Expressing concentration of solution

- **Molarity** : is the number of moles of solute dissolved in one liter of solution. The units, therefore are **moles per liter**, specifically it's **moles of solute** per **liter of solution**.

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liter of solution}}$$

Example 1. What is the molarity of a 5.00 liter solution that was made with 10.0 moles of KBr ?

Solution: We can use the original formula. Note that in this particular example, where the number of moles of solute is given, the identity of the solute (KBr) has nothing to do with solving the problem.

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

Given: # of moles of solute = 10.0 moles
Liters of solution = 5.00 liters

$$\text{Molarity} = \frac{\text{10.0 moles of KBr}}{\text{5.00 Liters of solution}} = 2.00\ M$$

Answer = 2.00 M

$$\text{Molarity} = \frac{\text{Weight (g)}}{\text{Molecular Weight (g/mol)}} \times \frac{1000}{\text{Volume (ml)}}$$

Molecular Weight = Sum. Of atomic weight

Example : Prepare 0.1 M of NaCl in 250 ml of D.Water from Solid?

Wt= M x M.wt. x V(ml) / 1000

= 0.1 x 55.5 x 250 / 1000

= 1.38 mol/L

- **Normality** : is the number of equivalents of solute dissolved in one liter of solution. The units, therefore are **equivalents per liter**, specifically it's **equivalents of solute** per **liter of solution.**

$$\text{Normality} = \frac{\text{No. of equivalents of solute}}{\text{liter of solution}}$$

$$\text{No. of equivalents} = \frac{\text{Weight (g)}}{\text{Equivalent Weight (g/eq)}}$$

$$\text{Normality} = \frac{\text{Weight(g)}}{\text{Equivalent weight (g/eq)}} \times \frac{1000}{\text{Volume(ml)}}$$

$$\text{Eq.Wt} = \frac{\text{M.Wt}}{n}$$

n = No. of (H) atoms for acids

for HCl n=1

n = No of OH groups for bases

for NaOH n=1

n = No of Cation atoms (M+) for salts

for Na_2CO_3 n= 2

n = No. of gained or lost electrons for oxidants and reductants

for $KMnO_4$ n= 7

- **Relationship between Molarity and Normality**

$$\text{Molarity} = \frac{\text{Weight (g)}}{\text{Molecular Weight (g/mol)}} \times \frac{1000}{\text{Volume (ml)}}$$

$$\text{Normality} = \frac{\text{Weight(g)}}{\text{Equivalent weight (g/eq)}} \times \frac{1000}{\text{Volume(ml)}}$$

$$\text{Eq.Wt} = \frac{\text{M.Wt}}{\text{N}}$$

$$N = M \times n$$

Q / what is the normality of 0.1 mol / l of Na_2SO_4 ?

N = 0.1 / 2 = 0.05 equivalents / L

- Weight – Volume Percentage (% w/v)

$$\% \, w / v = \frac{\text{Weight of solute (g)}}{\text{Volume of solution (ml)}} \times 100$$

- Weight – Weight Percentage (% w/w)

$$\% \, w / v = \frac{\text{Weight of solute (g)}}{\text{Weight of solution (g)}} \times 100$$

- Volume – Volume Percentage (% v / v)

$$\% \, w / v = \frac{\text{Volume of solute (ml)}}{\text{Volume of solution (ml)}} \times 100$$

Q/ What is the weight/volume percentage concentration of 250mL of aqueous sodium chloride solution containing 5g NaCl?

Calculate the weight/volume (%) = mass solute ÷ volume of solution x 100

mass solute (NaCl) = 5g

volume of solution = 250mL

w/v (%) = 5g ÷ 250mL x 100 = 2g/100mL (%)

Q / 2.0L of an aqueous solution of potassium chloride contains 45.0g of KCl. What is the weight/volume percentage concentration of this solution in g/100mL?

a. Convert the units (mass in grams, volume in mL):
   mass KCl = 45.0g
   volume of solution = 2.0L = 2.0 x 10^3mL = 2000mL
b. Calculate w/v (%) = mass solute (g) ÷ volume solution (mL) x 100
   w/v (%) = 45.0 ÷ 2000mL x 100 = 2.25g/100mL (%)

- **Mole Fraction**

The mole fraction, *X*, of a component in a solution is the ratio of the number of moles of that component to the total number of moles of all components in the solution.

To calculate mole fraction, we need to know:

- The number of moles of each component present in the solution.

The mole fraction of A, X_A, in a solution consisting of A, B, C, ... is calculated using the equation:

$$X_A = \frac{\text{moles of A}}{\text{moles of A} + \text{moles of B} + \text{moles of C} + \cdots}$$

To calculate the mole fraction of B, X_B, use:

$$X_B = \frac{\text{moles of B}}{\text{moles of A} + \text{moles of B} + \text{moles of C} + \cdots}$$

Molality

Molality, m, tells us the number of moles of solute dissolved in exactly one kilogram of solvent. (represented by a lower case m.)

We need two pieces of information to calculate the molality of a solute in a solution:

- The moles of solute present in the solution.
- The mass of solvent (in kilograms) in the solution.

To calculate molality we use the equation:

$$\text{Molality} = \frac{\text{moles of solute}}{\text{mass of solvent in kilograms}}$$

Q / If you have 10.0 grams of Br_2 and dissolve it in 1.00 L of cyclohexane, what is the molality of the solution? The density of cyclohexane is 0.779 kg/l at room temperature.

Solution /

First, work out the number of moles of bromine. Br_2 has a molecular weight of 159.8 g/mole, so we have

10 g / (159.8 g/mole) = 0.063 moles Br_2

Next, convert the volume of solvent to the weight of solvent using the density

1.0 L * 0.779 kg/l = 0.779 kg

Now just divide the two to get the molality

0.063 moles Br_2/ 0.779 kg cyclohexane = **0.080 molal**

- Parts per Millions (PPM)

$$\text{PPM} = \frac{\text{Weight of solute (g)}}{\text{Volume of Solution (ml)}} \times 10^{6}$$

Relationship between PPM and Molarity and Normality

$$PPM = M \times M.Wt \times 1000$$

$$PPM = N \times Eq.Wt \times 1000$$

Converting weight/volume (w/v) concentrations to ppm
ppm = $1g/m^3$ = 1mg/L = 1μg/mL

1. A solution has a concentration of 1.25g/L.
What is its concentration in ppm?

a. Convert the mass in grams to a mass in milligrams:
   1.25g = 1.25 x 1000mg = 1250mg
b. Re-write the concentration in mg/L = 1250mg/L = 1250ppm

2. A solution has a concentration of 0.5mg/mL.
What is its concentration in ppm?

a. Convert the volume to litres:
   volume = 1mL = 1mL ÷ 1000mL/L = 0.001L
b. Re-write the concentration in mg/L = 0.5mg/0.001L = 500mg/L = 500ppm

Converting weight/weight (w/w) concentrations to ppm
1ppm = 1mg/kg = 1μg/g

1. A solution has a concentration of 0.033g/kg.
What is its concentration in ppm?

a. Convert mass in grams to mass in milligrams:
   0.033g = 0.033g x 1000mg/g = 33mg

b. Re-write the concentration in mg/kg = 33mg/kg = 33ppm

2. A solution has a concentration of 2250μg/kg.
What is its concentration in ppm?

a. Convert mass in μg to mass in mg:
   2250μg = 2250μg ÷ 1000μg/mg = 2.25mg
b. Re-write the concentration in mg/kg = 2.25mg/kg = 2.25ppm

Parts Per Million (ppm) Concentration Calculations

1. 150mL of an aqueous sodium chloride solution contains 0.0045g NaCl.
Calculate the concentration of NaCl in parts per million (ppm).

a. ppm = mass solute (mg) ÷ volume solution (L)
b. mass NaCl = 0.0045g = 0.0045 x 1000mg = 4.5mg
   volume solution = 150mL = 150 ÷ 1000 = 0.150L
c. concentration of NaCl = 4.5mg ÷ 0.150L = 30mg/L = 30ppm

2. What mass in milligrams of potassium nitrate is present in 0.25kg of a 500ppm $KNO_{3(aq)}$?

a. ppm = mass solute (mg) ÷ mass solution (kg)
b. Re-arrange this equation to find the mass of solute:
   mass solute (mg) = ppm x mass solution (kg)
c. Substitute in the values:
   mass KNO_3 = 500ppm x 0.25kg = 125mg

3. A student is provided with 500mL of 600ppm solution of sucrose.
What volume of this solution in millilitres contains 0.15g of sucrose?

a. ppm = mass solute (mg) ÷ volume solution (L)
b. Re-arrange this equation to find volume of solution:
   volume solution (L) = mass solute (mg) ÷ ppm
c. Substitute in the values:
   volume solution (L) = (0.15g x 1000mg/g) ÷ 600 = 0.25L
d. Convert litres to millilitres: volume solution = 0.25L x 1000mL/L = 250mL

- **DILUTIONS**

Whenever you need to go from a more concentrated solution ["stock"] to a less concentrated one, you add solvent [usually water] to "dilute" the solution. No matter what the units of concentration are, you can always use this one formula

$C_1 V_1 = C_2 V_2$

[Concentration of the stock] x [Volume of the stock] = [Concentration of the final solution] x Volume of the final solution]

$N_1 V_1 = N_2 V_2$

$M_1 V_1 = M_2 V_2$

Q / What is the volume of 0.2 mol / L of NaOH that it required to dilute it to 0.05 mol /L in 100 ml ?

$N_1 V_1 = N_2 V_2$

$0.2 \times V_1 = 0.05 \times 100 \Longrightarrow V_1 = 25$ ml complete to 100 ml

- **Normality of Concentrated Reagents**

$$\text{Normality} = \frac{\text{Specific Gravity (g/l)} \times \text{Percentage (\%)} \times 1000}{\text{Equivalent Weight (g/ eq)}}$$

$$\text{Molarity} = \frac{\text{Specific Gravity (g/l)} \times \text{Percentage (\%)} \times 1000}{\text{Molecular Weight (g/ mol)}}$$

Q / Describe the preparation of 900 mL of 3.00 M $HN0_3$ from the commercial reagent that is 70.5% HN0 3 (w/w) and has a specific gravity of 1.42.

$$\text{Molarity} = \frac{\text{Specific Gravity (g/l) x Percentage (\%) x 1000}}{\text{Molecular Weight (g/ mol)}}$$

$$M_{HNO3} = \frac{1.42 \times (70.5/100) \times 1000}{63} = 15.9$$

$M_1 V_1 = M_2 V_2$

$15.9 \times V1 = 3 \times 900 \Longrightarrow V_1 = 159$ ml diluted to 900 ml

p-Functions

The p-function of a number *X* is written as p*X* and is defined as

$pX = -\log(X)$

$X = H^+, Cl^-$, …….etc.

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

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

$[H^+] + [OH^-] = 10^{-14} = K_w$

$PH + POH = 14$

What is pNa for a solution of 1.76×10^{-3} M Na_3PO_4?

SOLUTION

Since each mole of Na_3PO_4 contains three moles of Na^+, the concentration of Na^+ is

$$[Na^+] = \frac{3 \text{ mol } Na^+}{\text{mol } Na_3PO_4} \times 1.76 \times 10^{-3} \text{ M} = 5.28 \times 10^{-3} \text{ M}$$

and pNa is

$$pNa = -\log[Na^+] = -\log(5.28 \times 10^{-3}) = 2.277$$

Example /

What is the [H+] in a solution that has a pH of 5.16?

SOLUTION

The concentration of H+ is

$pH = -\log[H^+] = 5.16$

$\log[H^+] = -5.16$

$[H^+] = \text{antilog}(-5.16) = 10^{-5.16} = 6.9 \times 10^{-6}$ M