CHAPTER-4

Solutions



Cont'.....

- **1. Introduction**

- A **solution** is a *homogeneous mixture* of two or more substances of molecular dimensions, whose composition can be varied within certain limits.
- A solution has **two components** namely, ***solute*** and **solvent**.
- ✓ A **solute** is:
- ✓ a substance that dissolves.
- ✓ a substance *existing in small proportion*.
- ✓ the most active component chemically.
- A **solvent** is:
- A substance that ***does*** the ***dissolving process***.
- A substance *existing in larger proportion*.

Cont'.....

- ✓ A substance whose physical state doesn't change during the solution process.
- ✓ The least active component chemically.
- ❖ **NB: Solutes** and **solvents** may be of any form of matter: ***solid***, ***liquid*** **or** ***gas***.

❖ 4.2. Types of solutions

❖ There are ***three types*** of ***solutions***:

i. **Gaseous solution:** If the ***solvent is a gas***, only gases are dissolved under any given set of conditions.

ii. **Liquid solution:** If the ***solvent is a liquid***, then gases, liquids and solids can be dissolved.

iii. **Solid solution:** If the ***solvent is a solid***, then gases, liquids, and solids can be dissolved.

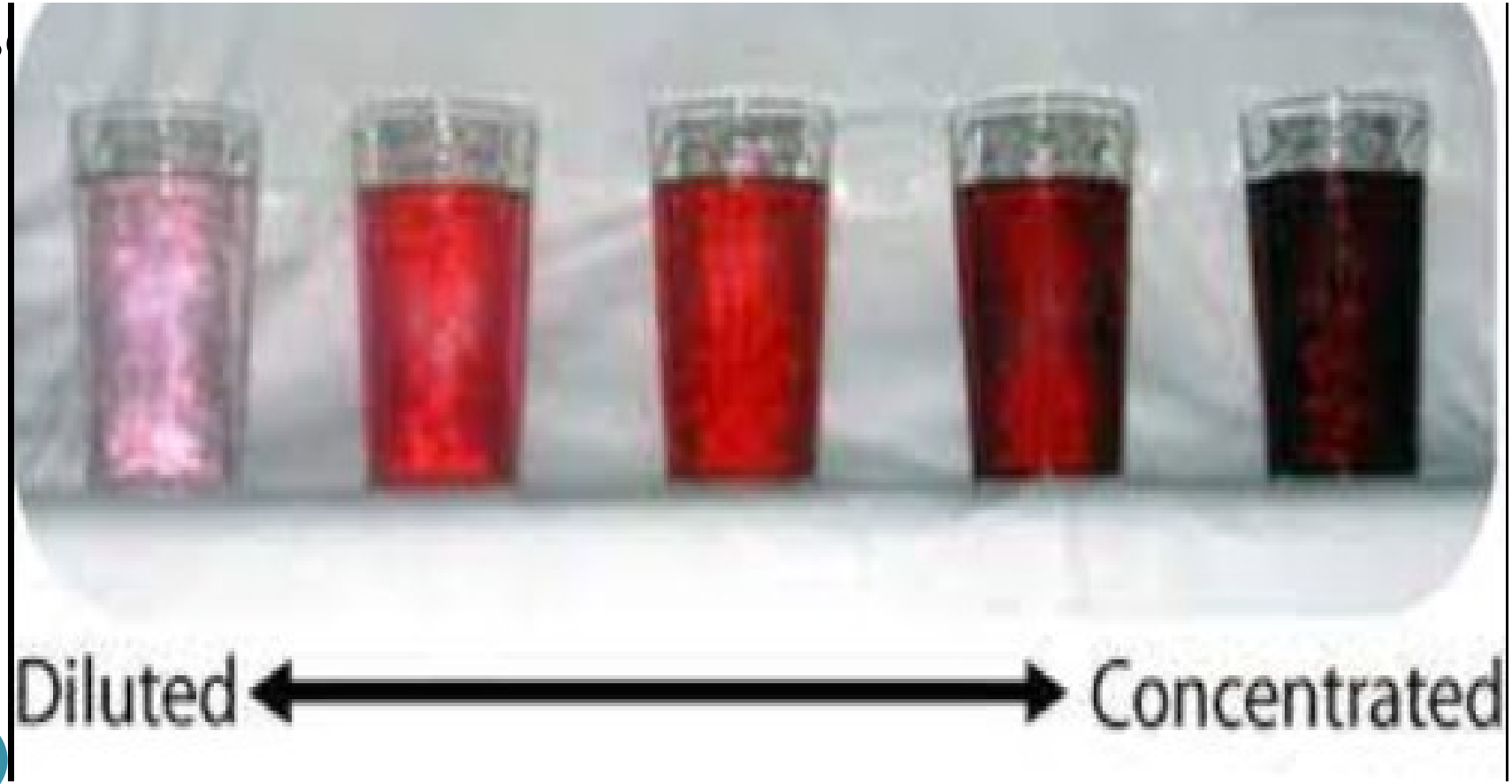
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**Some common examples for each [*types of*

Solute	Solvent	Types of solutions	Examples
Solid	Solid	Solid in solid	Alloys
Liquid	Solid	Liquid in solid	Hydrated salts
Gas	Solid	Gas in solid	Dissolved gas in minerals
Solid	Liquid	Solid in liquid	Salt solution in water
Liquid	Liquid	Liquid in liquid	Alcohol in water
Gas	Liquid	Gas in liquid	Aerated drinks
Solid	Gas	Solid in gas	Iodine vapor in air
Liquid	Gas	Liquid in gas	Humidity in air
Gas	Gas	Gas in gas	Air

Cont'.....

- **Measures of concentration**
- The concentration is a **measure of the amount of solute** that has been dissolved in a given amount of solvent or s



Emphasizing the amounts of solute in solution

- The amount of **solute** that has been dissolved in a given amount of **solvent** or solution can be expressed either **qualitatively** (diluted and concentrated) and **quantitatively**.
- A solution with a relatively **small amount of solute** is said to be **dilute**
- One with a **large amount** of the solutes is said to be **concentrated**
- Several **quantitative** expressions of solution concentration are used in chemistry including:
  - Percent by mass

❖ Percent by mass

✓ The percent by mass is the **ratio of the mass of a solute** to the mass of the solution, multiplied by 100.

$$\text{Percent by mass of solute} = \frac{\text{Mass of solute}}{\text{Mass of solute} + \text{Mass of solvent}} \times 100$$

or

$$\text{Percent by mass of solute} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$$

Example: What is the mass percentage of Fe in a piece of metal with 87.9 g of Fe in a 113 g sample?

$$\%m/m = \frac{87.9\,g\,Fe}{113\,g\,sample} \times 100\% = 77.8\%\,Fe$$

Cont'.....

- **Parts per million (ppm):** When the mass of **solute i**n the solution is very small, a widely used expression is

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

or

$$\text{ppm} = \text{Mass percentage of solute} \times 10^4$$

Example: A 2.5 g sample of ground water was found to cor
par

Solution:

Because 1 μg is 1×10^{-6} g, 5.4 μg = 5.4×10^{-6} g

$$\text{Thus, ppm} = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^6 = \frac{5.4 \times 10^{-6}\,\text{g}}{2.5\,\text{g}} \times 10^6 = 2.16\ \text{ppm}$$

Cont'.....

- For solutions that are even more dilute, parts per billion (ppb) is used.
- A conc g of solute per billion (

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

❖ Molarity (M)

- ✓ Molarity or m mber of **moles of solute** in 1 l

$$\text{Molarity (M)} = \frac{\text{Number of moles of solute}}{\text{Volume (L) of solution}}$$

$$\text{Number of moles of solute} = \frac{\text{Mass of solute}}{\text{Molar mass of solute}}$$

Example: What is the molar concentration of a solution containing 16.0 g of CH_3OH in 200 mL of

Solution:

Given:

Mass of solute (CH_3OH) = 16 g

Volume of solution = 200 mL = 0.20 L

Molar mass of CH_3OH = 32 g/mol

Required:

Molarity = ?

$$\text{Molarity} = \frac{\text{Mass of } CH_3OH}{\text{Molar mass of } CH_3OH \times \text{Volume (L) of Solution}}$$

$$= \frac{16\ \text{g}}{32\ \text{g/mol} \times 0.20\ \text{L}} = 2.50\ \text{mol/L} = 2.50\ \text{M}$$

❖ Molality (m)

- The molality of a solution is the number of moles of solute **per kilogram** of solvent con

$$\text{Molality (m)} = \frac{\text{Number of moles of solute}}{\text{Mass (kg) of solvent}}$$

Example: What is molality of a solution prepared by dissolving 32.0 g of $CaCl_2$ in 271 g of water?

Solution:

Given:

Mass of $CaCl_2 = 32.0$ g

Mass of $H_2O = 271$ g $= 0.271$ kg

Required:

Molality (m) =?

$$\text{Molality } (m) = \frac{\text{Mass of } CaCl_2}{\text{Molar mass of } CaCl_2 \times \text{Mass (kg) of } H_2O}$$

$$= \frac{32.0 \text{ g}}{111 \text{ g/mol} \times 0.271 \text{ kg}} = 1.06\ m\ CaCl_2$$

Cont'.....

❖ Mole fractions (X)

$$X_A = \frac{\text{moles of substance A}}{\text{total moles in solution (A+B+···)}}$$

Example: What is the **mole fraction of I_2** in a solution containing **30 g of I_2 in 400 g of CCl_4** ?

Solution:

First find the number of moles of I_2 and the number of moles of CCl_4

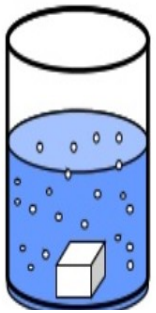
$$\text{Number of moles of } I_2 = \frac{\text{Mass of } I_2}{\text{Molar mass of } I_2} = \frac{30\text{ g}}{254\text{ g/mol}} = 0.12\text{ mol } I_2$$

$$\text{Number of moles of } CCl_4 = \frac{\text{Mass of } CCl_4}{\text{Molar mass of } CCl_4} = \frac{400\text{ g}}{254\text{ g/mol}} = 2.6\text{ mol } CCl_4$$

$$X_{I_2} = \frac{n_{I_2}}{n_{I_2} + n_{CCl_4}} = \frac{0.12\text{ mol}}{0.12\text{ mol} + 2.6\text{ mol}} = 0.044$$

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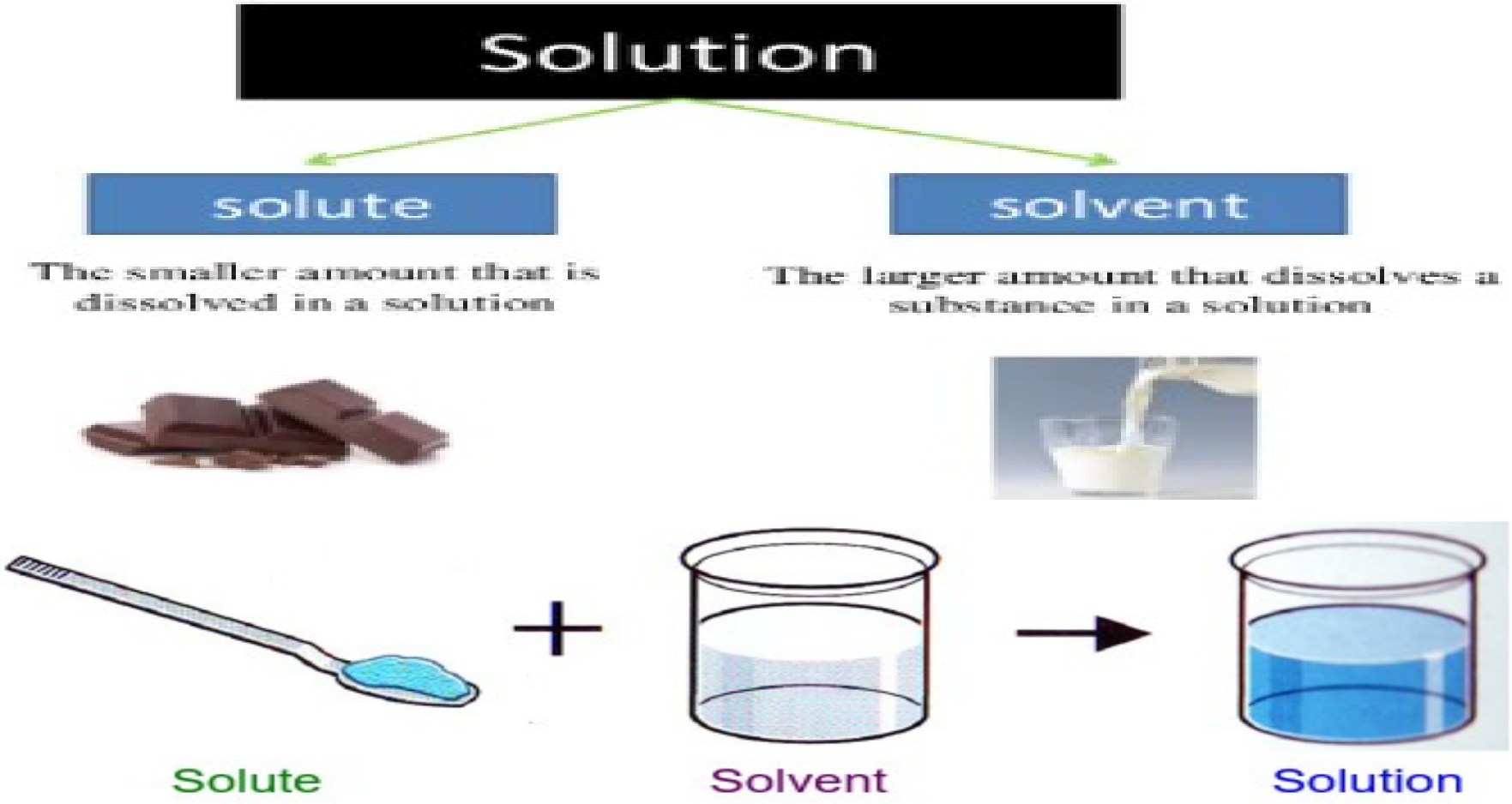
- **Emphasizing relative amounts of solute and solvent molecules**
- ✓ **Solubility**
  - Solubility of the substance is the amount of that substance that dissolve in a given amount of solvent.
  - Solubility, may be considered to be an equilibrium; equilibrium between a **solid** and **ions** in the solutions.
  - Any ionic solution is 100% ionized, ir it actually it dissolves:



$$Na^+(aq) + Cl^-(aq) \rightleftarrows NaCl(s)$$

  - Is the **maximum** quantity of solute that can be dissolved in certain quantity of solvent

Cont'.....

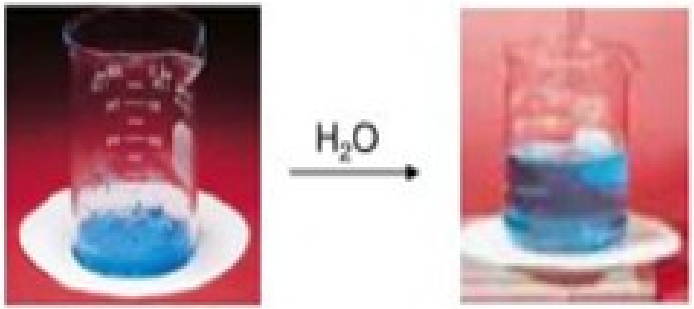


A **solute** is the substance to be dissolved (sugar).

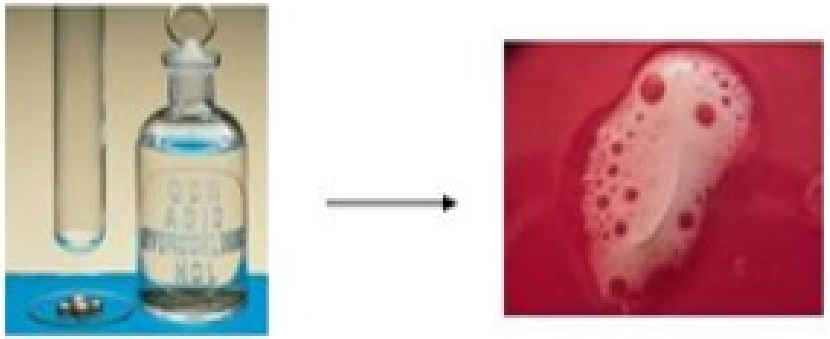
The **solvent** is the one doing the dissolving (water).

Cont'.....

❖ Dissolution Vs. Reaction



$$Cu(NO_3)_{2\,(s)} \xrightarrow{H_2O} Cu(NO_3)_{2\,(aq)}$$



$$Mg_{\,(s)} + 2\,HCl_{\,(aq)} \rightarrow H_{2\,(g)} + MgCl_{2\,(aq)}$$

Just because a substance disappears when it comes in contact with a solvent, it doesn't mean the substance dissolved.

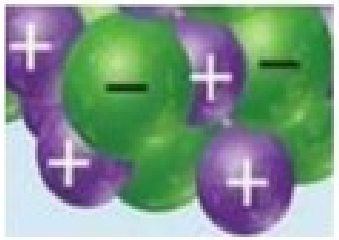
- Dissolution is a physical change—you can get back the original solute by evaporating the solvent.
- If you can't, the substance didn't dissolve, it reacted.

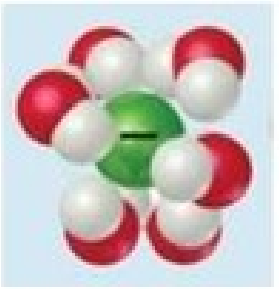
How does a solution form?

The intermolecular forces between solute and solvent particles must be strong enough to compete with those between solute particles and those between solvent particles.

Example: {NaCl + H_2O^*}

$$NaCl_{(s)} \xrightarrow[\text{dissociation}]{+ H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$$



As a solution forms, the solvent pulls solute particles apart and surrounds, or solvates, them.

- Simply put, three processes affect the energetics of the process:
  - Separation of solute particles
  - Separation of solvent particles
  - New interactions between solute and solvent



ΔH_1: Separation of solute molecules



ΔH_2: Separation of solvent molecules



ΔH_3: Formation of solute-solvent interactions

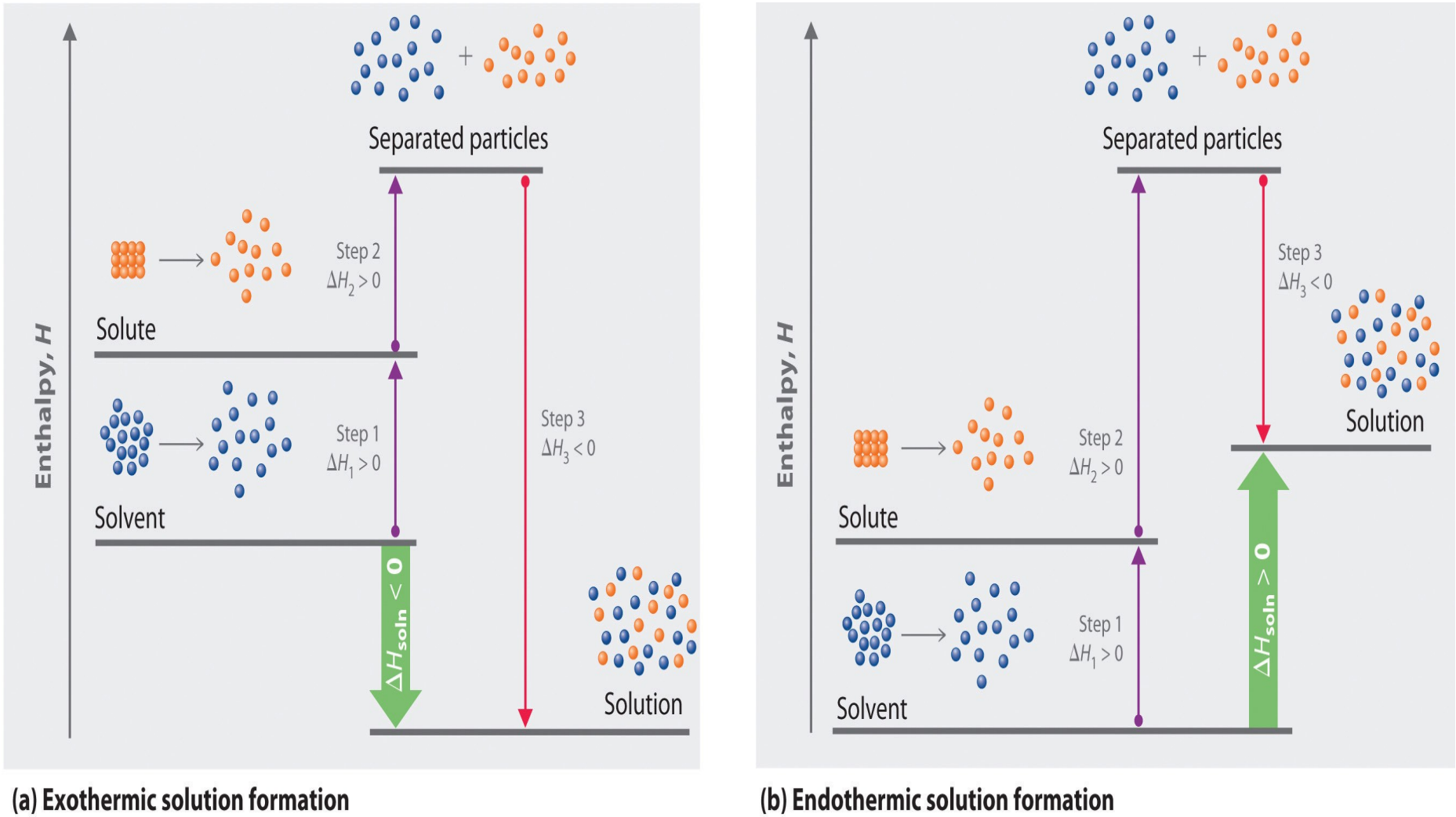


Figure -1 Enthalpy Changes That Accompany the Formation of a Solution.

Cont'.....

- ***Solvation*** *can be an exothermic or* ***endothermic*** *process depending on the nature of the solute and solvent*

- *From the* ***above figure****, In both cases,* ***step 1,*** *separation of the solvent particles, is energetically uphill (*$\Delta H_1 > 0$*), as is* ***step 2,*** *separation of the solute particles (*$\Delta H_2 > 0$*).*

- *In contrast, energy is released in step 3 (*$\Delta H_3 < 0$*) because of interactions between the solute and solvent.*

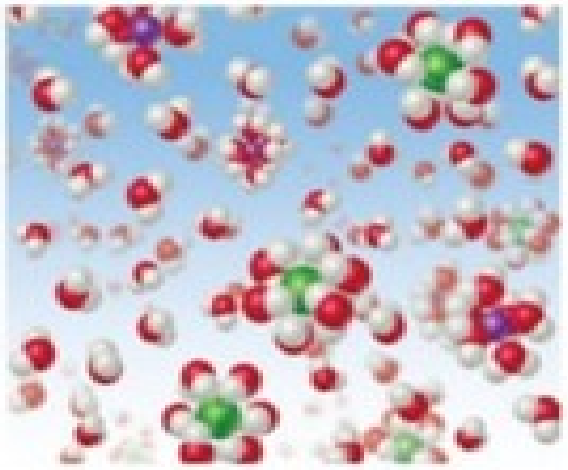
- *Figure, (a) When* ΔH_3 *is* ***larger in magnitude*** *than the sum of* ΔH_1 *and* ΔH_2*, the overall process is* ***exothermic*** ***(***$\Delta H_{soln} < 0$***),*** *as shown in the thermochemical cycle.*

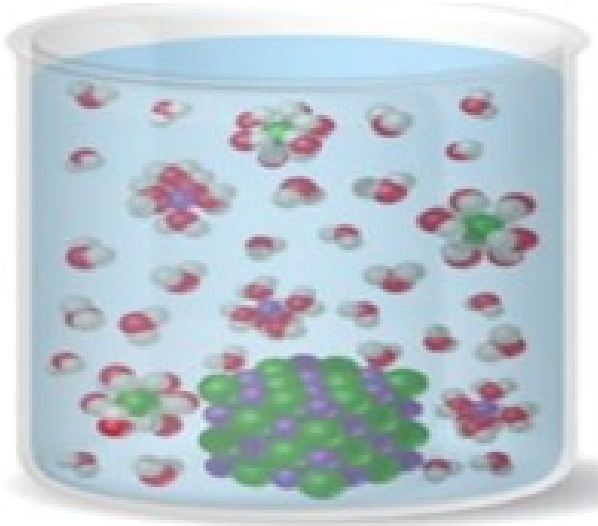
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- ***Figure (b)*** *When* ΔH_3 *is* ***smaller in magnitude*** *than the sum of* ΔH_1 *and* ΔH_2*, the overall process is* ***endothermic*** *(ΔHsoln > 0).*
- A positive value for ***ΔH soln*** does **not mean that a solution will not form.**
- A high ***ΔH soln*** is usually an indication that the substance is **not very soluble**

❖ Saturation and solubility

- Unsaturated
  - ➢ Less than the maximum amount of solute for that temperature is dissolved in the solvent.





- Saturated
  - ➢ Solvent holds as much solute as is possible at that temperature.
  - ➢ Dissolved solute is in dynamic equilibrium with solid solute particles.

- A solution is considered saturated when adding additional solute no longer increases the concentration
  - **Supersaturated**
    - Solvent holds more solute than is normally possible at that temperature.
    - These solutions are unstable; crystallization can usually be stimulated by adding a "seed crystal" or scratching the side of the flask.

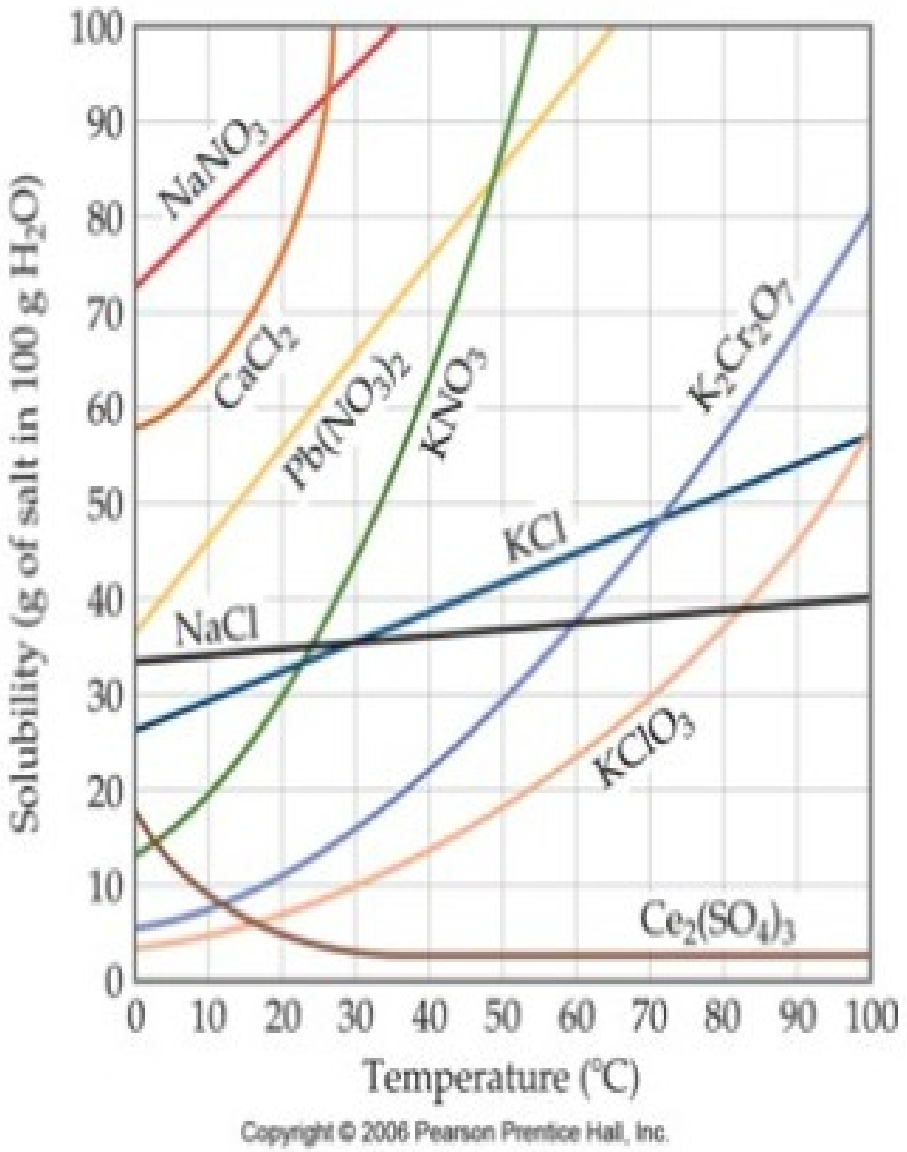
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- **Factors that affect solubility**
- Solubility of solute in a solvent depends on:

1) Temperature:

For solids &liquids: as temperature increases, solubility increases usually

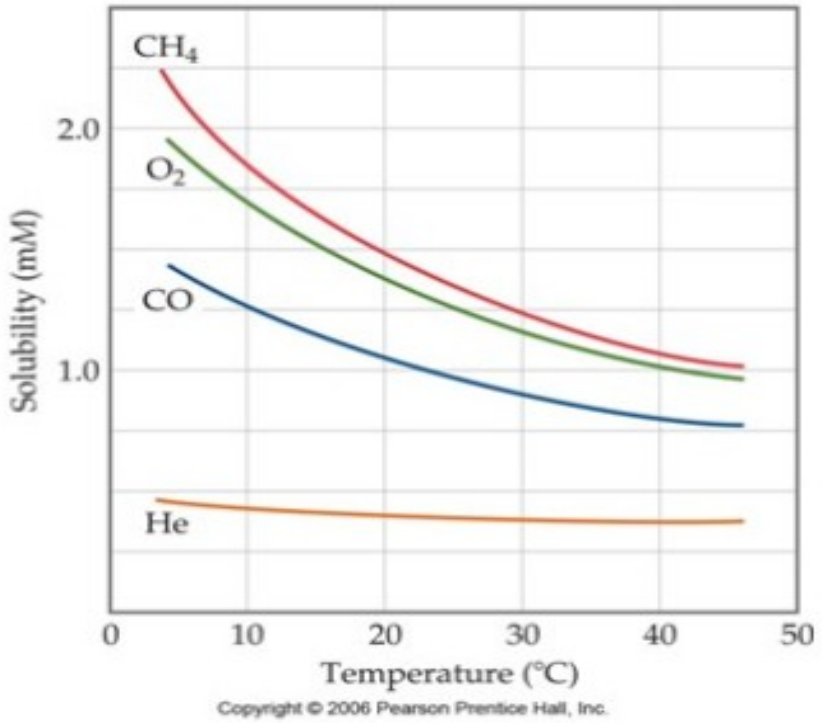
- **Many salts** show a large increase in solubility with temperature.
- Some solutes exhibit solubility that is fairly independent of temperature.
- A few, such as cerium(III) sulfate, become less soluble in water as temperature increases. ***Why?***



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- Usually, solubility of ionic solids increases with temperature

- **For gases**: Solubility of a gas in **water** tends to decrease with increasing temperature while solubility of a gas in an **organic solvent** tends to increase with increasing t



- Solubility of gases typically decreases with increasing temperature

All solubilities were measured with a constant pressure of

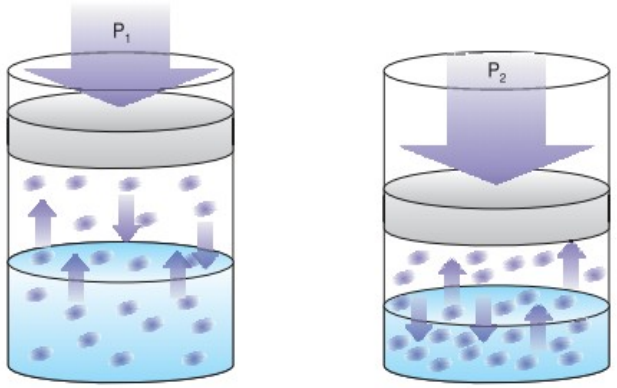
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2) Pressure

- **William Henry**, an English chemist, showed that the solubility of a gas increased with **increasing pressure.**
- The increase in solubility based on pressure will depend on which gas is being dissolved and must be determined experimentally for each gas
- **Henry's law**: States that the solubility of a gas in a liquid is directly proportional to the **partial pressure** of the gas above the liquid.
- He discovered the following relationship: $C = k \times P_{gas}$

➢ *Where* C is the concentration of the **gas in solution**, which is a measure of its **solubility**, k is a proportionality constant that has been experimentally determined, and P_{gas} is the partial pressure of the gas

- If you push down the piston (increasing pressure to P_2) more gas dissolves



Example: The partial pressure of carbon dioxide gas inside a bottle of mineral water is 4 atm at 25°C. What is the solubility of CO_2?.

Solution:

Given:

- P_{CO2} = 4 atm
- k= 3.3 x10^{-2} mol/L.atm

Required:

C_{CO2} = ?

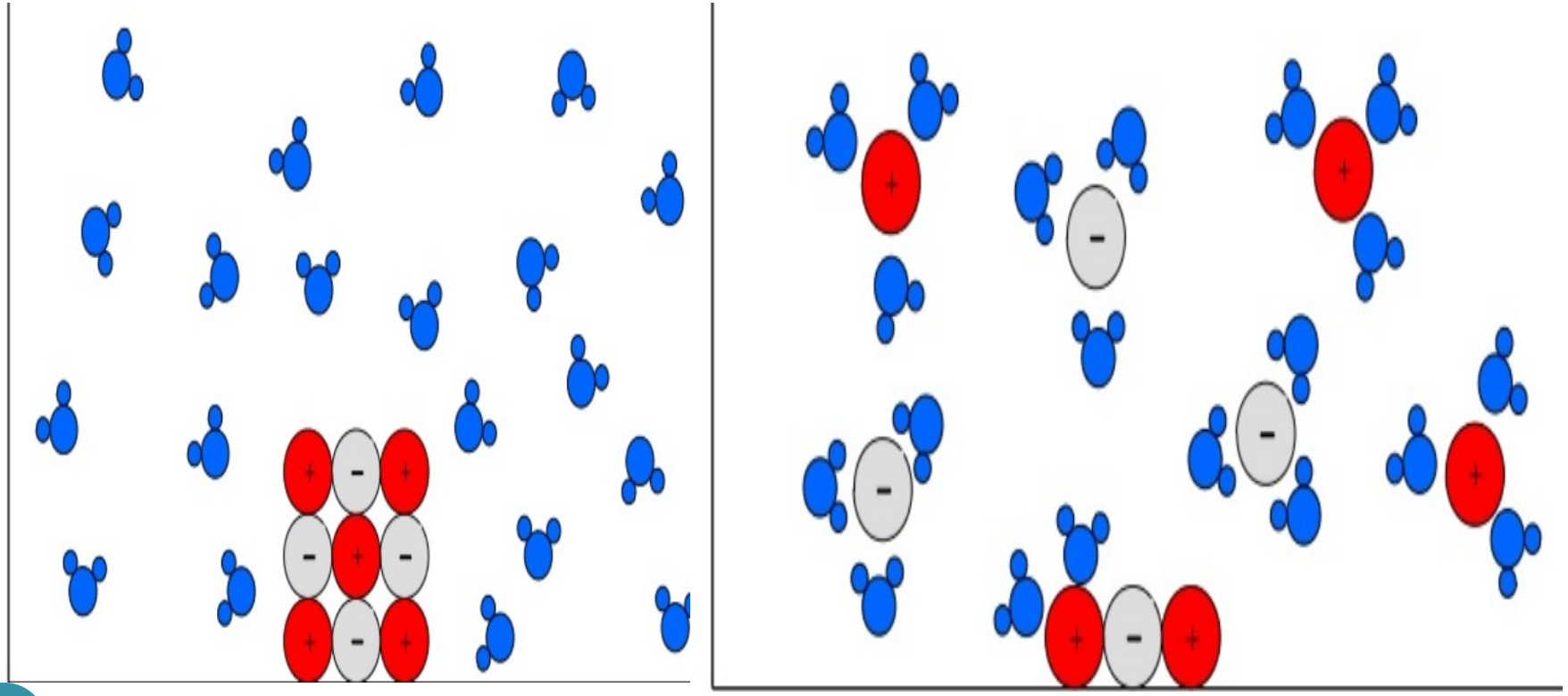
$C_{CO2} = k \cdot P_{CO2} = 3.3 \times 10^{-2}$ mol/L.atm . 4 atm = **0.1mol/L**

3) The nature of solute and solvent(Polarity)

- The phrase,'' **Like dissolve like''** is used to summarize how well a solute will dissolve in a solvent.
- ***Polar solvents generally dissolve polar solute very well and***
- ***Non-polar solvents generally dissolve non-polar solute very well.***
- How ever, polar solvents do not dissolve non-polar solute very well and non-polar solvents do not dissolve polar solute very well.

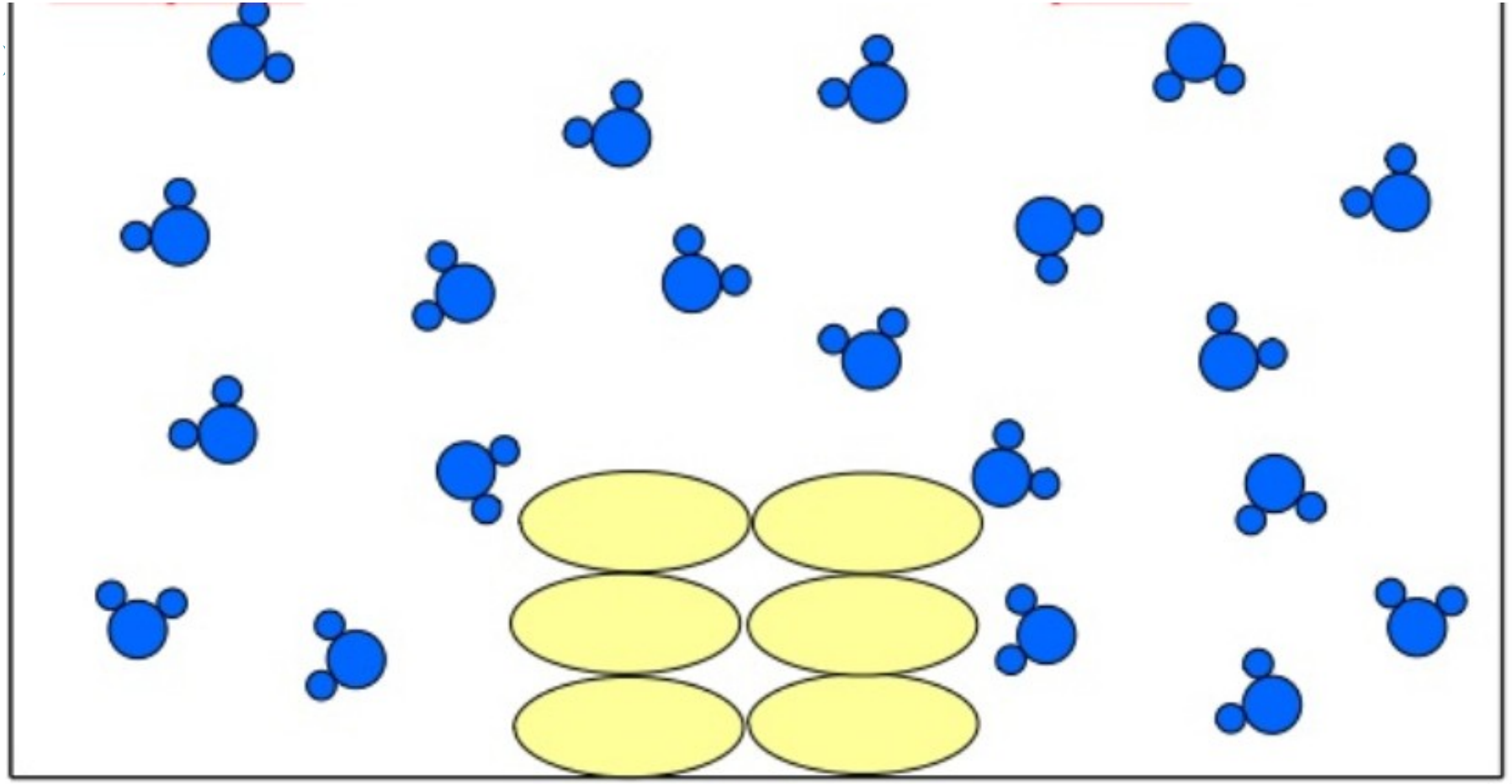
Example-1: Dissolving a salt(ionic solute) in water.

- **Like dissolve Like**
- All ionic solids are classified as polar solutes.



Cont'.....

Example-2: Trying to dissolve Wax(non-polar solute) in water.



Cont'.....

- **Preparation of solutions**

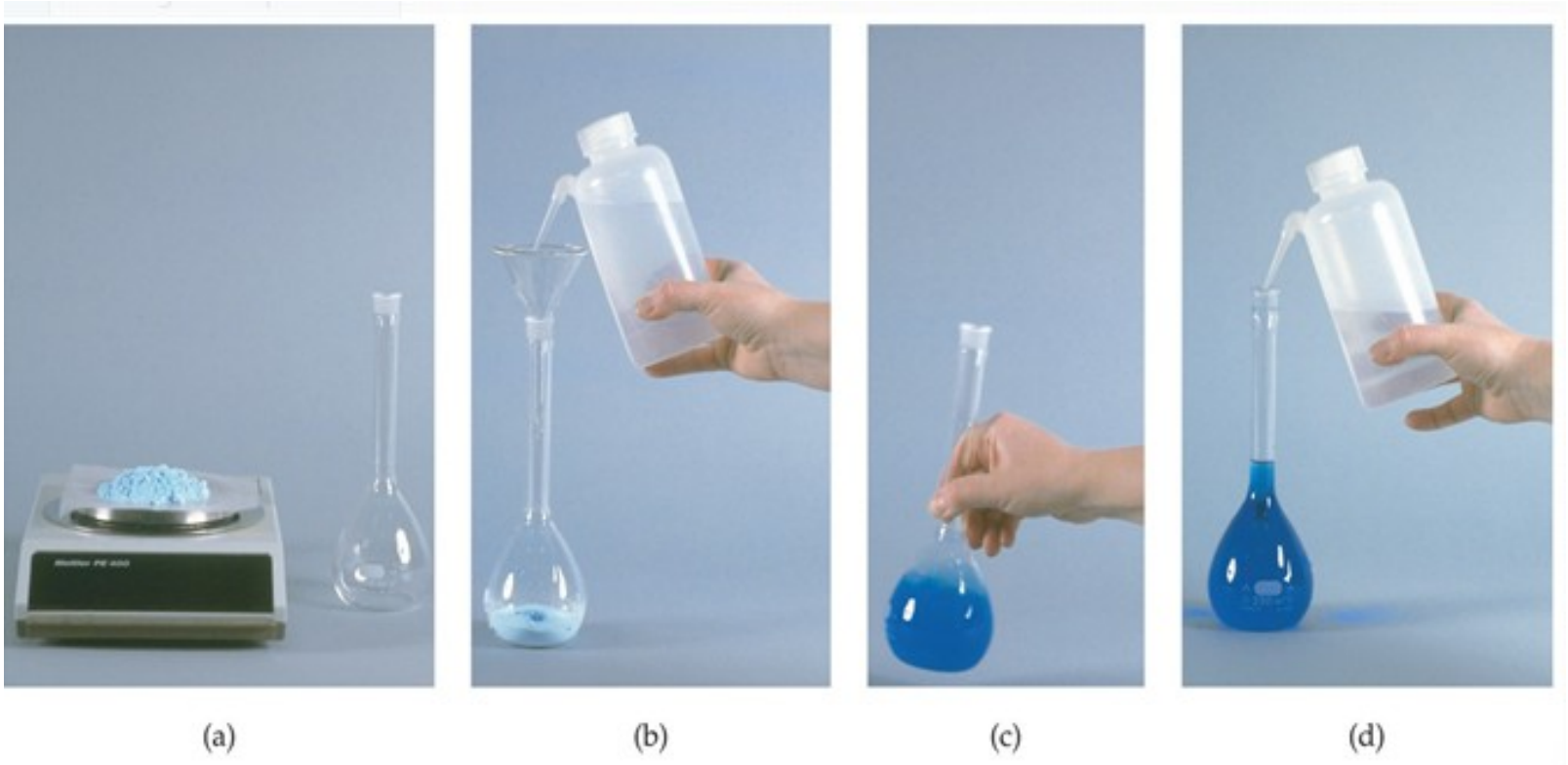
❖ **Preparing Stock Solutions**

- A ***stock solution*** is prepared by weighing out an appropriate portion of a *pure solid* or by measuring out an appropriate volume of a *pure liquid* and diluting to a known volume.
- Exactly how this is done *depends on* the required *concentration units*.

- **If starting with a solid, use the following procedure:**
  - Calculate the amount of solid required
  - Weigh out the solid
  - Place in an appropriate volumetric flask
  - Fill flask about half full with water and mix (shake).
  - Fill to the mark with water and invert to mix.

Cont'.....

- **Grams of solute to measure = V_D X M_D X M solute**
- Where, V_D-desired volume, M_D - molarity M_{solute}-molecular weight of solute



Cont'.....

Describe how you would prepare the solutions:

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

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:

Given: V_D = 500 ml =0.5L

Molarity = 0.2 M = 0.2 mol/L

Molar mass = 40g/mol

- Grams of solute to measure = 0.5 L x 0.2 mol/L x 40g/mol = 4gm.

Therefore; to prepare the solution we place 4.0 g of NaOH in a bottle or beaker and add approximately

Preparing Solutions by Dilution

$$C_B \times V_B = C_A \times V_A \quad \text{.....Dilution Law}$$

Where :

- ✓ C_B is the concentration of the stock solution,
- ✓ V_B is the volume of the stock solution being diluted (before dilution),
- ✓ C_A is the concentration of the dilute solution, and
- ✓ V_A is the volume of the solution after dilution.

NB: The total amount of solute is the **same** before and after dilution

Cont'.....

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).

Given : C_B = 14.8 M, C_A = 0.1 M, V_A= 250 ml, required V_B =?

Solution:

$$V_B = \frac{0.10M\ X\ 0.25\ L}{14.8\ M} = 1.69X10^{-3}L = 1.7\ mL$$

- Substituting known volumes in above equation, 14.8 M x V_B = 0.10 M x 0.25 L

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