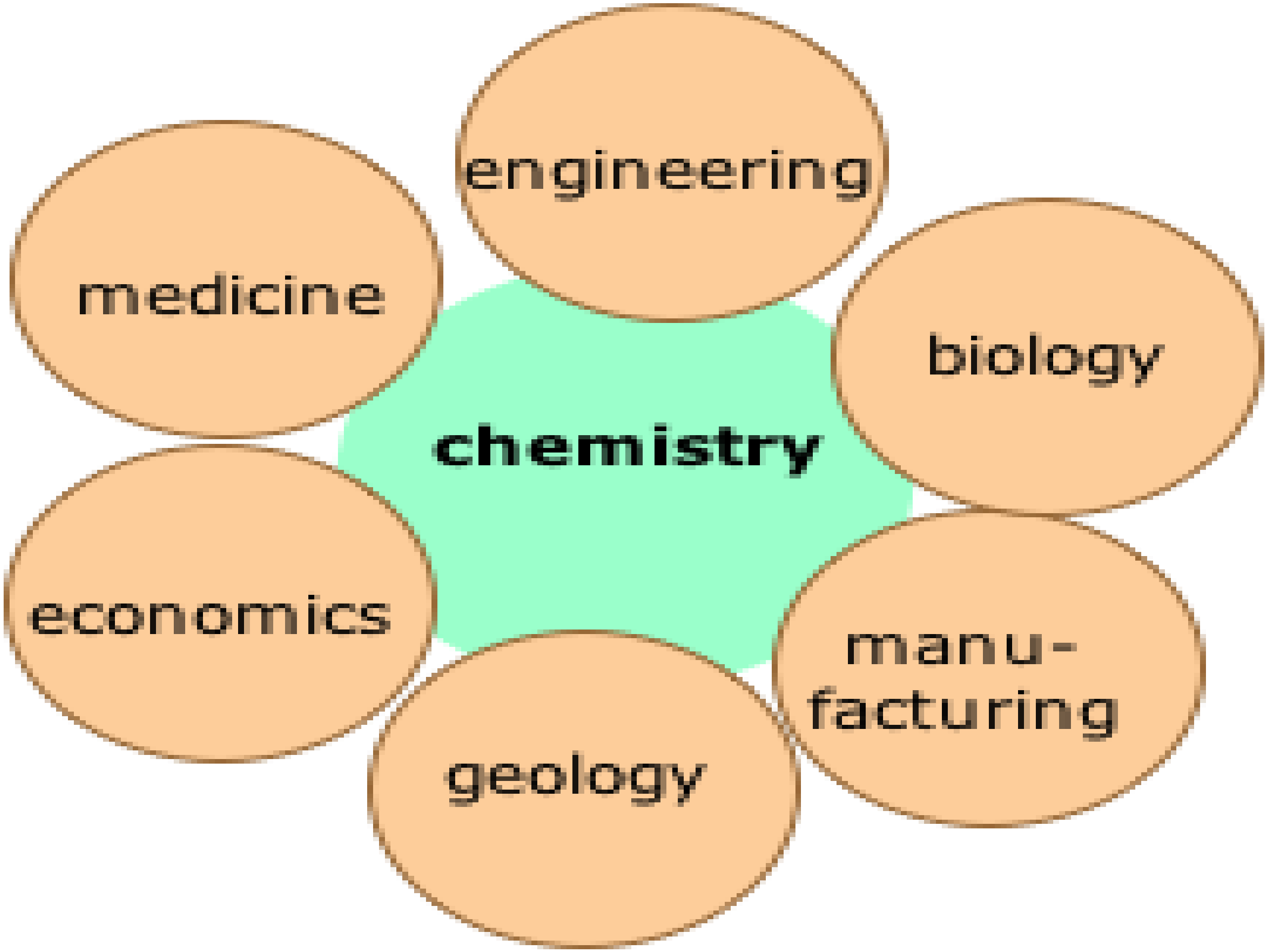
Chemistry: the central science



What is chemistry

Chemistry is that branch of science that deals with chemicals.

The chemicals include all the day-today things you touch, see and smell, eg. your body , the book you touch , the air you smell . Chemicals are everywhere (except in the vacuum).

Many chemicals occur :

a.naturally; metals ,oil, sand.

b.produced by living things ; animals (wool, honey)

c. vegetables (cotton, sugar).

d. synthetic chemicals are also known (**drugs** , clothes).

- Despite the benefits of chemistry , we have to be careful in dealing with chemicals . Many chemicals are toxic, other are potential cancer producer . Therefore chemicals must be handled with control.
- **What are chemicals composed of ?**

Chemicals are composed of different substances (substance mean chemical material of which an object is composed eg. Ice is composed of substance water , or chemical material water).

- An important aspect of chemistry is the study of *chemical reactions* , that is changes that occur when chemicals interact with each other to form a new and entirely different substances .
- Each chemical has its characteristic properties, eg sodium metal combine very rapidly with oxygen and moisture in the air forming crust (sodium oxide) .

  Sodium + oxygen → sodium oxide
- if sodium metal is places in water , it reacts violently producing hydrogen gas and sodium hydroxide .

  Sodium + water → sodium hydroxide + hydrogen
- these characters known as chemical properties

(C.F physical properties as shape, color , odor , taste specific gravity) .

The study of chemistry

- ***Chemistry*** is the study of *substances*; their properties, structure, and the changes they undergo.

  the study of chemistry involves three basic steps :

A – observation

B – representation

C – interpretation

Chemical Composition

- **Classification of Matter**

A. Matter can be classified by its state.

- **Solids** have a definite volume and shape.
- **Liquids** have a definite volume, but change shape.
- **Gases** have neither definite volume nor shape.

B.Matter can also be classified by its composition.

- An **element** is a pure substance made up of atoms with the same number of protons. As of 2007, 117 elements have been observed, 92 of which occur naturally. Carbon (C), oxygen (O), hydrogen (H) are examples of elements. The **periodic table** is a tabular representation of the known elements .
- A **compound** consists of two or more chemical elements that are chemically bonded together. Water (H_2O) and table sugar ($C_{12}H_{22}O_{11}$) are examples of chemical compounds. The ratio of the elements in a compound is always the same. For example in water, the number of H atoms is always twice the number of O atoms .
- A **mixture** consists of two or more substances (element or compound) mixed together without any chemical bond. Salad is a good example. A mixture can also be separated into its individual components by mechanical means .

Elements and compounds

- get substances known as ***elements*** that cannot be reduced to any simpler forms by ordinary chemical or physical means.
- the compound mercuric oxide can be broken down by heating into two other substances:

$$2\ HgO \rightarrow 2\ Hg + O_2$$

- ... but the two products, metallic mercury and dioxygen, cannot be decomposed into simpler substances, so they must be elements.
- Similarly $2\ H_2O \rightarrow 2\ H_2 + O_2$

the compound water can be decompose (electrolysis) into the elements oxygen and hydrogen.

Elements and atoms

The ***atom***, by contrast, is a *microscopic* concept which in modern chemistry relates the unique character of every chemical element to an actual physical particle.

Formula and structure

- **The formula** of a substance expresses the relative number of atoms of each element it contains. (eg. $C_6H_{12}O_6$) (H_2O)
- ***Structure***, which in its greatest detail reveals the relative locations (in two or three dimensional space) of each atom within the minimum collection needed to define the structure of the substance.
- The ordinary chemical formula does not tell us about the order in which the component atoms are connected, whether they are grouped into discrete units (***molecules***) or are two- or three dimensional extended structures,

Measurement

- Different instruments enable us to measure substances properties.
- Buret , pipet ,graduated cylender , volumetric flask <u>measure volum</u>.
- Balance <u>measure mass</u>
- Thermometer <u>measure temperature</u>
- A measured quantity is usually written as a number with appropriate unit.

International system of unit
SI Units

Base quantity	name of unit	symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electric current	ampere	A
Temperature	kelvin	K
Amount of substance	**mole**	**mol**
Luminous intensity	candela	cd

Prefixes used with units

Prefix	symbol	meaning	example
Giga	G	10^9	1 gigameter(Gm) = 1×10^9 m
Mega	M	10^6	1 megamete (Mm) = 1×10^6m
Kilo	k	10^3	1 kilometer (km) = $1 \times 10^{3\ m}$
deci	d	10^{-1}	1 decimeter (dm) = 0.1 m
centi	c	10^{-2}	1 centimeter (cm) = 0.01 m
Milli	m	10^{-3}	1 millimeter (mm) = 0.001 m
Micro	µ	10^{-6}	1 micrometer (µm) = 1×10^{-6} m
Nano	n	10^{-9}	1 nanometer (nm) = 1×10^{-9} m
Pico	p	10^{-12}	1 picomer (pm) = 1×10^{-12} m

The SI unit of mass is kg :

$1\ kg = 1 \times 10^3\ g$

$gm = 1 \times 10^3\ mg$

$mg = 1 \times 10^3\ ug$

$um = 1 \times 10^3\ ng$

$ng = 1 \times 10^3\ pico\ g$

The SI unit of length is m:

$1m = 100\ cm$ $1cm = 10\ mm$

$1\ cm^3 = (1 \times 10^{-2}\ m)^3 = 1 \times 10^{-6}\ m^3$

$1\ dm^3 = (1 \times 10^{-1}\ m)^3 = 1 \times 10^{-3}\ m^3$

A common unit of volum is the liter (L)

$1\ L = 1000\ ml$

$= 1000\ cm^3$

$= 1\ dm^3$

$1\ ml = 1\ cm^3$

Chemical symbol

- **Elements :**

Li, lithium; Na , sodium
F, fluorine; , k , potassium
P, phosphorus; . ca , calcium
Cu, copper; H , hydrogen
As, arsenic; O , oxygen
Zn, zinc; Cl, chlorine;
Pt, platinum; Mg, magnesium;
U, uranium; Al, aluminum;
Si, silicon; Ne, neon.

Compounds :

water , H_2O , hydrochloric acid HCl , sodium chloride NaCl

Lecture 2

The Mole Concept (Avogadro's Number)

- Molecules and atoms are extremely small objects - both in size and mass. Consequently, working with them in the laboratory requires a large collection of them. A standard needs to be introduced. This standard is the ***"mole".*** The mole is based upon the carbon-12 isotope.
- How many carbon-12 atoms are needed to have a mass of exactly 12 g. That number is NA - ***Avogadro's number***. Thus, NA is defined by
- NA = 6.0221367×10^{23}.
- NA x (mass of carbon-^{12} atom) = 12 g

- Careful measurements yield a value for

  $NA = 6.0221367x10^{23}$.
- This is an incredibly large number .
- A convenient name is given when there is an Avogadro's number of objects - it is called a ***"mole".***

  1 mole = NA objects

  NA x 12 amu = 12 g

  (amu : mass of average atom)
- Thus, a mole of carbon-^{12} atoms has a mass of just 12 g.

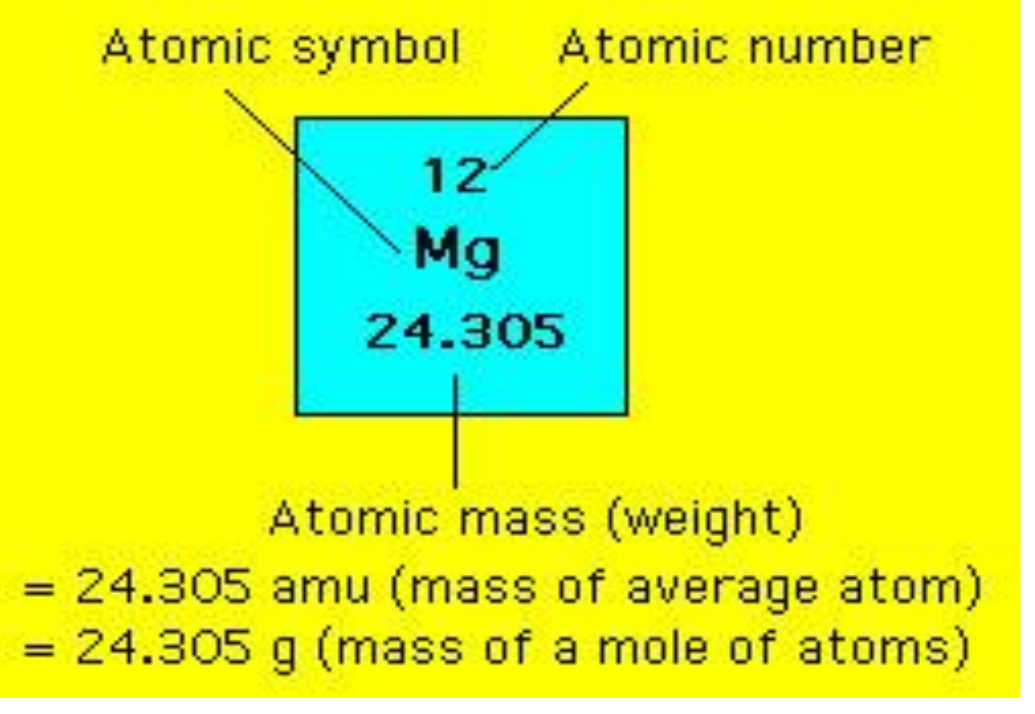
The mole is a unit that is defined as 6.023×10^{23} particles.

It is also equal to the formula mass of a substance expressed in grams.

For example:

1 mole of H_2O = 6.023×10^{23} molecules of H_2O

= 18 grams H_2O



Molar mass of an element= no of atoms X the atomic weight of each element. And,
molar mass of the compound= the sum total of molar mass of all the elements.
That is, the molar mass of NaCl
Sodium (Na) = 1 X 22.98= 22.98 g/mol
Chlorine (Cl) = 1 X 35.45= 35.45 g/mol
Molar Mass of NaCl = (22.98 + 35.45) g/mol = 58.43 g/mol

After calculating the atomic mass of each element and its individual atom, the molecular mass is determined which comes out to be around 58.4 after rounding off to the nearest decimal.

Sodium Sulfate

Sodium sulfate is an ionic compound with the formula Na_2SO_4.

Answer and Explanation:

The molar mass of sodium sulfate is 142.04214 g/mol.

Determining the molar mass of sodium sulfate :

requires us to sum the molar masses for each element in the compound,

considering how many of each element are in the compound:

- Two sodium atoms:
  - (2)(22.989770 g/mol) = 45.97954 g/mol
- One sulfur atom:
  - (1)(32.065 g/mol) = 32.065 g/mol
- Four oxygen atoms:
  - (4)(15.9994 g/mol) = 63.9976 g/mol

Now we sum these values:

45.97954 g/mol + 32.065 g/mol + 63.9976 g/mol = 142.04214 g/mol

- The mass of 1 mole of a compound is called its *molar weight* or *molar mass*.
- The units for molar weight or molar mass are grams per mole.
- **Molecular weight = gram molecular weight**

  = 1 mol of compound

- Here is the formula to determine the number of moles of a sample:
- **No .of mol = weight of sample (g) /molar weight**

  Wt. in gm = No .of mol X molar weight

- Note that the ratios in which individual atoms combine to form molecules are exactly the same as the ratios in which moles of these atoms combine.
- Example : in carbon tetrachloride CCl_4

1 atom C + 4 atoms Cl → 1 molecule CCl_4

Thus 1 mol C + 4 mols Cl → 1 mol CCl_4

2 mol C + 8 mols Cl → 2 mol CCl_4

5 mol C + 20 mols Cl → 5 mol CCl_4

0. 1mol C + 0. 4mols Cl → 0.1 mol CCl_4

0.5 mol C + 2 mols Cl → 0.5 mol CCl_4

Mole to gram conversion

- Determine the number of grams in 4 moles of H_2O
- Formula mass H_2O = (2 x 1.0) + (1 x 16)

  = 18
- 1 mole H_2O = formula mass H_2O

  = 18 grams H_2O
- 4 moles H_2O x 18 grams / 1 mole

  = 72 grams H_2O

Problem

- The balanced equation for the synthesis of ammonia is

$$3\ H_2(g) + N_2(g) \dashrightarrow 2\ NH_3(g).$$

Calculate:

a. the mass in grams of NH_3 formed from the reaction of 64.0 g of N_2

b. the mass in grams of N_2 required to form 1.00 kg of NH_3

Solution

From the balanced equation, it is known that:

$$1 \text{ mol } N_2 \ \alpha\ 2 \text{ mol } NH_3$$

- Use the periodic table to look of the atomic weights of the elements to calculate the weights of the reactants and products:

1 mol of N_2 = 2(14.0 g) = 28.0 g

1 mol of NH_3 is 14.0 g + 3(1.0 g) = 17.0 g

- These relations can be combined to give the conversion factors needed to calculate the mass in grams of NH_3 formed from 64.0 g of N_2:
- mass NH_3 = 64.0 g N_2 x 1 mol N_2 /28.0 g NH_3 (wt/MW)
  x 2 mol NH_3 /1mol NH_3 x 17.0 g NH_3 /1 mol NH_3

  mass NH_3 = 77.7 g NH_3

- To obtain the answer to the second part of the problem, the same conversions are used, in a series of three steps:
- (1) grams NH_3 --> moles NH_3 (1 mol NH_3 = 17.0 g NH_3)
- (2) moles NH_3 --> moles N_2 (1 mol N_2 $\propto$ 2 mol NH_3)
- (3) moles N_2 --> grams N_2 (1 mol N_2 = 28.0 g N_2)
- mass N_2 = 1.00 x 103 g NH_3 x 1 mol NH_3 /17.0 g NH_3 x 1 mol N_2 /2 mol NH_3 x 28.0 g N_2 /1 mol N_2
- mass N_2 = 824 g N_2

- **Answer**
- a. mass NH_3 = 77.7 g NH_3
  b. mass N_2 = 824 g N_2

Gram to Mole conversion

- Determine the number of moles in 88 grams of CO_2
- Formula Mass CO_2 = (1 x 12) + (2 x 16)
  = 44
- 1 mole of CO_2 = formula mass CO_2
  = 44 grams CO_2
- 88 grams CO_2 x 1 mole CO_2 / 44 grams CO_2
  = 2 moles CO_2

- **Problem**
- Determine the number of moles of CO_2 in 454 grams.
- **Solution**
- First, look up the atomic masses for carbon and oxygen from the Periodic Table. The atomic mass of C is 12.01 and the atomic mass of O is 16.00.
- The formula mass of CO_2 is: $12.01 + 2(16.00) = 44.01$
- Thus, one mole of CO_2 weights 44.01 grams.
- This relation provides a conversion factor to go from grams to moles. Using the factor 1 mol/44.01 g:
- moles CO_2 = 454 g x 1 mol/44.01 g = 10.3 moles
- **Answer**
- 10.3 moles CO_2

Concentration Units

- There are numerous ways of expressing concentrations. It will be important to know the units used to express each concentration, as these units essentially define the concentration. Let's look at some ways to express concentration.
- **Standard Solution :**

is that solution of Known concentration (حفظ)

- **Parts-per" concentration**
- One common method of expressing the concentration is based on **the quantity of solute in a fixed quantity of solution.** The "quantities" referred to here can be expressed in weight, in volume, or both (the weight of solute in a given volume of solution.) In order to distinguish among these posibilities, the abbreviations (w/w), (v/v) and (w/v) are used.
- In commerce, medicine, and other applied fields of Chemistry, (w/w) measure is often used, and is commonly expressed as *weight-percent concentration*, or simply "percent concentration".
- For example, a solution containing 5 g of NaCl and 95 g of H_2O is a 5% solution of NaCl.

1. Weight/Weight Percent (w/w %)

– This unit of concentration is often used for concentrated solutions, typically acids and bases. If you were to look on a bottle of a concentrated acid or base solution the concentration expressed as a weigh/weight percent. A weight/weight percent is defined as

:

$$\text{w/w\%} = \frac{\text{grams of solute}}{\text{grams of solution}} \text{ x } 100$$

- <u>2. Molarity (M)</u> :This unit of concentration relates the moles of solute per liter of solution.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{L solution}}$$

- <u>3.Molality (m)</u>:
- This unit of concentration relates the moles of solute per kilogram of solvent

$$\text{Molality} = \frac{\text{moles of solute}}{\text{kg solvent}}$$

4 - Normality (N) :

is the number of equivalent weight of solute per liter.

$$\text{normality} = N = \frac{\text{number of equivalents of solute}}{\text{1 liter of solution}} = \frac{\text{equivalents}}{\text{liter}}$$

where

$$\text{number of equivalents of solute} = \frac{\text{grams of solute}}{\text{equivalent weight of solute}}$$

then

$$N = \frac{\text{grams of solute}}{\text{eq wt solute} \times \text{L solution}} = \frac{\text{grams}}{\text{eq wt} \times \text{L}}$$

$$N = \frac{\text{Wt.}}{\text{Eq. Wt.}} \times \frac{1000}{\text{V (ml)}}$$

$$\text{Equivalent weight} = \frac{\text{Molecular weight}}{\text{Valence}}$$

- **Valance** (n): number of hydrogen ions in acid or number of hydroxide ions in base
- or number of electrons transport in a reaction.

Consider the following reactions in which an excess of HCl is present. Hydrogen actually exits as H_2 molecules, but for convenience in considering the data, the hydrogen produced is shown as the number of atomic weights of hydrogen released per atomic weight of metal reacting.

$$Na_{(s)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H^{\circ}_{(g)}$$
$$Ca_{(s)} + 2\ HCl_{(aq)} \longrightarrow CaCl_{2(aq)} + 2\ H^{\circ}_{(g)}$$
$$Al_{(s)} + 3\ HCl_{(aq)} \longrightarrow AlCl_{3(aq)} + 3\ H^{\circ}_{(g)}$$

The table below summarizes the pertinent data for these reactions:

metal	atomic weight (amu)	number of atomic weights of hydrogen liberated per atomic weight of metal	equivalent weight of metal (amu)
Na	23.0	1	23.0/1 = 23.0
Ca	40.1	2	40.1/2 = 20.0
Al	27.0	3	27.0/3 = 9.0

$$\text{eq wt} = \frac{\text{at wt Na}}{1} = \frac{\text{at wt Ca}}{2} = \frac{\text{at wt Al}}{3} = \frac{\text{at wt H}}{1}$$

The equivalent weight of a substance may be variable; its value is dependent on the reaction that the substance is undergoing. Consider the following reactions:

$$NaOH + H_2SO_4 \longrightarrow NaHSO_4 + H_2O$$
$$2\,NaOH + H_2SO_4 \longrightarrow Na_2SO_4 + 2\,H_2O$$

In the first reaction, 1 mole of sulfuric acid furnishes 1 gram-atomic weight of hydrogen. Therefore the equivalent weight of sulfuric acid is the formula weight (98.1 grams). In the second reaction, the equivalent weight of sulfuric acid is ½ the formula weight (49.0 grams).

- Parts per million (ppm):
- This unit of concentration may be expressed in a number of ways. It is often used to express the concentration of very dilute solutions. The "technical" definition of parts per million is:

$$\text{ppm} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 10^6$$

$$\text{ppm} = \frac{\text{mg of solute}}{\text{L solution}}$$

- Parts per billion (ppb):

This concentration unit is also used for very dilute solutions. The "technical" definition is as follows:

$$\text{ppb} = \frac{\text{g of solute}}{\text{grams of solution}} \times 10^{9}$$

$$\text{ppb} = \frac{\mu\text{g of solute}}{\text{L solution}}$$

- **<u>Problem Example 1</u>**
- **Find the percent (w/w) concentration of a solution containing 4.5 g of sucrose in 90 mL of water.**
- ***<u>Solution</u>:*** **90 mL of water has a mass of 90 g, so the concentration will be**

4.5 g /94.5 g × 100% = 4.8% (w/w) sucrose

- ***<u>Percent</u>*** **<u>means "parts per 100";</u>**
- **we can also use <u>parts per thousand</u> (ppt) for expressing concentrations in grams of solute per kilogram of solution.**
- **For more dilute solutions, <u>parts per million (10^6, ppm) and parts per billion (10^9; ppb)</u> are used. These terms are widely employed to express the amounts**

- **<u>Problem Example 2</u>**
- **Describe how you would prepare 30 g of a 20 percent (w/w) solution of KCl in water.**
- ***<u>Solution:</u>* The weight of potassium chloride required is 20% of the total weight of the solution,**

  or 20/100 × 30 g = 6.0 g of KCl.
- **The remainder of the solution (30–6 = 24) g consists of water.**
- **Thus you would dissolve 6.0 g of KCl in 24 g of water.**

- **Example 3**
- **Determine the molarity of a solution made by dissolving 20.0 g of NaOH in sufficient water to yield a 482 cm^3 solution.**
- **Solution**

  Molarity is an expression of the moles of solute (NaOH) per liter of solution (water).

  1 mol NaOH weighs 23.0 g + 16.0 g + 1.0 g = 40.0 g

  So the number of moles in 20.0 g is:
- **moles NaOH = 20.0 g x 1 mol/40.0 g = 0.500 mol**
- **1 liter is 1000 cm^3, so the volume of solution is:**
- **liters solution = 482 cm^3 x 1 liter/1000 cm^3 = 0.482 liter**
- **molarity = 0.500 mol / 0.482 liter = 1.04 mol/liter**

  = 1.04 M

- Problem Example 4
- How would you make 120 mL of a 0.10 M solution of potassium hydroxide in water?
- ***Solution:***

M = moles of solute / liters of solution

or moles = M x L

(no.mol = V x M) (wt. in gm = no.mol x fm)

- The amount of KOH required is (0.120 L) × (0.10 mol L^{-1}) = 0.012 mol.

The molar mass of KOH is 56.1 g, so the weight of KOH required is .012 mol × 56.1 $gmol^{-1}$ = 0.67 g.

We would dissolve this weight of KOH in a volume of water that is *less than* 120 mL, and then add sufficient water to bring the total volume up to 120 mL.

(Remember: molarity is defined in terms of the volume of the *solution*, not of the *solvent*.)

Problem Example 5

- Calculate the molarity of a 60% (w/w) solution of ethanol (C_2H_5OH) in water whose density is 0.8937 g mL^{-1}.
- *Solution:* One liter of this solution has a mass of 893.7 g, of which (0.60 × 893.7 g) = 536.2 g consists of ethanol.
- The molecular weight of C_2H_5OH is 46.0, so the number of moles of ethanol present in one litre (that is, the molarity) will be (536.2 g) ÷ 46.0 g mol^{-1}) 1L
- = 11.6 mol L^{-1}
- Although molar concentration is widely employed, it suffers from one serious defect: since volumes are temperature-dependent (substances expand on heating), so are molarities;
- a 0.100 *M* solution at 0 ◦C will have another concentration at 50 ◦C. For this reason, molarity is not the preferred concentration measure in applications where physical properties of solutions and the effect of temperature on these properties is of importance.

Example:

- How many ml of 0.25M solution of NaOH are needed to provide a 0.02 mol NaOH ?
- 0.25 M NaOH means 0.25 mol NaOH / L solution

  0.25 mol NaOH / 1000 ml solution

  0.02 mol NaOH $\leftrightarrow$? Ml solution

  0.02 mol NaOH x 1000 ml sol. / 0.25 mol = 80 ml

Thus if we measure 80 ml of this solution it will contain the desired 0.02 mol of NaOH .

Example :

- how many gm of silver nitrate $AgNO_3$ are needed to prepare 500 ml 0.3 M $AgNO_3$ sol.?
- 0.3 M $AgNO_3$ means 0.3 mol $AgNO_3$ /1000 ml solution.
- In the final solution , the amount of $AgNO_3$ that must be present :
- 500 ml x 0.3 mol $AgNO_3$ / 1000 ml sol.=0.15 mol $AgNO_3$ (no.mol = V x M)
- The formula mass of $AgNO_3$ is 170 g/mol

  (wt. in gm = no.mol x fm)
- Therefore 0.15 mol $AgNO_3$ x 170 gm $AgNO_3$ /1 mol $AgNO_3$ = 25.5 gm $AgNO_3$.

Determine concentration of a solution in which 6.081 g $NaNO_3$ is dissolved to a total volume of 843 mL.

1. Calculate moles of solute

moles of $NaNO_3$

$6.081\ g / 84.994\ g\ mol^{-1} = 7.155 \times 10^{-2}$ mol

2. Calculate molarity of solution

a. convert volume to L

843 mL = 0.843 L

b. calc M

$7.155 \times 10^{-2}\ mol / 0.843\ L = \boxed{8.49 \times 10^{-2}\ M}$

Dilution

Preparation of diluted acid from concentrated acid by using **dilution low**

$$\mathbf{M1 \times V1 = M2 \times V2}$$

Concentrated acid   diluted acid

Prepare 100 ml. of diluted HCl (0.1M) from 5M concentrated HCl.

$$M1 \times V1 = M2 \times V2$$

Concentrated acid   diluted acid

$$5 \times V1 = 0.1 \times 100 \rightarrow\rightarrow V1 = \frac{0.1 \times 100}{5}$$

$$V1 = \frac{10}{5} \rightarrow\rightarrow V1 = 2ml.$$

We measure (2ml) of concentrated HCl and adding D.W until complete the final volume which is 100 ml.

Dilutions

50.0 mL of .650 M NaCl solution is diluted with 1000.0 mL of water. Determine concentration of final solution.

1. $C_1V_1 = (0.650\ M)(0.0500\ L) = 0.0325$ mol NaCl
2. final volume $= V_2 = 50.0\ mL + 1000.0\ mL$

   $= 1050.0\ mL$
3. $C_2V_2 = C_1V_1 = 0.0325$ mol NaCl

   $C_2 = C_1V_1 / V_2 = 0.0325\ mol / 1.0500\ L$

   $= 3.10 \times 10^{-2}\ M$

You prepare a Cu stock solution with a concentration of 209.5 ppm. You need to now prepare a standard solution with a concentration of 7.5 ppm in a 25 mL volumetric flask. How do you prepare the standard?

$C_1 = 209.5$ ppm

$C_2 = 7.5$ ppm

$V_1 = ?$

$V_2 = 25$ mL

$$V_1 = \frac{C_2 V_2}{C_1} = \frac{(7.5\ \text{ppm})(25\ \text{mL})}{(209.5\text{ppm})} = 0.895\ \text{mL}$$