TYPES OF SOLUTIONS

- A ***solution*** is a ***homogeneous*** mixture of two substances: a ***solute*** and a ***solvent***.
- ***Solute***: substance being dissolved; present in lesser amount.
- ***Solvent***: substance doing the dissolving; present in ***larger*** amount.
- ***Solutes*** and ***solvents*** may be of any form of matter: ***solid, liquid or gas***.

Some Examples of Solutions

Type	Example	Solute	Solvent
Gas in gas	Air	Oxygen (gas)	Nitrogen (gas)
Gas in liquid	Soda water	CO_2 (gas)	Water (liquid)
Liquid in liquid	Vinegar	Acetic acid (liquid)	Water (liquid)
Solid in liquid	Seawater	Salt (solid)	Water (liquid)
Liquid in solid	Dental amalgam	Mercury (liquid)	Silver (solid)
Solid in solid	Brass	Zinc (solid)	Copper (solid)

- Solutions form between solute and solvent molecules because of similarities between them. (***Like dissolves Like***)
- ***Ionic*** solids dissolve in ***water*** because the ***charged ions (polar)*** are attracted to the ***polar*** water molecules.
- ***Non-polar*** molecules such as oil and grease dissolve in ***non-polar solvents*** such as kerosene.

ELECTROLYTES & NON-ELECTROLYTES

- Solutions can be characterized by their ability to conduct an electric current. Solutions containing ***ions*** are ***conductors*** of electricity and those that contain ***molecules*** are ***non-conductors***.

- Substances that dissolve in water to form ions are called ***electrolytes***. The ions formed from these substances conduct electric current in solution, and can be tested using a conductivity apparatus (diagram below).

- Electrolytes are further classified as ***strong electrolytes*** and ***weak electrolytes***.

- In water, a ***strong electrolyte*** exists ***only as ions***. Strong electrolytes make the light bulb on the conductivity apparatus glow brightly. ***Ionic substances*** such as NaCl are ***strong electrolytes***.

$$NaCl\,(s) \xrightarrow{H_2O} Na^{+}\,(aq) + Cl^{-}\,(aq)$$

(only ions present after solution)

- Solutions containing ***weak electrolytes*** contain only a ***few ions***. These solutions make the light bulb on the conductivity apparatus glow dimly. Weak acids and bases that dissolve in water and produce few ions are ***weak electrolytes.***

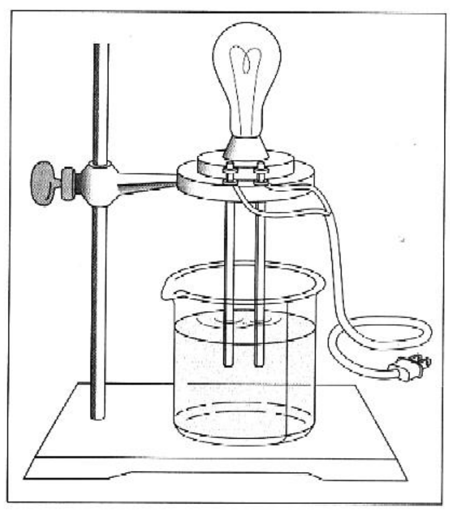
$$HF\,(aq) \underset{\text{Recombination}}{\overset{\text{Dissociation}}{\rightleftarrows}} H^{+}\,(aq) + F^{-}\,(aq)$$

(few ions present after solution)

- Substances that ***do not form any ions*** in solution are called ***non-electrolytes***. With these solutions the bulb on the conductivity apparatus does not glow. ***Covalent molecules*** that dissolve in water but do not form ions, such as sugar, are ***non-electrolytes***.

$$C_{12}H_{22}O_{11}\,(s) \xrightarrow{H_2O} C_{12}H_{22}O_{11}\,(aq)$$

(no ions present after solution)



ELECTROLYTES & NON-ELECTROLYTES

Classification of Solutes in Aqueous Solutions

Type of Solute	Dissociation	Particles in Solution	Conducts Electricity?	Examples
Strong electrolyte	Complete	Ions only	Yes	Ionic compounds such as NaCl, KBr, $MgCl_2$, $NaNO_3$; NaOH, KOH; HCl, HBr, HI, HNO_3, $HClO_4$, H_2SO_4
Weak electrolyte	Partial	Mostly molecules and a few ions	Yes, but poorly	HF, H_2O, NH_3, $HC_2H_3O_2$ (acetic acid)
Nonelectrolyte	None	Molecules only	No	Carbon compounds such as CH_3OH (methanol), C_2H_5OH (ethanol), $C_{12}H_{22}O_{11}$ (sucrose), CH_4N_2O (urea)

Examples:

1. Identify the predominant particles in each of the following solutions and write the equation for the formation of the solution:

   a) NH_4Br

   b) Urea (CH_4N_2O)

   c) HClO (weak acid)

EQUIVALENTS OF ELECTROLYTES

- Body fluids typically contain a mixture of several electrolytes, such as Na^+, Cl^-, K^+ and Ca^{2+}.

- Each individual ion is measured in terms of an equivalent (Eq), which is the amount of that ion equal to 1 mole of positive or negative electrical charge. For example, 1 mole of Na^+ ions and 1 mole of Cl^- ions are each 1 equivalent (or 1000 mEq) because they each contain 1 mole of charge.

- An ion with a charge of 2+ or 2– contains 2 equivalents per mole. Some examples of ions and their equivalents are shown below:

Ion	Electrical Charge	Number of Equivalents in 1 Mole
Na^+	1+	1 Eq
Ca^{2+}	2+	2 Eq
Fe^{3+}	3+	3 Eq
Cl^-	1–	1 Eq
SO_4^{2-}	2–	2 Eq

- In body, the charge of the positive ion is always balanced by the charge of the negative ion. For example, a solution containing 25 mEq/L of Na^+ and 4 mEq/L of K^+ must have 29 mEq/L of Cl^- to balance.

- Shown below are examples of some common intravenous solutions and their ion concentrations:

Solution	Electrolytes (mEq/L)	Use
Sodium chloride (0.9%)	Na^+ 154, Cl^- 154	Replacement of fluid loss
Potassium chloride with 5% dextrose	K^+ 40, Cl^- 40	Treatment of malnutrition (low potassium levels)
Ringer's solution	Na^+ 147, K^+ 4, Ca^{2+} 4, Cl^- 155	Replacement of fluids and electrolytes lost through dehydration
Maintenance solution with 5% dextrose	Na^+ 40, K^+ 35, Cl^- 40, lactate$^-$ 20, HPO_4^{2-} 15	Maintenance of fluid and electrolyte levels
Replacement solution (extracellular)	Na^+ 140, K^+ 10, Ca^{2+} 5, Mg^{2+} 3, Cl^- 103, acetate$^-$ 47, citrate^{3-} 8	Replacement of electrolytes in extracellular fluids

EQUIVALENTS OF ELECTROLYTES

Examples:

1. Indicate the number of equivalents in each of the following:

   a) 2 mol K^{+} ________Eq

   b) 0.5 mol Mg^{2+} ________Eq

   c) 3 mol CO_3^{2-} ________Eq

2. A typical concentration for Ca^{2+} in blood is 8.8 mEq/L. How many moles of Ca^{2+} are present in 0.50 L of blood?

$$0.50 \text{ L} \times \frac{8.8 \text{ mEq } Ca^{2+}}{1 \text{ L}} \times \frac{1 \text{ Eq}}{10^3 \text{ mEq}} \times \frac{1 \text{ mol } Ca^{2+}}{2 \text{ Eq } Ca^{2+}} = 0.0022 \text{ mol } Ca^{2+}$$

3. An IV solution contains 155 mEq/L of Cl^{-}. If a patient received 1250 mL of the IV solution, how many moles of chloride were given to him?

4. A sample of Ringer's solution contains the following concentrations (mEq/L) of cations: Na^{+} 147, K^{+} 4, and Ca^{2+} 4. If Cl^{-} is the only anion in the solution, what is the concentration of Cl^{-} in mEq/L?

SOLUBILITY AND SATURATION

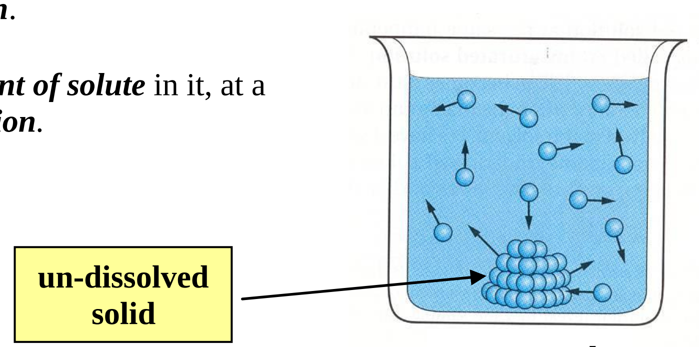
- ***Solubility*** refers to the ***maximum amount of solute*** that can be dissolved in a given ***amount of solvent***.

- Many factors such as ***type of solute***, ***type of solvent*** and ***temperature*** affect the solubility of a solute in a solution.

- Solubility is measured in ***grams of solute*** per ***100 grams of solvent*** at a ***given temperature***.

- A solution that ***does not contain the maximum amount of solute*** in it, at a given temperature, is called an ***unsaturated solution***.

- A solution that ***contains the maximum amount of solute*** in it, at a given temperature, is called a ***saturated solution***.



Saturated solution

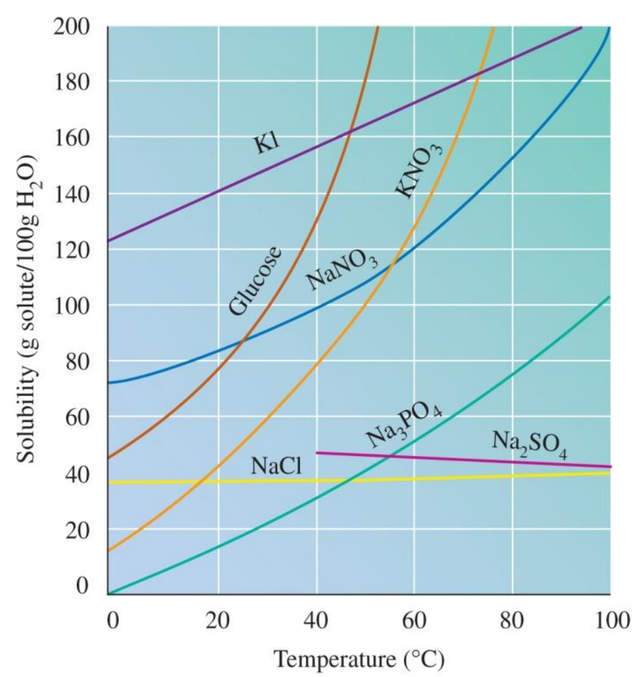
- Solubility of most ***solids*** in water ***increases*** as ***temperature increases***.

- Using solubility chart shown below, the solubility of a solute at a given temperature can be determined.

- For example, KNO_3 has a solubility of 80 g/100 g H_2O (80%) at 40 °C.

- Solubility of ***gases*** in water ***decreases*** as ***temperature increases***. At higher temperatures more gas molecules have the energy to escape from solution.

- ***Henry's law*** states that the solubility of a gas is directly proportional to the pressure above the liquid. For example, a can of soda is carbonated at high pressures in order to increase the solubility of CO_2. Once the can is opened, the pressure is reduced and the excess gas escapes from the solution.



SOLUBLE & INSOLUBLE SALTS

- Many ionic solids ***dissolve*** in water and are called ***soluble salts***. However, some ionic solids ***do not dissolve*** in water and do not form ions in solution. These salts are called ***insoluble salts*** and remain solid in solution.

1. Chemists use a set of ***solubility rules*** to ***predict*** whether a salt is ***soluble or insoluble***. These rules are summarized below:

SOLUBLE	NO_3^-	**No exceptions**
	Na^+, K^+, NH_4^+	**No exceptions**
	Cl^-, Br^-, I^-	**Excepts for those containing Ag^+, Pb^{2+}**
	SO_4^{2-}	**Except for those containing Ba^{2+}, Pb^{2+}, Ca^{2+}**
INSOLUBLE	S^{2-}, CO_3^{2-}, PO_4^{3-}	**Except those containing Na^+, K^+, NH_4^+**
	OH^-	**Except those containing Na^+, K^+, Ca^{2+}, NH_4^+**

<u>Examples:</u>

1. Use the solubility rules to determine if each of the following salts are soluble or insoluble:

   a) K_3PO_4 ____________________

   b) $CaCO_3$ ____________________

2. Using the solubility chart, determine if each of the following solutions is saturated or unsaturated at 20°C:

   a) 25 g NaCl in 100 g of water ______________________

   b) 11 g $NaNO_3$ in 25 g of water ______________________

   c) 400. g of glucose in 125 g of water ______________________

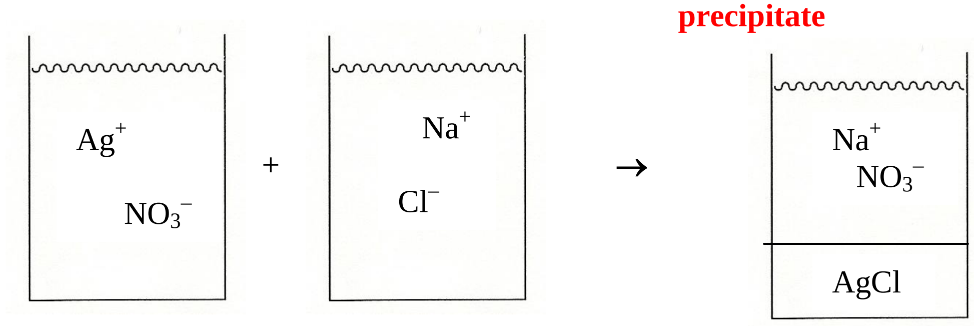
FORMATION OF A SOLID

- Solubility rules can be used to predict whether a solid, called a ***precipitate***, can be formed when two solutions of ionic compounds are mixed.

- A solid is formed when two ions of an insoluble salt come in contact with one another.

- For example, when a solution of $AgNO_3$ is mixed with a solution of NaCl, a white insoluble salt AgCl is produced.

$$\mathbf{AgNO_3\,(aq) + NaCl\,(aq) \rightarrow \underset{\text{precipitate}}{AgCl\,(s)} + NaNO_3\,(aq)}$$



- Double replacement reactions in which a precipitate is formed are called ***precipitation*** reactions.

- The solubility rules can be used to predict whether a precipitate forms when two solutions of ionic compounds are mixed together. The stepwise process is outlined below:

    1. Write the ***molecular equation*** for the reaction by predicting the products formed from the combination of the reactants. Use the solubility rules to determine if any of the products are insoluble. Label all the states and balance the equation.

    2. Using the molecular equation above, write the ***complete ionic equation*** by breaking all of the soluble compounds into their corresponding ions; leave the precipitate compound together as molecular.

    3. Using the complete ionic equation above, write the ***net ionic equation (NIE)*** by cancelling all ions that appear as the same on both sides of the equation. These ions are called ***spectator*** ions.

    4. If no precipitate forms in step 2, write "NO REACTION" after the arrow and stop.

PRECIPITATION REACTIONS

- For example, the reaction of $AgNO_3$ and NaCl, can be predicted as shown below:

Step 1: $AgNO_3\,(aq) + NaCl\,(aq) \rightarrow$?????????

$AgNO_3\,(aq) + NaCl\,(aq) \rightarrow AgCl\,(?) + NaNO_3\,(?)$

$AgNO_3\,(aq) + NaCl\,(aq) \rightarrow AgCl\,(s) + NaNO_3\,(aq)$ (**molecular equation**)

Step 2: $Ag^+ + NO_3^- + Na^+ + Cl^- \rightarrow AgCl\,(s) + Na^+ + NO_3^-$ (**complete ionic eq**)

Step 3: $-Ag^+ + \cancel{NO_3^-} + \cancel{Na^+} + Cl^- \rightarrow AgCl\,(s) + \cancel{Na^+} + \cancel{NO_3^-}$

Net Ionic Equation: $\mathbf{Ag^+ + Cl^- \rightarrow AgCl\,(s)}$
precipitate

Examples:
Predict the products for each reaction shown below and write molecular, complete ionic and net ionic equations. If no reaction occurs, write "No Reaction" after the arrow.

1. $Na_2SO_4\,(aq) + Pb(NO_3)_2\,(aq) \rightarrow$???????

Step 1:

Step 2

:

Step 3:

PRECIPITATION REACTIONS

2. $PbCl_2$ (aq) + KI (aq) $\rightarrow$???????????

Step 1:

Step 2:

Step 3:

3. NH_4Cl (aq) + KNO_3 (aq) $\rightarrow$???????????

PERCENT CONCENTRATION

- The ***amount of solute*** dissolved in a certain ***amount of solution*** is called ***concentration***.

$$\textbf{Concentration} = \frac{\textbf{amount of solute}}{\textbf{amount of solution}}$$

MASS PERCENT:

- ***Mass percent*** (***% m/m***) of a solution is the ***mass of solute*** divided by the ***mass of solution***.

$$\textbf{mass \% (m/m)} = \frac{\textbf{mass of solute}}{\textbf{mass of solution}}\,\textbf{x100}$$

$$\textbf{mass of solution = mass of solute + mass of solvent}$$

MASS/VOLUME PERCENT:

- ***Mass/Volume percent*** (***% m/v***) of a solution is the ***mass of solute*** divided by the ***volume of solution***.

$$\textbf{mass \% (m/v)} = \frac{\textbf{mass of solute}}{\textbf{volume of solution}}\,\textbf{x100}$$

Examples:

1. What is the mass % (m/m) of a solution prepared by dissolving 30.0 g of NaOH in 120.0 g of water?

   mass of solution =

   $$\text{mass \% (m/m)} = \frac{\text{mass of NaOH}}{\text{mass of solution}}\,\text{x100} =$$

2. What is the mass % (m/v) of a solution prepared by dissolving 5.0 g of KI to give a final volume of 250 mL?

   $$\text{mass \% (m/v)} = \frac{\text{mass of KI}}{\text{volume of solution}}\,\text{x100} =$$

USING PERCENT CONCENTRATION

- In the preparation of solutions, one often needs to calculate the amount of solute or solution. To achieve this, percent composition can be used as a conversion factor.

- Some examples of percent compositions, their meanings, and possible conversion factors are shown in the table below:

Percent Concentration	Meaning	Conversion Factors		
15% (m/m) KCl	There are 15 g of KCl in 100 g of solution.	$\frac{15 \text{ g KCl}}{100 \text{ g solution}}$	and	$\frac{100 \text{ g solution}}{15 \text{ g KCl}}$
5% (m/v) glucose	There are 5 g of glucose in 100 mL of solution.	$\frac{5 \text{ g glucose}}{100 \text{ mL solution}}$	and	$\frac{100 \text{ mL solution}}{5 \text{ g glucose}}$

Examples:

2. A topical antibiotic solution is 1.0% (m/v) Clindamycin. How many grams of Clindamycin are in 65 mL of this solution?

   65 mL of solution x ________________ = g of Clindamycin

3. How many grams of KCl are in 225 g of an 8.00% (m/m) solution?

4. How many grams of solute are needed to prepare 150 mL of a 40.0% (m/v) solution of $LiNO_3$?

MOLARITY

- The most common unit of concentration used in the laboratory is **molarity (M).** Molarity is defined as:

$$\textbf{Molarity} = \frac{\textbf{moles of solute}}{\textbf{Liter of solution}}$$

Examples:

1. What is the molarity of a solution containing 1.4 mol of acetic acid in 250 mL of solution?

$$250\ \cancel{\text{mL}} \times \frac{1\text{ L}}{1000\ \cancel{\text{mL}}} = 0.25\text{ L}$$

$$\text{Molarity} = \frac{1.4\text{ mol of acetic acid}}{0.25\text{ L of solution}} = 5.6\text{ M}$$

2. What is the molarity of a solution prepared by dissolving 60.0 g of NaOH in 0.250 L of solution?

First, calculate the number of moles of solute:

$$60.0\text{ g of NaOH} \times \frac{1\text{ mol}}{40.0\text{ g}} = 1.50\text{ mol of NaOH}$$

Next, calculate the molarity of solution:

$$\text{M} = \frac{1.50\text{ mol of NaOH}}{0.250\text{ L of solution}} = 6.00\text{ M}$$

3. What is the molarity of a solution that contains 75 g of KNO_3 in 350 mL of solution?

Calculate moles of solute:

Calculate molarity:

USING MOLARITY

Calculating moles or mass of solute:

1. How many moles of nitric acid are in 325 mL of 16 M HNO_3 solution?

$$325\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 0.325\ \text{L}$$

$$0.325\ \text{L of solution} \times \frac{16\ \text{mol}}{1\ \text{L of solution}} = 5.2\ \text{mol of } HNO_3$$

2. How many grams of KCl would you need to prepare 0.250 L of 2.00 M KCl solution?

First, calculate the number of moles of solute:

$$0.250\ \text{L of solution} \times \frac{2.00\ \text{mol}}{1\ \text{L of solution}} = 0.500\ \text{mol of KCl}$$

Next, calculate the mass of solute:

$$0.500\ \text{mol of KCl} \times \frac{74.6\ \text{g}}{1\ \text{mol}} = 37.3\ \text{g of KCl}$$

3. How many grams of $NaHCO_3$ are in 325 mL of 4.50 M solution of $NaHCO_3$?

Calculate moles of solute:

Calculate mass of solute:

USING MOLARITY

Calculating volume of solutions:

4. What volume (L) of 1.5 M HCl solution contains 6.0 moles of HCl?

$$6.0 \text{ mol HCl x } \frac{1 \text{ L solution}}{1.5 \text{ mol HCl}} = 4.0 \text{ L of solution}$$

5. What volume (mL) of 2.0 M NaOH solution contains 20.0 g of NaOH?

First, calculate the number of moles of NaOH:

$$20.0 \text{ g NaOH x } \frac{1 \text{ mol}}{40.0 \text{ g}} = 0.500 \text{ mol of NaOH}$$

Next, calculate the volume of solution:

$$0.500 \text{ mol NaOH x } \frac{1 \text{ L solution}}{2.0 \text{ mol NaOH}} = 0.25 \text{ L of solution}$$

$$0.25 \text{ L x } \frac{1000 \text{ mL}}{1 \text{ L}} = 250 \text{ mL}$$

6. How many mL of a 0.300 M glucose ($C_6H_{12}O_6$) intravenous solution is needed to deliver 10.0 g of glucose to the patient?

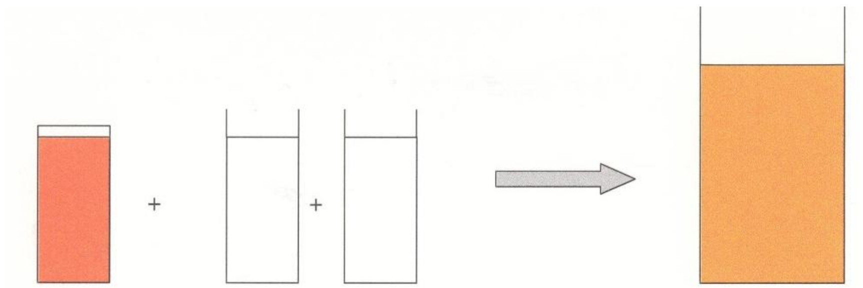
Calculate mole of solute:

Calculate volume of solution:

DILUTION

- Solutions are often prepared from more concentrated ones by adding water. This process is called ***dilution***.

- When more water is added to a solution, the ***volume increases***, causing a ***decrease in concentration***. However, the ***amount of solute does not change***.



Volume and Concentration are inversely proportional

- The amount of solute depends on the concentration and the volume of the solution. Therefore,

$$M_1 \times V_1 = M_2 \times V_2$$

Examples:

1. What is the molarity of the final solution when 75 mL of 6.0 M KCl solution is diluted to 150 mL?

$M_1 = 6.0$ M $\quad V_1 = 75$ mL

$M_2 = ???$ $\quad V_2 = 150$ mL

$$M_2 = \frac{M_1 V_1}{V_2} = \frac{(6.0\text{ M})(75\text{ mL})}{150\text{ mL}} = 3.0\text{ M}$$

2. What volume (mL) of 0.20 M HCl solution can be prepared by diluting 50.0 mL of 1.0 M HCl?

$M_1 =$ $\quad V_1 =$

$M_2 =$ $\quad V_2 =$

OSMOLARITY

- Many important properties of solutions depend on the number of particles formed in solution.

- Recall that when ionic substances (strong electrolytes) dissolve in water they form several particles for each formula unit. For example:

$$NaCl\,(s) \xrightarrow{H_2O} Na^{+}\,(aq) + Cl^{-}\,(aq)$$

1 formula unit → 2 particles

$$CaCl_2\,(s) \xrightarrow{H_2O} Ca^{2+}\,(aq) + 2\ Cl^{-}\,(aq)$$

1 formula unit → 3 particles

- When covalent substances (non- or weak electrolytes) dissolve in water they form only one particle for each formula unit. For example:

$$C_{12}H_{22}O_{11}\,(s) \xrightarrow{H_2O} C_{12}H_{22}O_{11}\,(aq)$$

1 formula unit → 1 particle

- Osmolarity of a solution is its molarity multiplied by the number of particles formed in solution.

Osmolarity = i x Molarity

Solution	**i**	**Osmolarity**
0.10 M NaCl		
0.10 M $CaCl_2$		
0.10 M $C_{12}H_{22}O_{11}$		

TONICITY OF SOLUTIONS

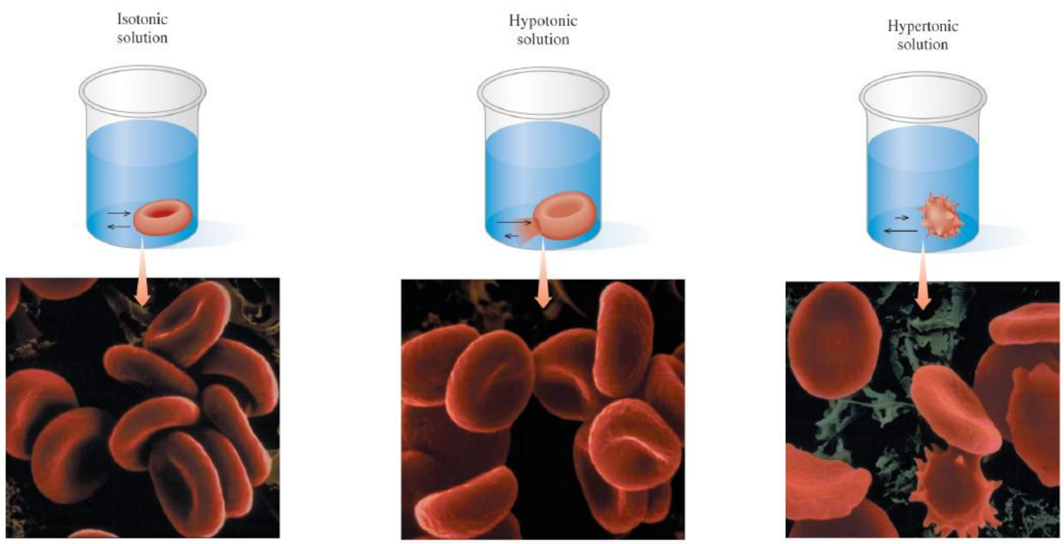
- Because the cell membranes in biological systems are semipermeable, particles of solute in solutions can travel in and out of the membranes. This process is called ***osmosis***.

- The direction of the flow of solutions in or out of the cell membranes is determined by the relative osmolarity of the cell and the solution. The comparison of osmolarity of a solution with those in body fluids determines the ***tonicity*** of a solution.

- Solutions with the ***same*** osmolarity as the cells (0.30) are called ***isotonic***. These solutions are called physiological solutions and allow red blood cells to retain their normal volume. See diagram (a) below.

- Solutions with ***lower*** osmolarity than the cells are called ***hypotonic***. In these solutions, water flows into a red blood cell, causing it to swell and burst (***hemolysis***). See diagram (b) below.

- Solutions with ***greater*** osmolarity than the cells are called ***hypertonic***. In these solutions, water leaves the red blood cells causing it to shrink (***crenation***). See diagram (c) below.



(a) (b) (c)

<u>Examples:</u>

1. Determine the tonicity of the solutions shown below:

Solution	Osmolarity	Tonicity
0.10 M NaCl		
0.10 M $CaCl_2$		
0.10 M $C_{12}H_{22}O_{11}$		