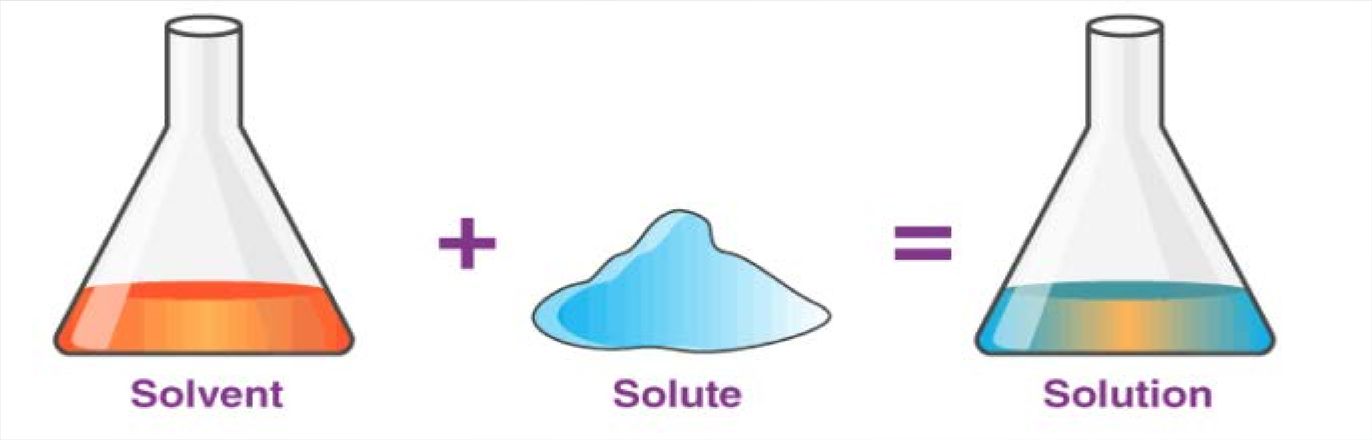
Higher technological institute
Chemical engineering department

Chapter 2 Solutions

A solution is a type of mixture.

- **Solution:** a type of mixture that is the same throughout.
  - A solution can be physically separated
  - All portions of the solution have the same properties
- **Homogeneous:** the same
- **Examples: salt water, blood, sugar water, gasoline**



- **Solute:** a substance that is dissolved to make a solution.
  - **when dissolved it separates into individual particles**
- **Solvent:** is a substance that dissolves a solute.
  - when the solute is dissolved into the solvent it is not possible to identify the solvent and solute as individual parts

Solute is the substance present in the smallest amount.

Solvent is the substance present in the greatest amount.

Aqueous solutions are solutions in which water is the solvent.

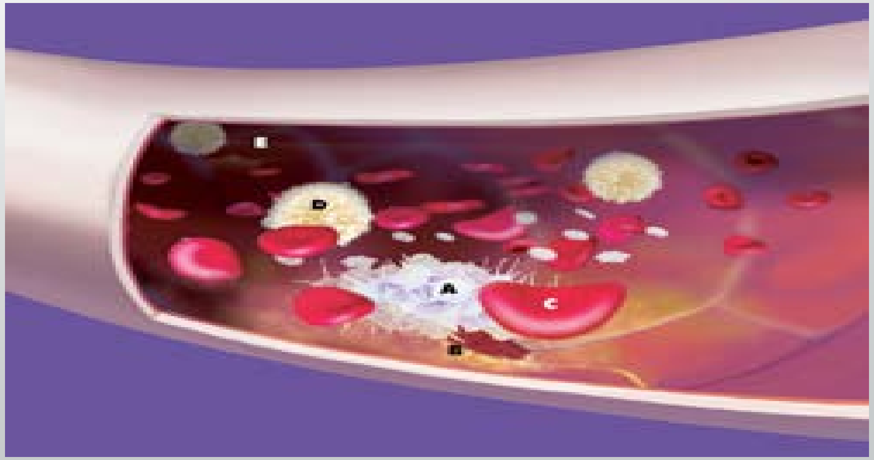
The solute and solvent can either be a solid, liquid, or gas in a solution.

Examples of Solutes and Solvents

- Saltwater: salt=solute, water=solvent

saltwaterfishing365.com

- Blood: calcium ions, sugars=solutes, water =solvent



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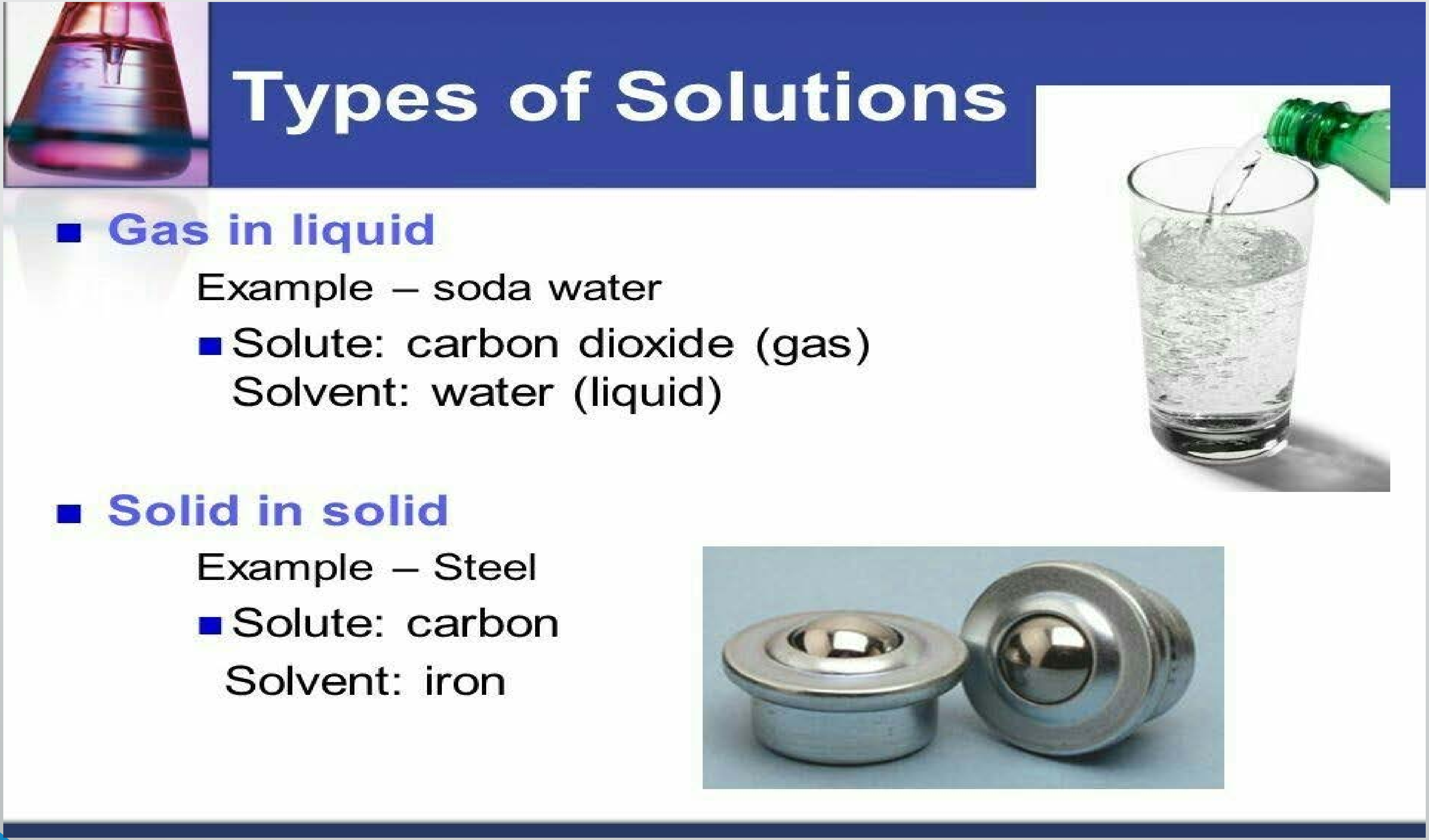
Types of Solutions

- **Solutions can be made from solids, liquids, and gases**

- **<u>Air:</u> solute=oxygen, solvent=nitrogen (oxygen is dissolved in nitrogen)**

- **Humidity: solute=liquid, solvent=gas (water is dissolved into air)**

- **Stainless steel: solute=chromium metal, solvent=iron (chromium metal is dissolved in iron to form a shiny steel)**



Types of Solutions

- **Gas in liquid**

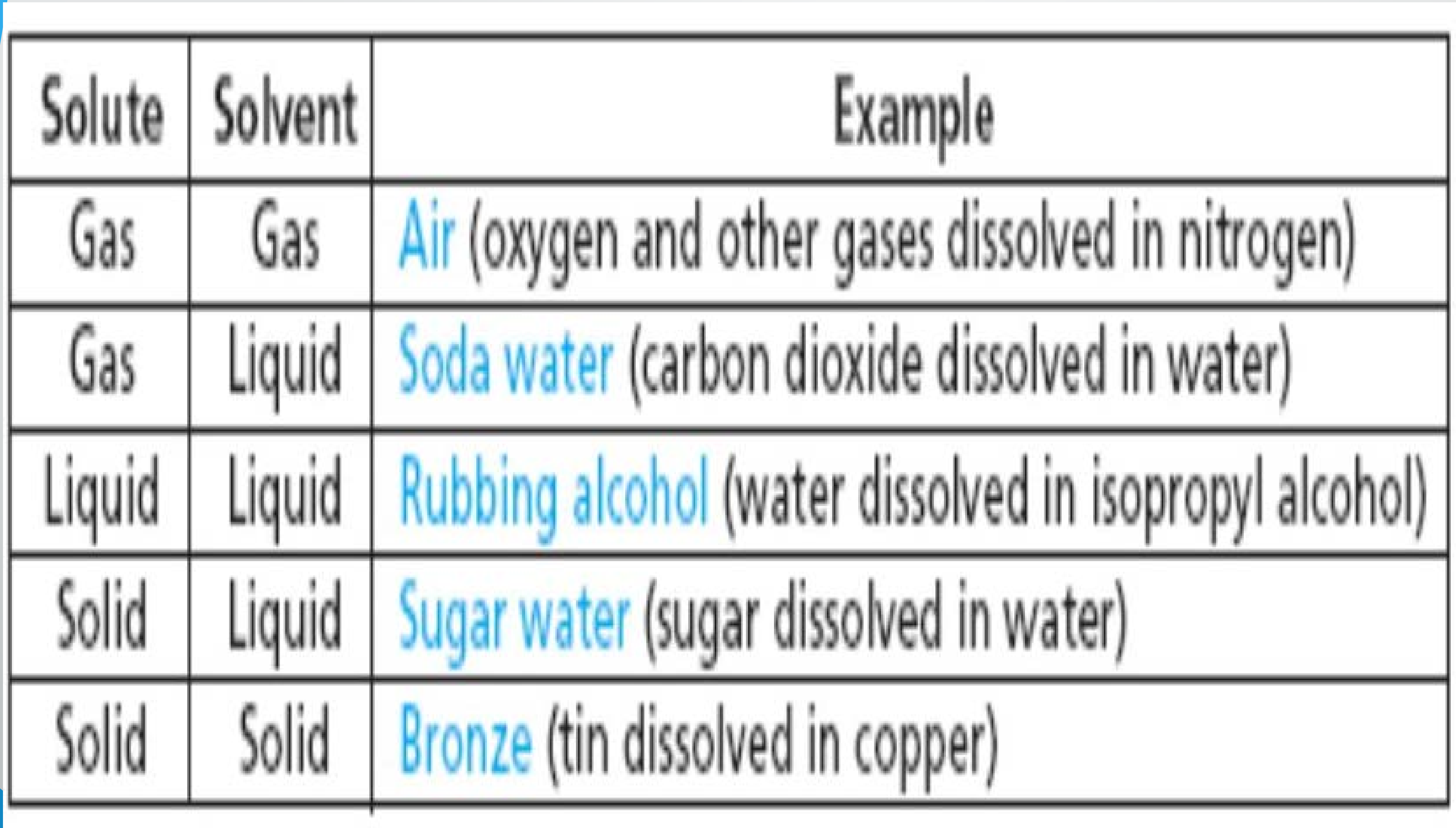
  Example – soda water

  - Solute: carbon dioxide (gas)
    Solvent: water (liquid)

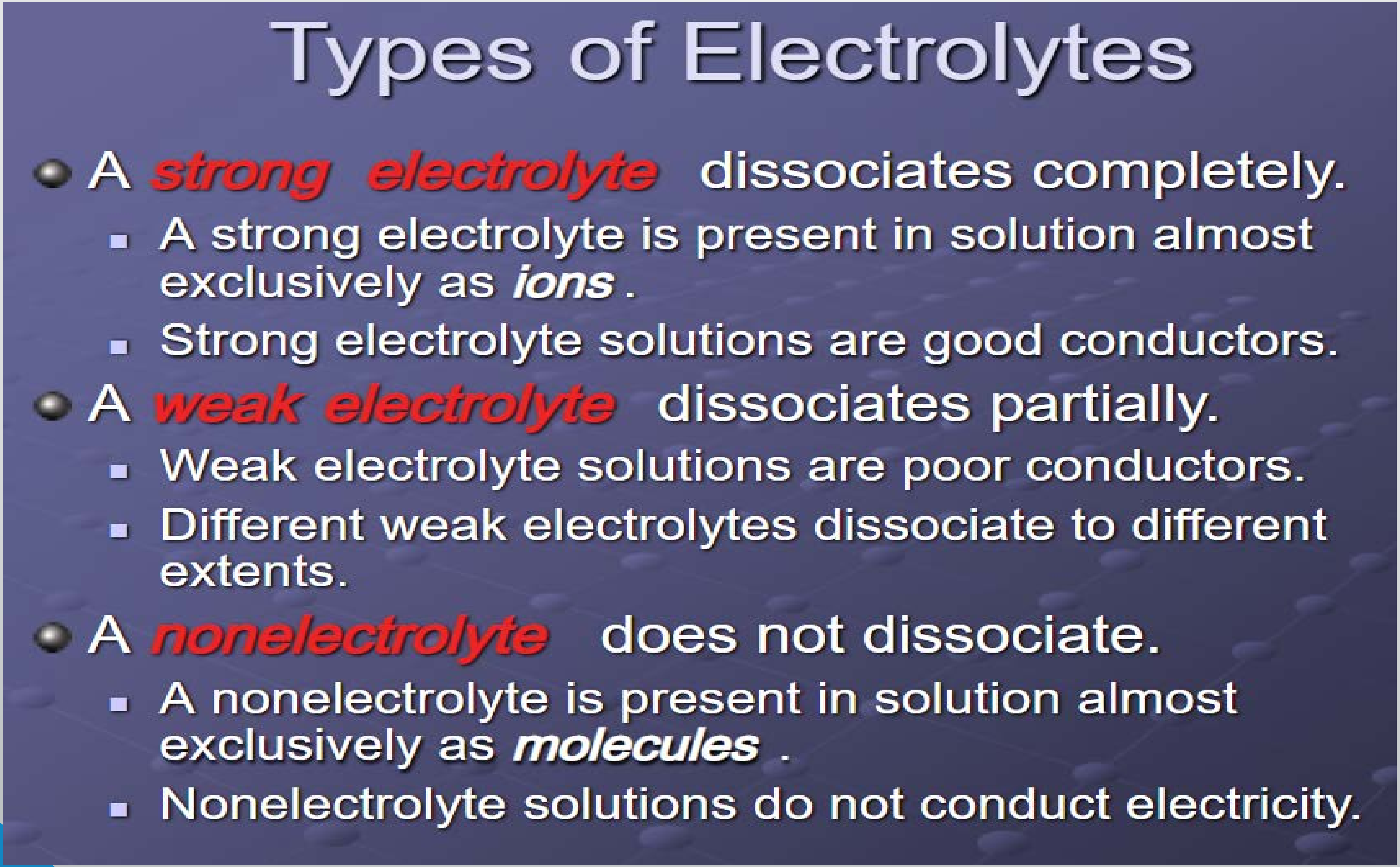
- **Solid in solid**

  Example – Steel

  - Solute: carbon
    Solvent: iron

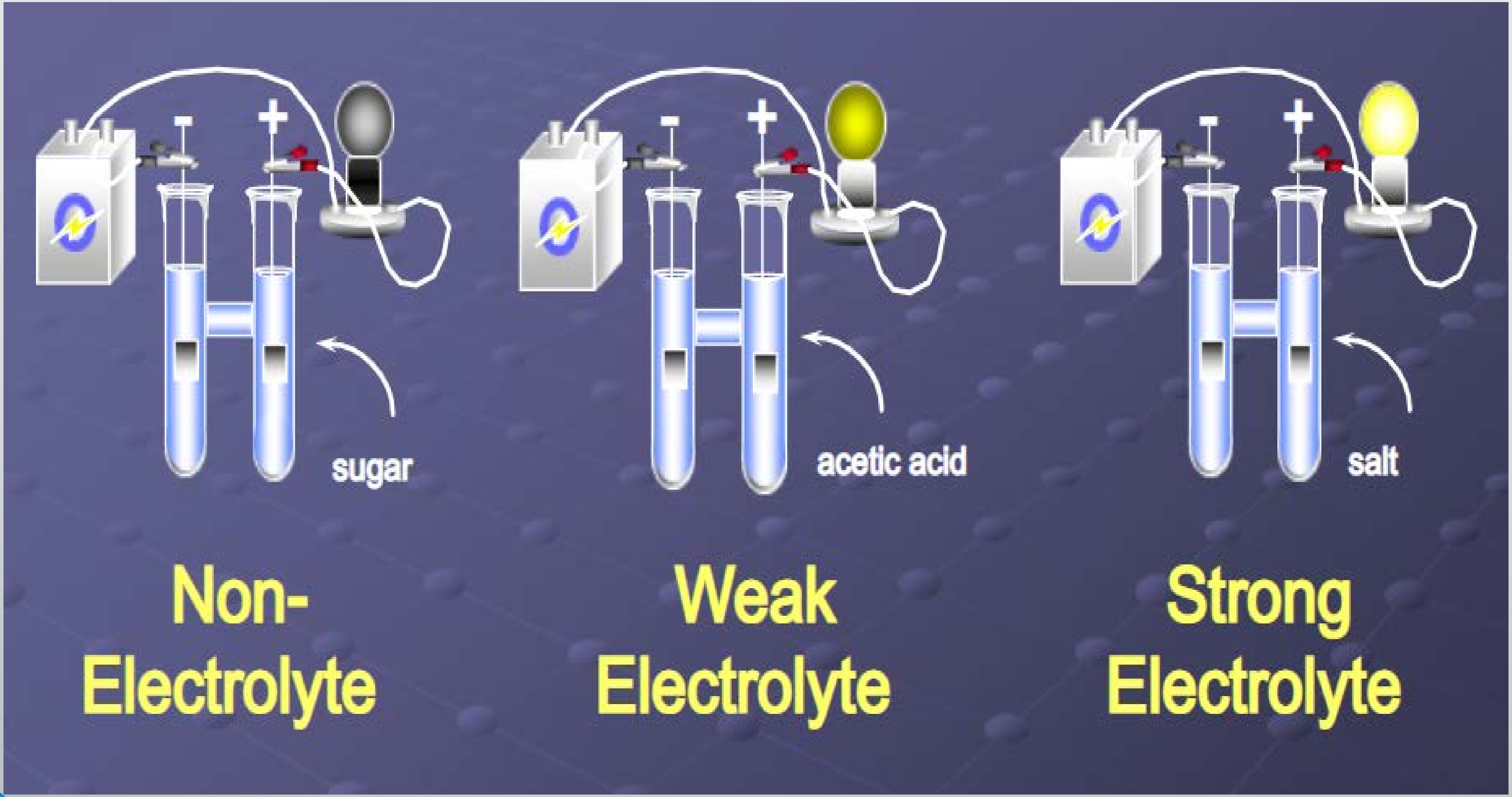


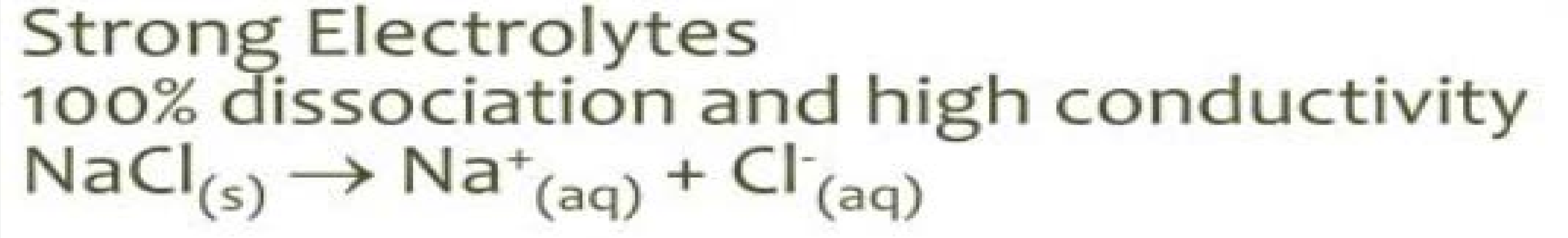
Solute	Solvent	Example
Gas	Gas	Air (oxygen and other gases dissolved in nitrogen)
Gas	Liquid	Soda water (carbon dioxide dissolved in water)
Liquid	Liquid	Rubbing alcohol (water dissolved in isopropyl alcohol)
Solid	Liquid	Sugar water (sugar dissolved in water)
Solid	Solid	Bronze (tin dissolved in copper)



Types of Electrolytes

- A ***strong electrolyte*** dissociates completely.
  - A strong electrolyte is present in solution almost exclusively as *ions* .
  - Strong electrolyte solutions are good conductors.
- A ***weak electrolyte*** dissociates partially.
  - Weak electrolyte solutions are poor conductors.
  - Different weak electrolytes dissociate to different extents.
- A ***nonelectrolyte*** does not dissociate.
  - A nonelectrolyte is present in solution almost exclusively as *molecules* .
  - Nonelectrolyte solutions do not conduct electricity.


-
+
sugar
Non-Electrolyte
-
+
acetic acid
Weak Electrolyte
-
+
salt
Strong Electrolyte



Strong Electrolytes

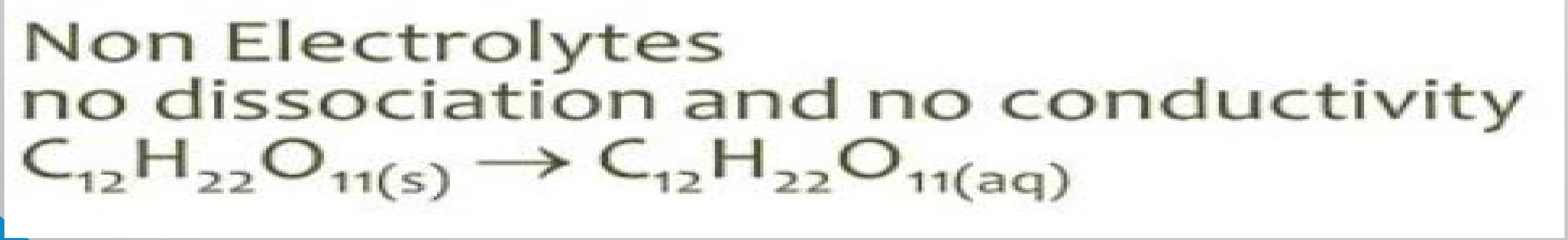
100% dissociation and high conductivity

$NaCl_{(s)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$

Weak Electrolytes

partial dissociation and partial conductivity

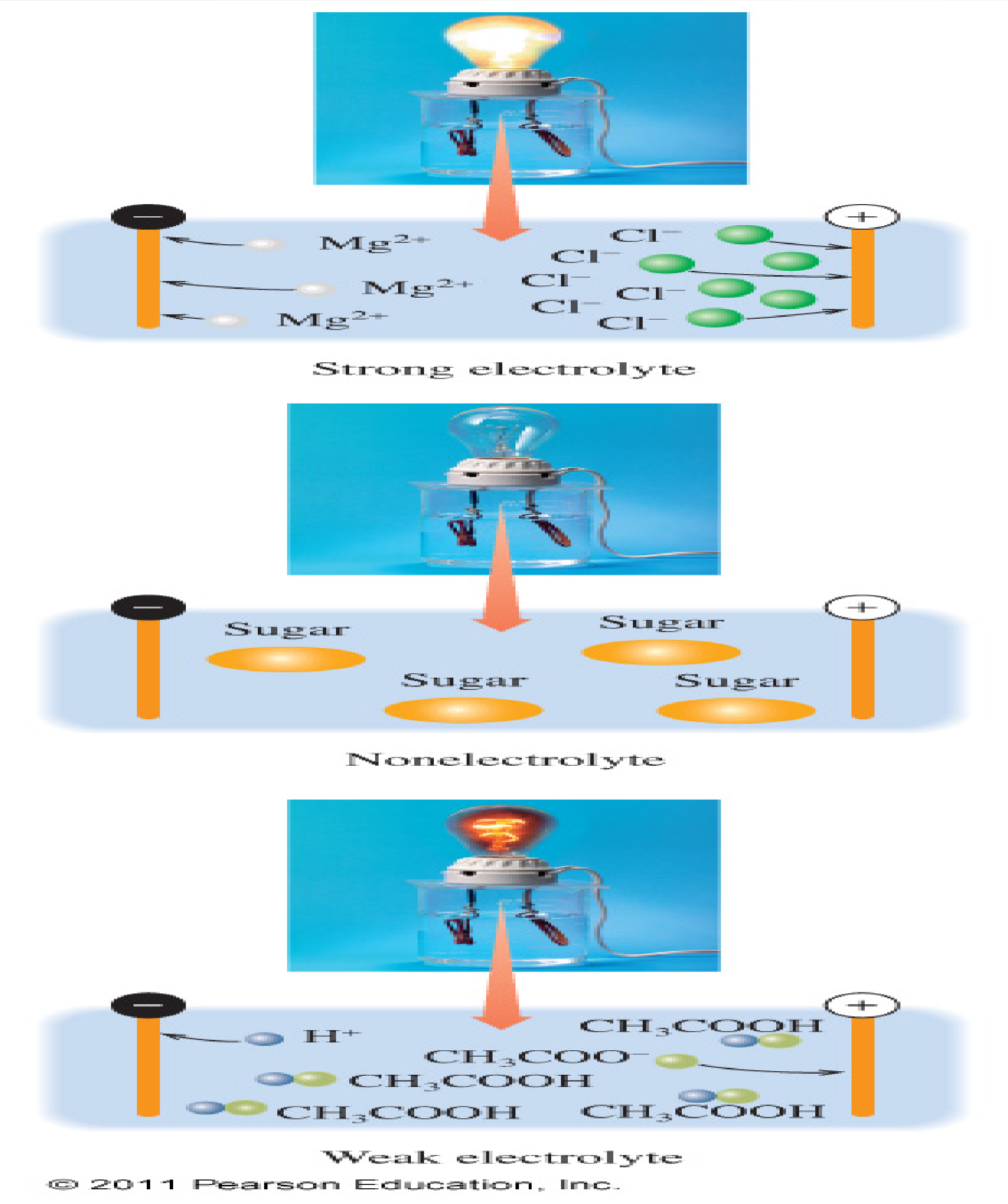
$CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^{-}_{(aq)} + H^{+}_{(aq)}$



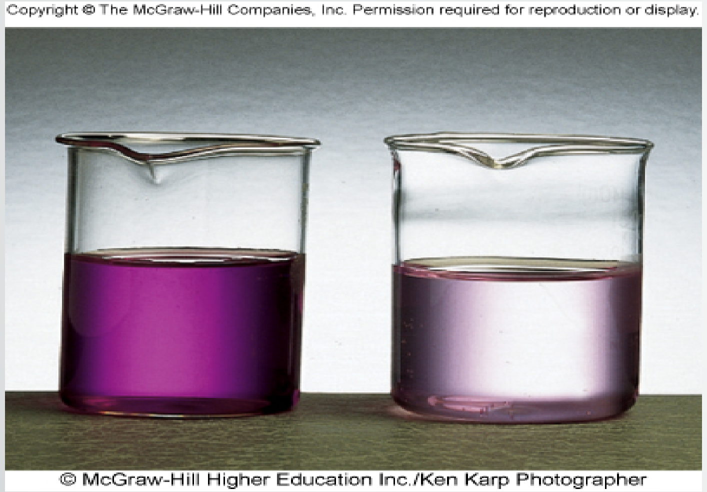
Non Electrolytes

no dissociation and no conductivity

$C_{12}H_{22}O_{11(s)} \rightarrow C_{12}H_{22}O_{11(aq)}$


Mg^{2+}
Mg^{2+}
Mg^{2+}
Cl$^-$
Cl$^-$
Cl$^-$
Cl$^-$
Cl$^-$
Cl$^-$
Strong electrolyte
Sugar
Sugar
Sugar
Sugar
Nonelectrolyte
H$^+$
CH_3COOH
CH_3COO^-
CH_3COOH
CH_3COOH
CH_3COOH
Weak electrolyte

Solubility is the property of a substance (solute) to dissolve in another substance (solvent); solutions are characterized by their capacity to do so



aqueous solutions of $KMnO_4$

CdS PbS $Ni(OH)_2$ $Al(OH)_3$

- Soluble the solutes ability to dissolve in a solvent

- Insoluble the solute is not able to dissolve in a solvent

Solubility Units

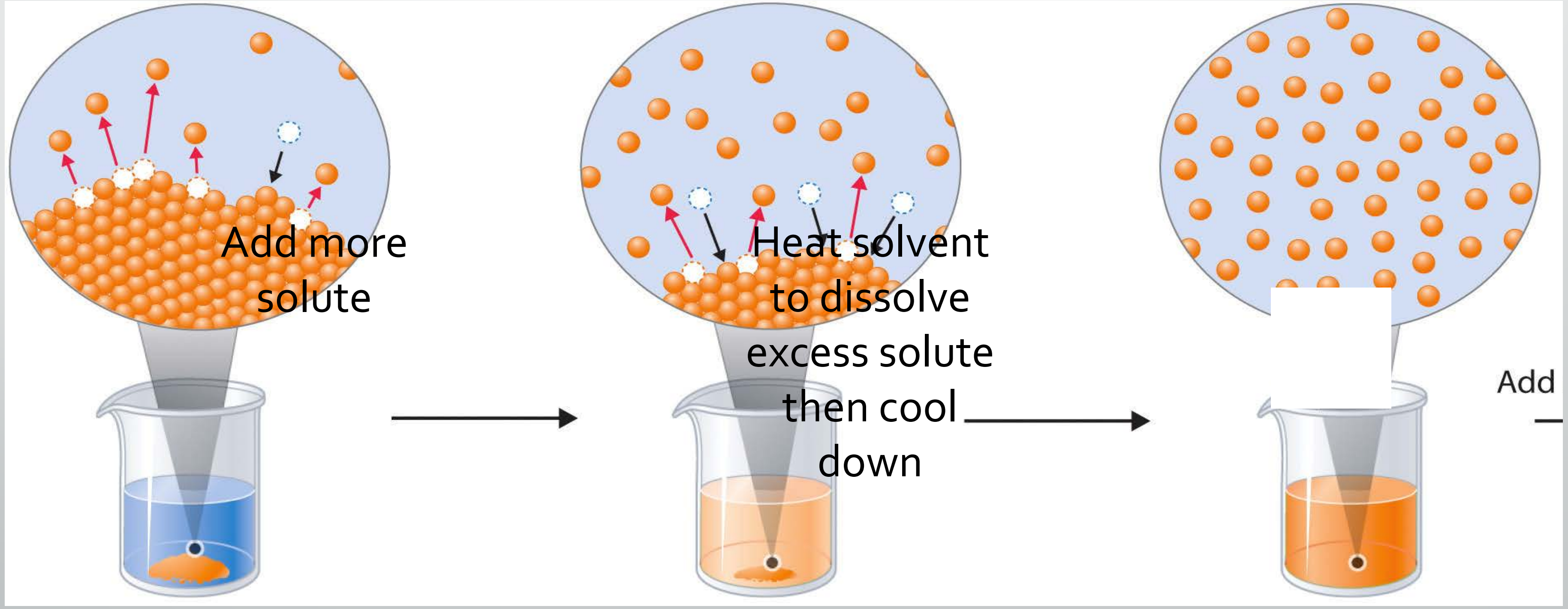
$$\frac{grams\ of\ solute}{100\ grams\ of\ solvent}\ or\ \frac{grams\ of\ solute}{1\ Liter\ of\ solution}$$

Saturation the state that occurs when no more of something can be absorbed, combined, or added

Unsaturated

Saturated

Supersaturated



Add more solute

Heat solvent to dissolve excess solute then cool down

A ***saturated solution*** contains the maximum amount of a solute that will dissolve in a given solvent at a specific temperature.

An ***unsaturated solution*** contains less solute than the solvent has the capacity to dissolve at a specific temperature.

A ***supersaturated solution*** contains more solute than is present in a saturated solution at a specific temperature.

UNSATURATED SOLUTION

more solute dissolves

SATURATED SOLUTION

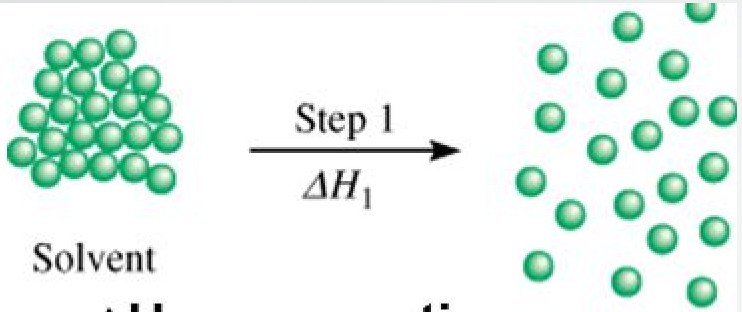
no more solute dissolves

SUPERSATURATED SOLUTION

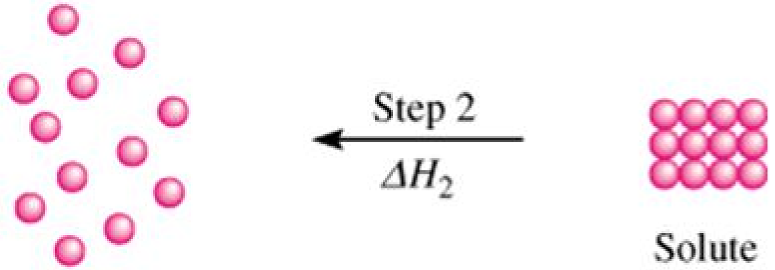
becomes unstable, crystals form

Three types of interactions in the solution process:

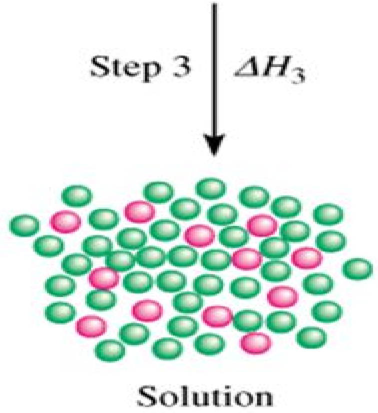
- ***solvent-solvent*** interaction
- ***solute-solute*** interaction
- ***solvent-solute*** interaction



ΔH_1 = separation of solvent molecules



ΔH_2 = separation of solute molecules



ΔH_3 = formation of solute-solvent interaction

$$\Delta H_{soln} = \Delta H_1 + \Delta H_2 + \Delta H_3$$

ΔH_{soln} can either be positive or negative depending on the intermolecular forces.

- If the solute-solvent attraction is stronger than the solvent-solvent attraction and solute-solute attraction,

The solution process is favorable, or exothermic ($\Delta H_{soln} < 0$).

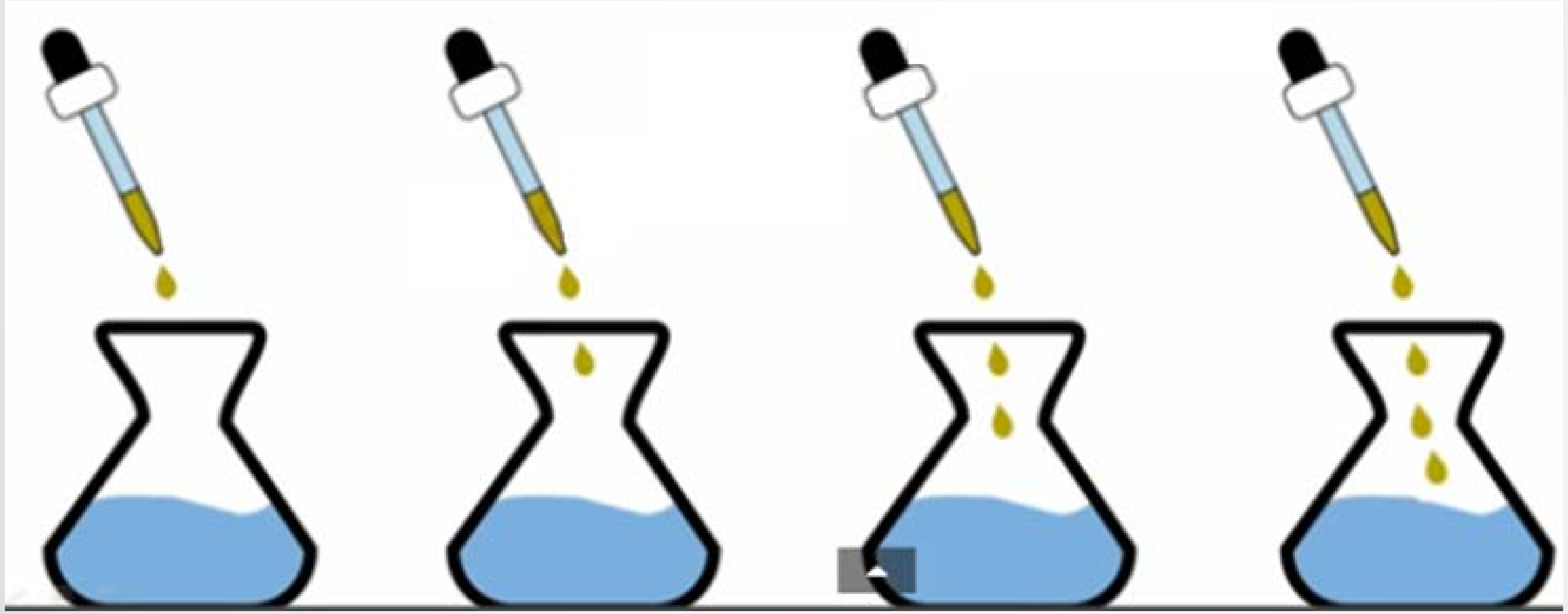
- When energy is released, the reaction is exothermic; ΔH is negative (−).

- If the solute-solvent interaction is weaker than the solvent-solvent and solute-solute interactions,

The solution process is endothermic ($\Delta H_{soln} > 0$).

- When energy is absorbed, the reaction is endothermic; ΔH is positive (+).

The **Heat of Solution** is the amount of heat energy absorbed (endothermic) or released (exothermic) when a specific amount of solute dissolves in a solvent. Molar heat of solution is the amount of energy absorbed or released per one mole of the solute.

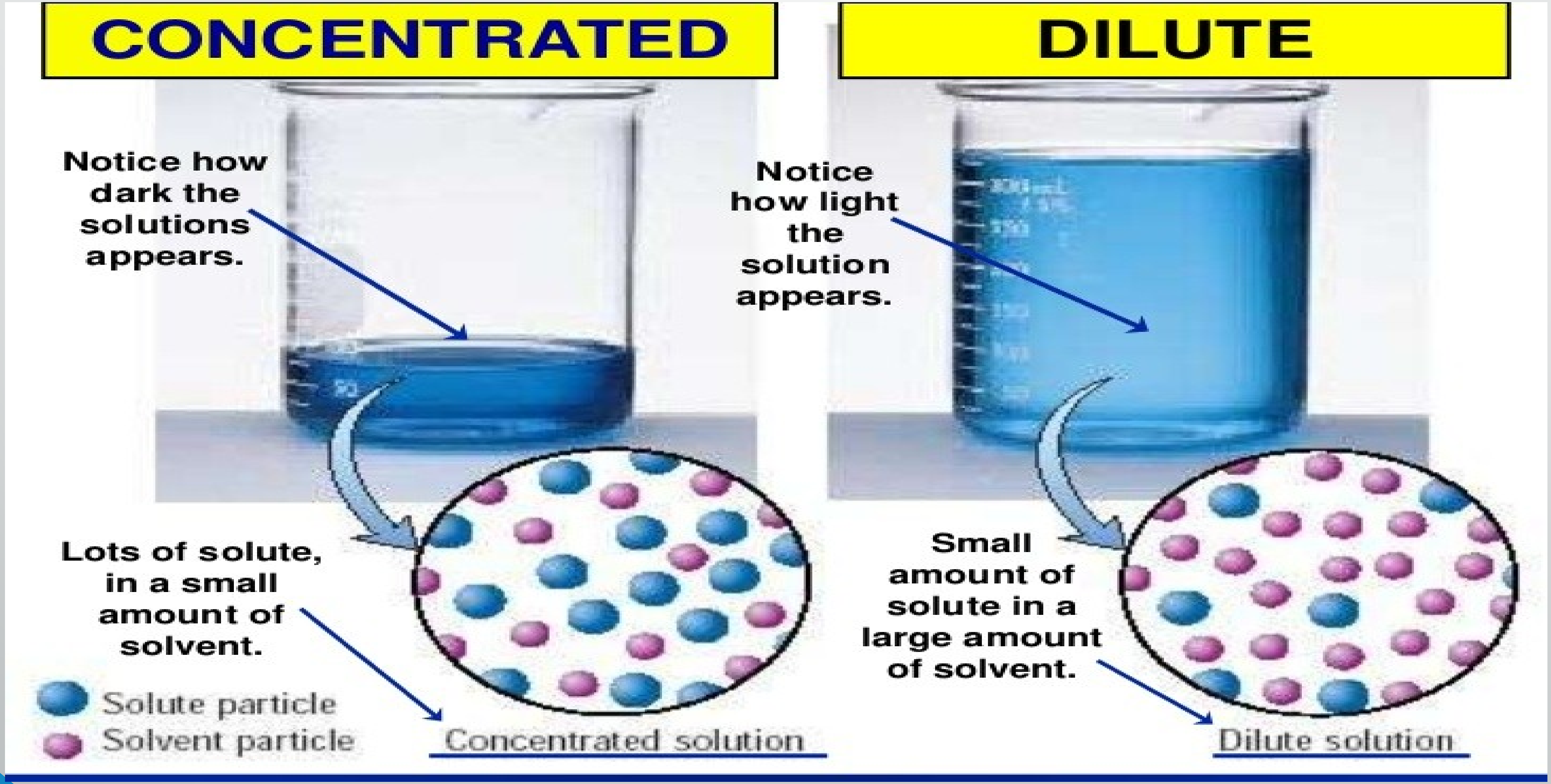


Are the solutions the same?
How can we describe them?

Concentration is the amount of solute present for a given amount of solvent or solution; relative or absolute

Concentrated contains greater amount of solute per volume of solution

Dilute contains relatively small amounts of solute per volume of solution



Notice how dark the solutions appears.

Notice how light the solution appears.

Lots of solute, in a small amount of solvent.

Small amount of solute in a large amount of solvent.

Concentration Units

The *concentration* of a solution is the amount of solute present in a given quantity of solvent or solution.

Percent by Mass or by Wight

$$\% \text{ by mass} = \frac{\text{mass of solute (g)}}{\text{mass of solute (g)} + \text{mass of solvent(g)}} \times 100\%$$

$$= \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

Unit %

Parts per Million and Parts per Billion

Parts per Million (ppm)

$$\mathbf{ppm} = \frac{\text{mass of solute in solution}}{\text{total mass of solution}} \times 10^6$$

Parts per Billion (ppb)

$$\text{ppb} = \frac{\text{mass of solute in solution}}{\text{total mass of solution}} \times 10^9$$

Mole Fraction (*X*)

$$X_A = \frac{moles\ of\ solute}{moles\ of\ solute + moles\ of\ solvent}$$

- In some applications, one needs the mole fraction of *solvent*, not solute—make sure you find the quantity you need!

Molarity (*M*)

$$M = \frac{\text{mol of solute}}{\text{L of solution}}$$

- Because volume is temperature dependent, molarity can change with temperature.

molality (*m*)

$$m = \frac{\text{mol of solute}}{\text{kg of solvent}}$$

Because neither moles nor mass change with temperature, molality (unlike molarity) is *not* temperature dependent.

Mass/Mass

$$Mass\% = \frac{mass\ A}{total\ mass\ of\ solution} \times 100$$

$$ppm = \frac{mass\ A}{total\ mass\ of\ solution} \times 10^6$$

$$ppb = \frac{mass\ A}{total\ mass\ of\ solution} \times 10^9$$

Moles/Moles

$$X_A = \frac{moles\ of\ solute}{total\ moles}$$

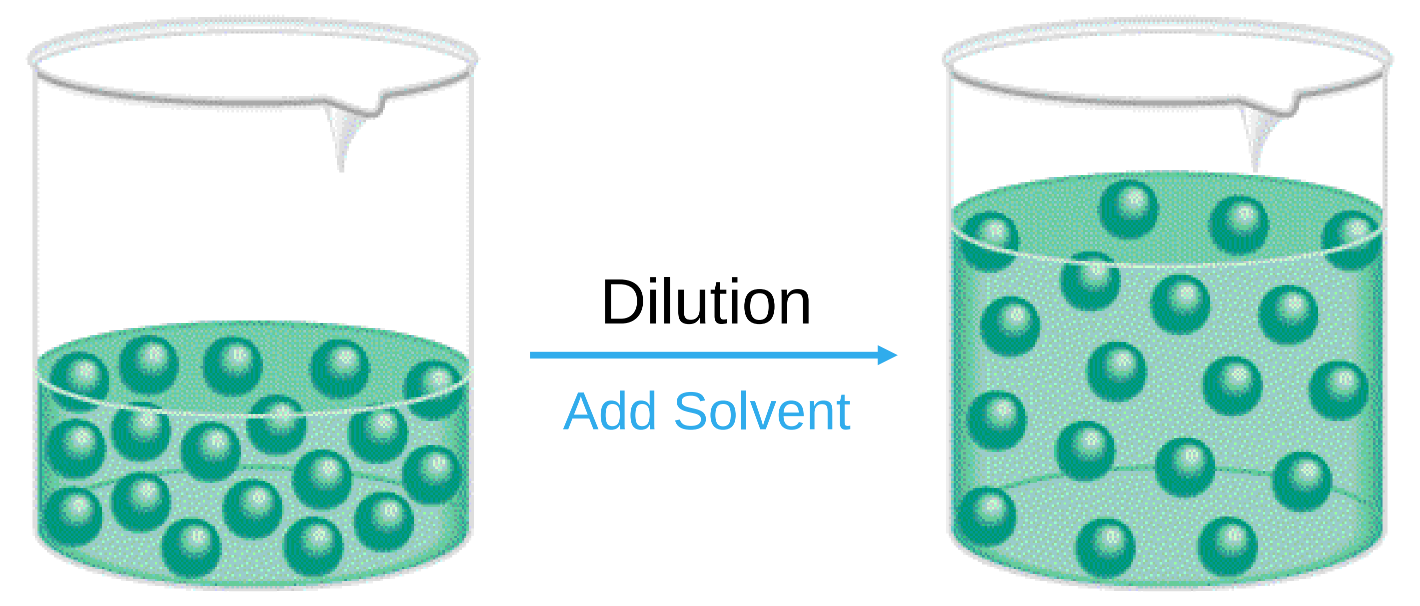
Moles/L

$$M = \frac{moles\ of\ solute}{L\ of\ solution}$$

Moles/Mass

$$m = \frac{moles\ of\ solute}{kg\ of\ solvent}$$

Dilution is the procedure for preparing a less concentrated solution from a more concentrated solution.



No of moles of concentrated solution = no of moles of diluted solution

NV) con = NV)dil

MV) con = MV)dil

$$M_iV_i = M_fV_f$$

Titrations

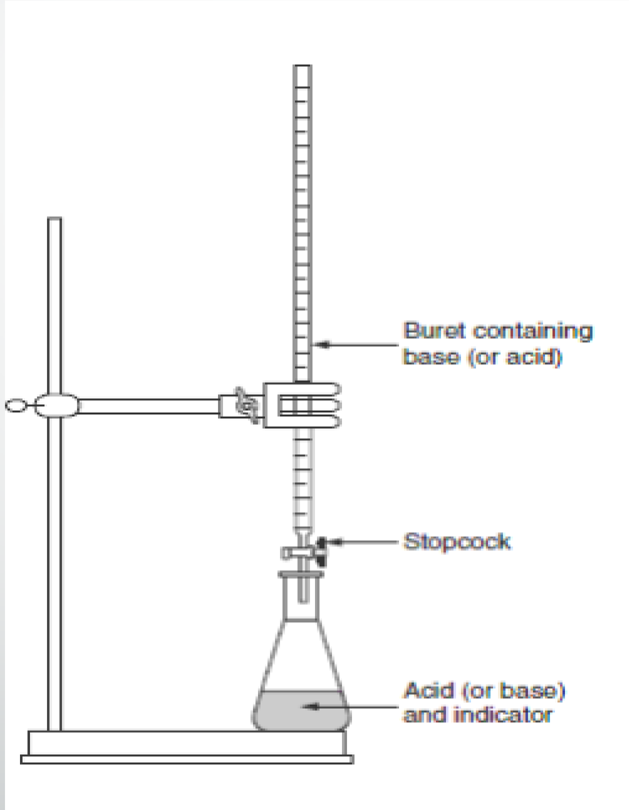
In **Titration** a solution of accurately known concentration **(TITRANT)** is added gradually to another solution of unknown concentration **(ANALYTE)** until the chemical reaction between the two solutions is complete.

Equivalence point

the point at which
the reaction is complete

Indicator

substance that changes
color at the
equivalence point



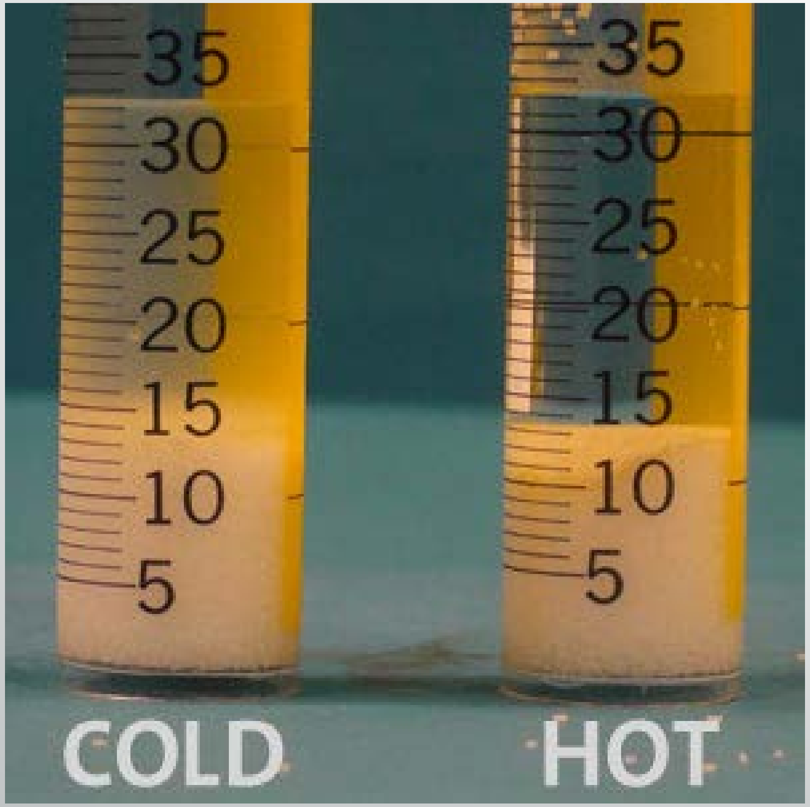
Factors Affecting Solubility

Temperature

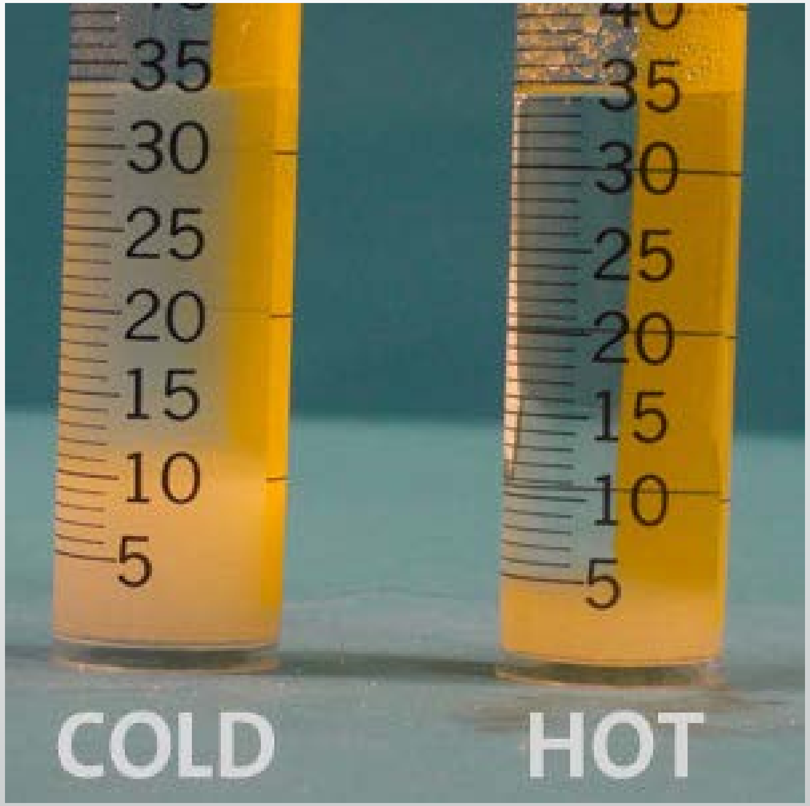
- ***<u>For solids</u>***: Temperature ↑ solubility ↑ Temperature increases kinetic energy of *solvent* particles therefore more solute can be dissolved
- ***<u>For gases</u>***: Temperature ↑ solubility ↓ Temperature increases the kinetic energy of *solute* particles therefore more particles ***escape*** from solution

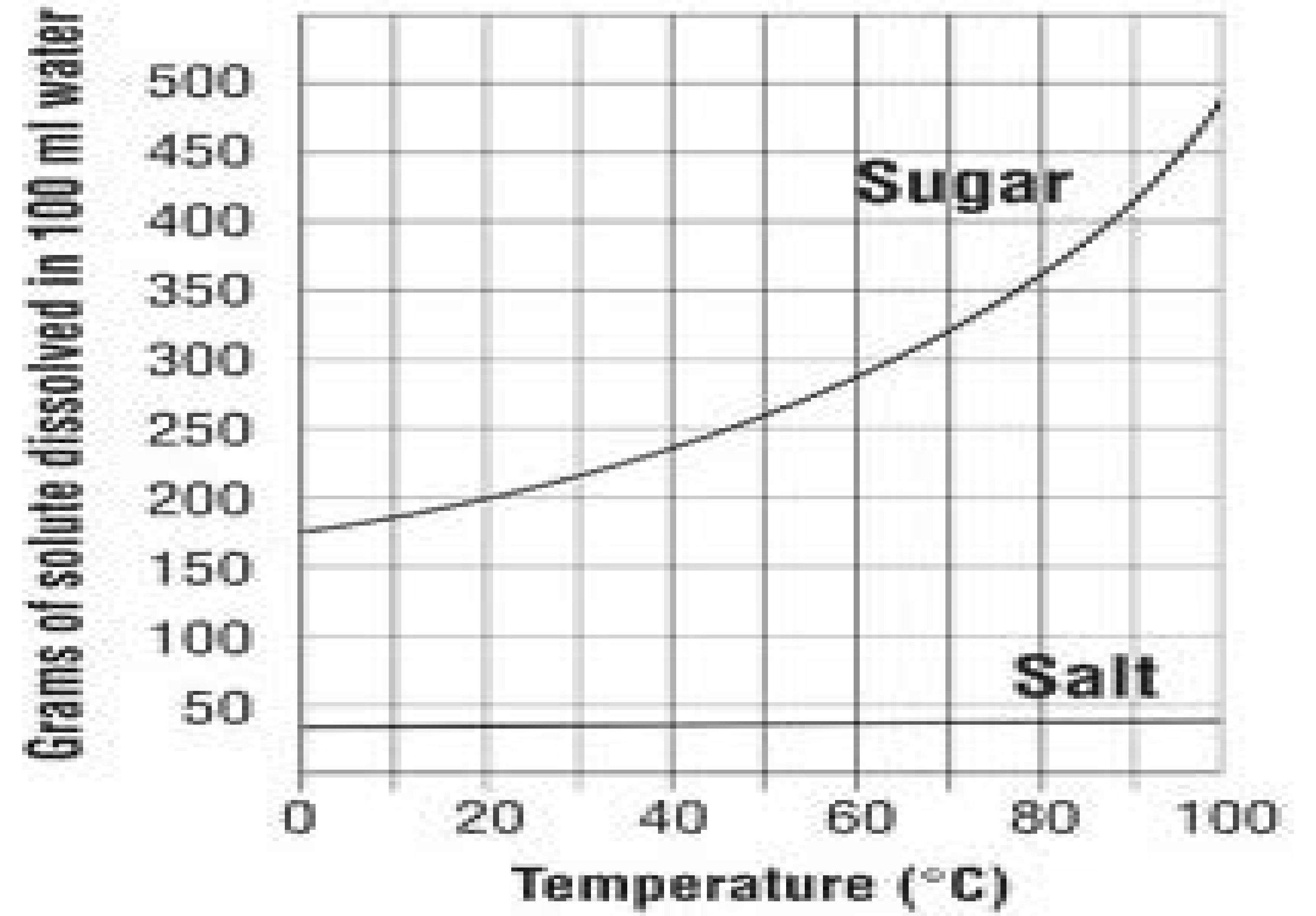
Solid Solubility & Temp

Salt

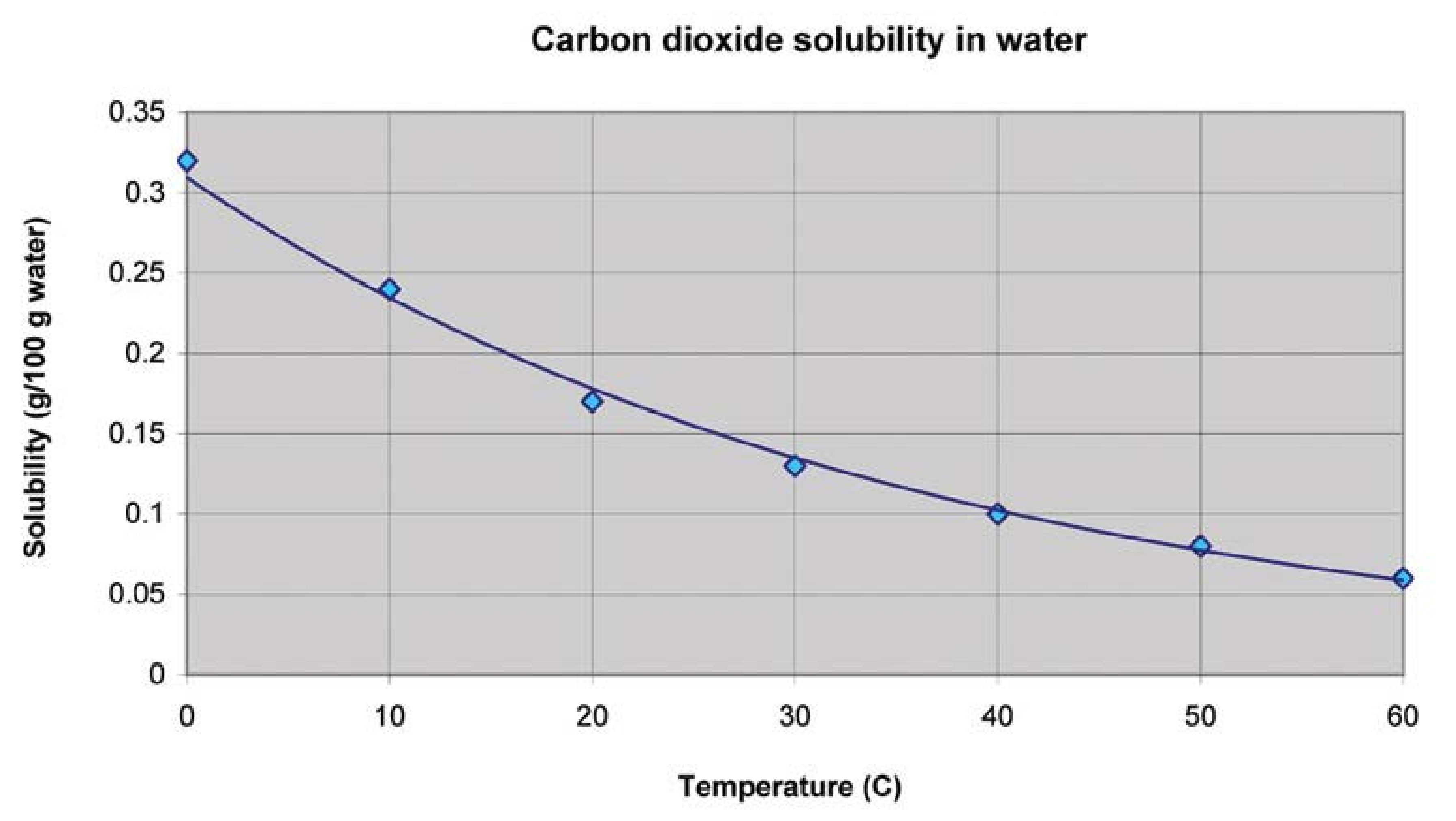


Sugar



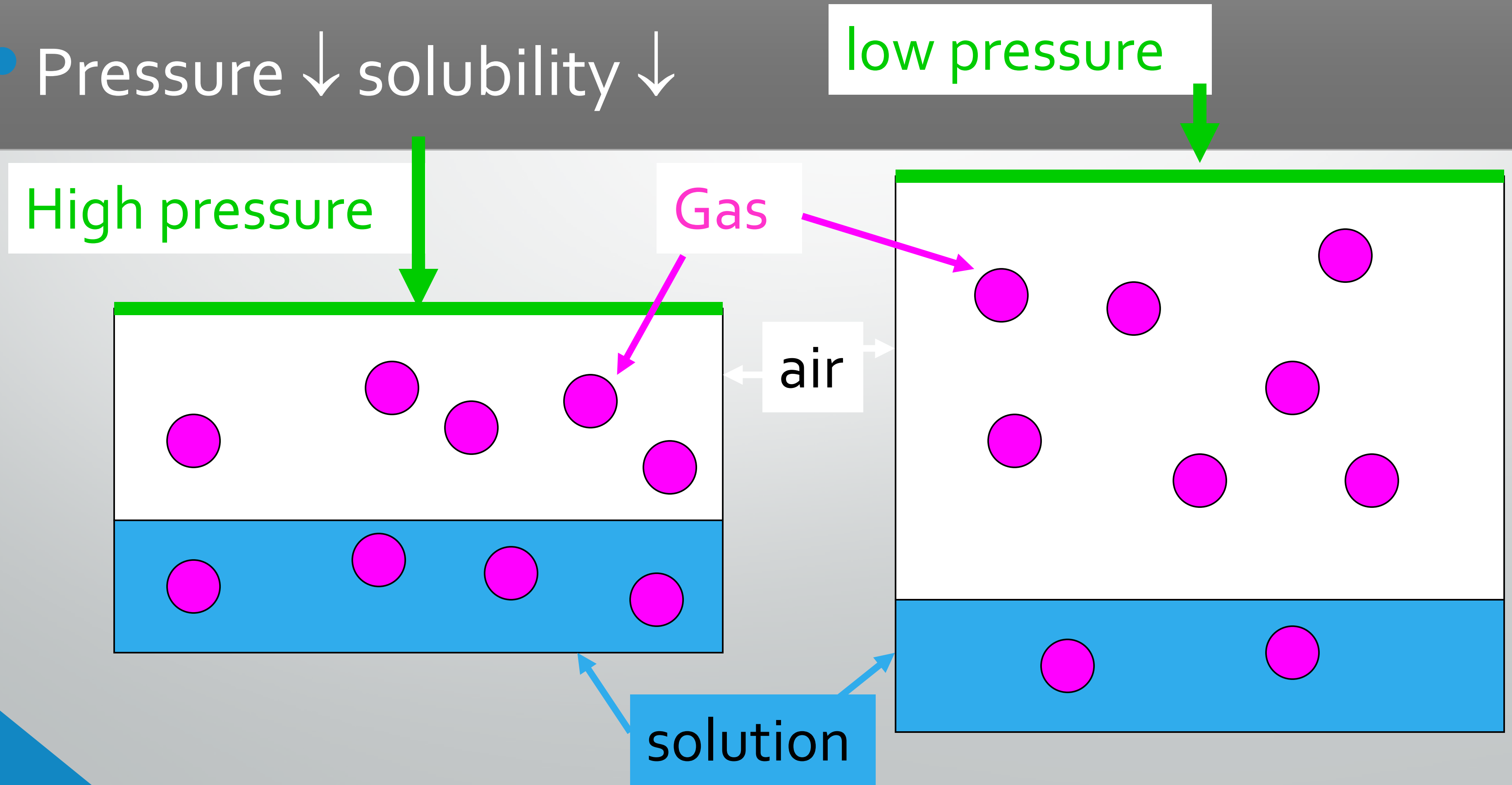

Grams of solute dissolved in 100 ml water
500
450
400
350
300
250
200
150
100
50
Sugar
Salt
0
20
40
60
80
100
Temperature (°C)

Gas Solubility & Temp

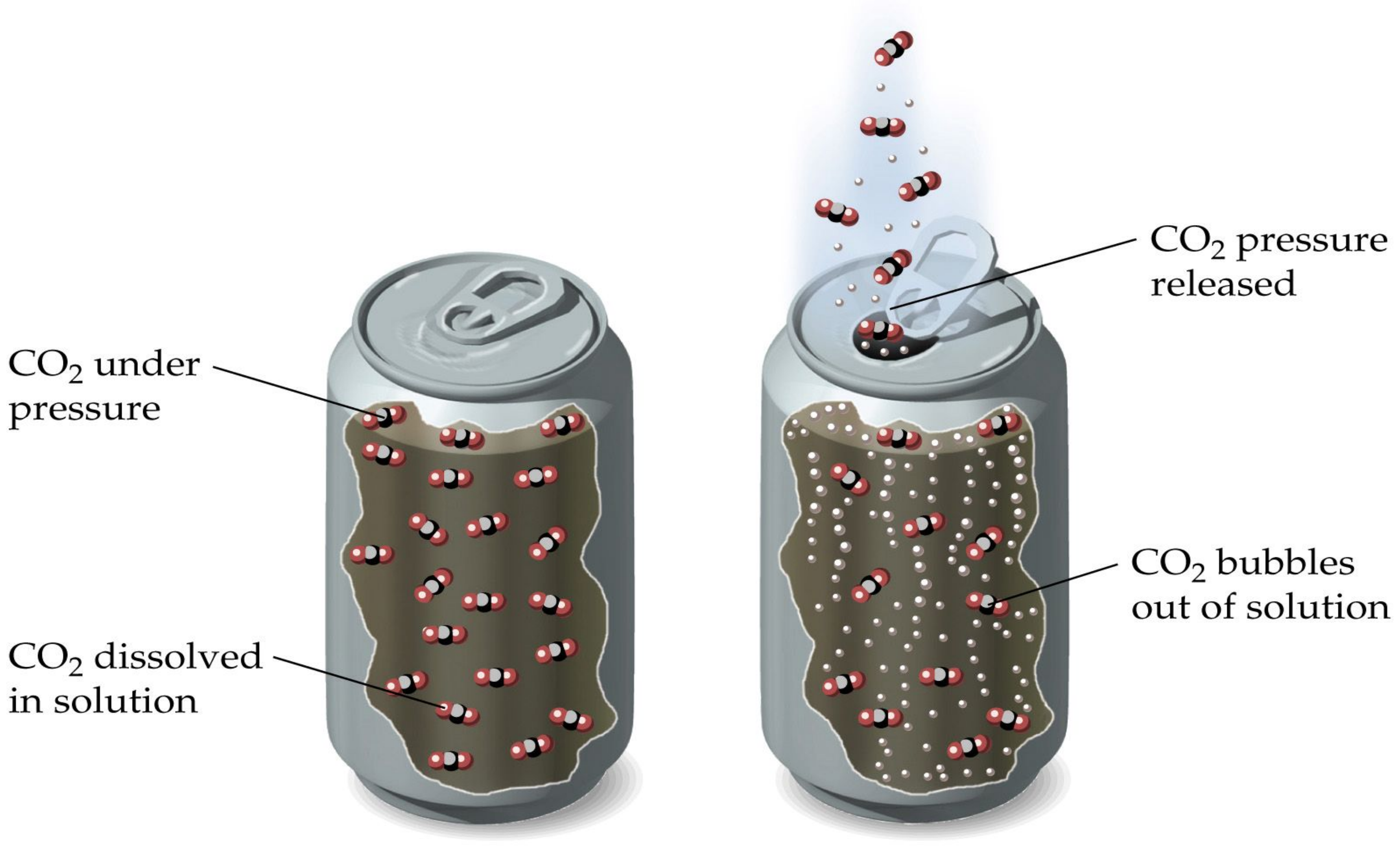


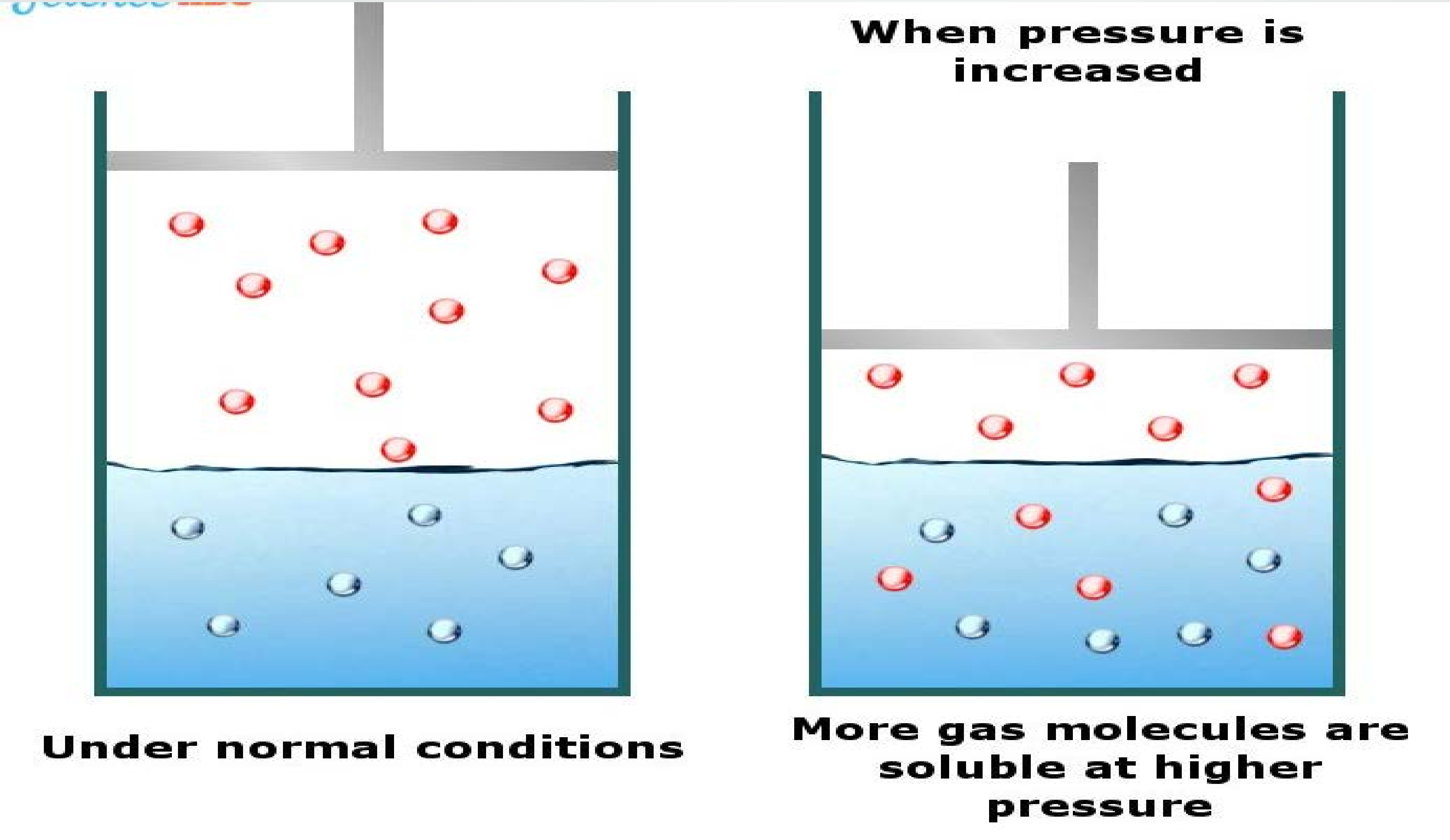
Pressure (Gasses Only!)

- Increasing pressure forces more gas into solution
- Pressure ↑ solubility ↑
- Pressure ↓ solubility ↓



Gas Solubility & Pressure





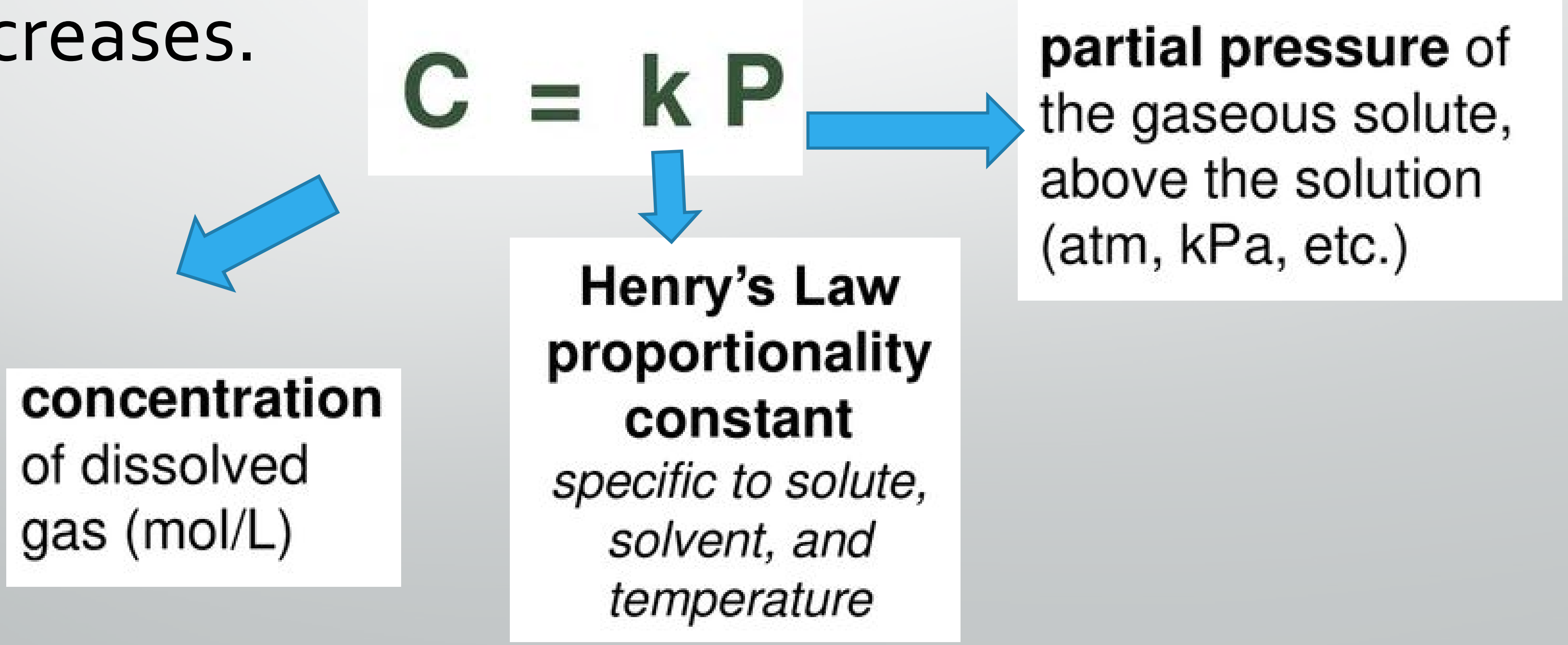
When pressure is increased

Under normal conditions

More gas molecules are soluble at higher pressure

Solubility and Pressure—Henry's Law

- ***Henry's law*** states that the solubility of a gas in a liquid is directly related to the pressure of the gas over the liquid. That is, the amount of gas that can dissolve in a liquid increases as the pressure of the gas in the space above the liquid increases.



$$C = kP$$

concentration of dissolved gas (mol/L)

Henry's Law proportionality constant *specific to solute, solvent, and temperature*

partial pressure of the gaseous solute, above the solution (atm, kPa, etc.)

What is the concentration of dissolved CO_2 in a 2-liter bottle of soda at 25°C if the partial pressure of CO_2 used in the bottling process is 3.20 atm. Assume the solubility of CO_2 in the soda is the same as in that of pure water. The Henry's law constant of CO_2 in water at 25°C is 3.1×10^{-2} M/atm.

$$C_g = kP_g$$

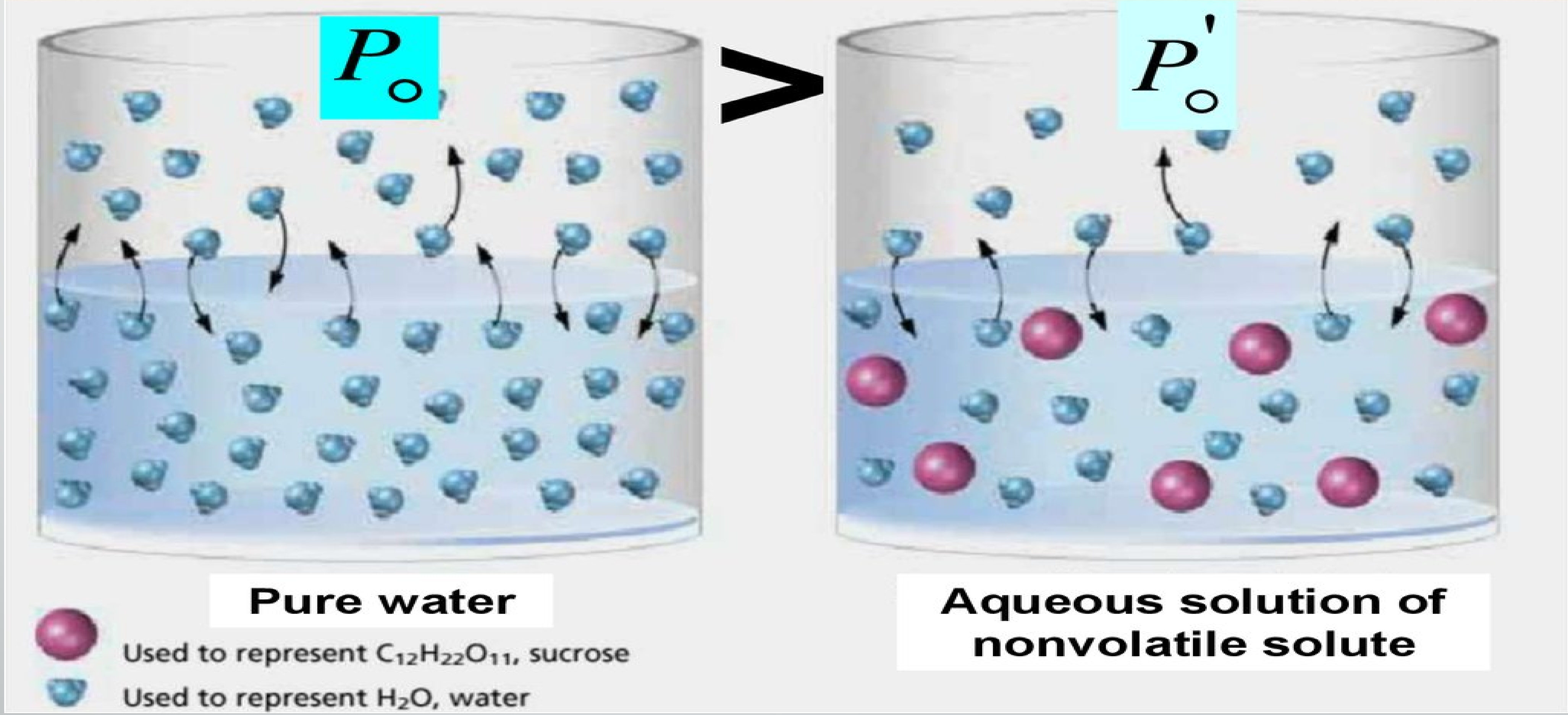
$$C_g = (3.1 \times 10^{-2}\ \text{M}/\cancel{\text{atm}})(3.20\ \cancel{\text{atm}})$$

Colligative Properties

Physical properties of solutions dictated by the number of solute particles present. Their chemical structures are not factors in determining these properties!

- vapor pressure lowering
- boiling point elevation
- freezing point depression
- osmotic pressure

VAPOR PRESSURE LOWERING



Quantitative Treatment of VP Lowering

Raoult's Law (vapor pressure of a solvent above a solution, P_{solvent})

$$P_{\text{solvent}} = \chi_{\text{solvent}} \times P^{o}_{\text{solvent}}$$

where P^{o}_{solvent} = vapor pressure of the pure solvent

How does the amount of solute affect the magnitude of the VP lowering?

(substitute 1- χ_{solute} for χ_{solvent} in the above equation and rearrange)

$$P^{o}_{\text{solvent}} - P_{\text{solvent}} = \Delta P = \chi_{\text{solute}} \times P^{o}_{\text{solvent}}$$

(change in VP is proportional to the *mole fraction* of solute)

Using Raoult's Law to find the vapor pressure lowering

PROBLEM: Calculate the vapor pressure lowering, ΔP, when 10.0 mL of glycerol ($C_3H_8O_3$) is added to 500. mL of water at 50. °C. At this temperature, the vapor pressure of pure water is 92.5 torr and its density is 0.988 g/mL. The density of glycerol is 1.26 g/mL.

PLAN: Find the mol fraction, χ, of glycerol in solution and multiply by the vapor pressure of water.

SOLUTION:

$$10.0 \text{ mL } C_3H_8O_3 \times \frac{1.26 \text{ g } C_3H_8O_3}{\text{mL } C_3H_8O_3} \times \frac{\text{mol } C_3H_8O_3}{92.09 \text{ g } C_3H_8O_3} = 0.137 \text{ mol } C_3H_8O_3$$

$$500.0 \text{ mL } H_2O \times \frac{0.988 \text{ g } H_2O}{\text{mL } H_2O} \times \frac{\text{mol } H_2O}{18.02 \text{ g } H_2O} = 27.4 \text{ mol } H_2O$$

$\chi = 0.00498$

$$\Delta P = \boxed{\frac{0.137 \text{ mol } C_3H_8O_3}{0.137 \text{ mol } C_3H_8O_3 + 27.4 \text{ mol } H_2O}} \times 92.5 \text{ torr} = 0.461 \text{ torr}$$

Boiling-Point Elevation

The boiling point of the solution is the temperature at which its vapor pressure
equal the external atmospheric pressure.

Because the presence of adding a nonvolatile solute **lowers** the vapor pressure
of solution.

It was found that the **decrease** in vapor pressure leads to increase in the boiling
Point

Ex.. Dissolve sugar in the water

Quantitative Treatment of BP Elevation

The magnitude of the effect is proportional to solute concentration.

$$\Delta T_{\mathbf{b}} = K_{\mathbf{b}} m$$

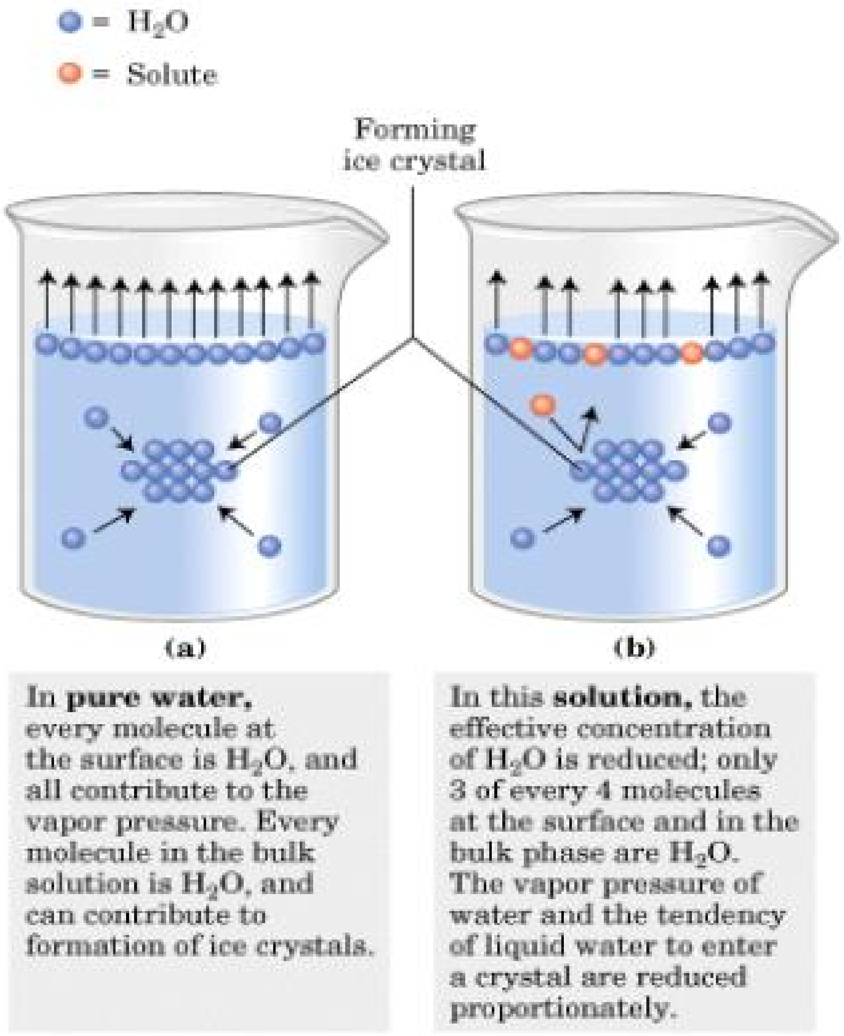
(m = solution molality, $K_{\mathbf{b}}$ = molal BP elevation constant, $\Delta T_{\mathbf{b}}$ = BP elevation)

$$\Delta T_{\mathbf{b}} = T_{\mathbf{b}}\,(\text{solution}) - T_{\mathbf{b}}\,(\text{solvent})$$

Freezing point depression and boiling point elevation: adding solute to a solvent lowers the freezing point and raises the boiling point of the solvent.

- Solute particles disrupt crystallization and evaporation.
- The change in freezing point or boiling point is directly proportional to the ***molality*** of solute particles.
- A solute that produces more ions in solution has the greatest effect.

Colligative properties = properties affected by the number, not the nature, of solute particles



In **pure water,** every molecule at the surface is H_2O, and all contribute to the vapor pressure. Every molecule in the bulk solution is H_2O, and can contribute to formation of ice crystals.

In this **solution,** the effective concentration of H_2O is reduced; only 3 of every 4 molecules at the surface and in the bulk phase are H_2O. The vapor pressure of water and the tendency of liquid water to enter a crystal are reduced proportionately.

For 1 mol solute/1 kg H_20
(=1 molal):
freezing point = -1.86 °C
melting point = +0.543 °C

Quantitative Treatment of FP Depression

$$\Delta T_f = K_f m$$

The magnitude of the effect is proportional to solute concentration.

(m = solution molality, K_f = molal FP depression constant , ΔT_f = FP depression)

$$\Delta T_f = T_f\text{(solvent)} - T_f\text{(solution)}$$

Molal Boiling Point Elevation and Freezing Point Depression Constants of Several Solvents

solvent	boiling point (°C)*	K_b (°C/m)	melting point (°C)	K_f (°C/m)
acetic acid	117.9	3.07	16.6	3.90
benzene	80.1	2.53	5.5	4.90
carbon disulfide	46.2	2.34	-111.5	3.83
carbon tetrachloride	76.5	5.03	-23	30.
chloroform	61.7	3.63	-63.5	4.70
diethyl ether	34.5	2.02	-116.2	1.79
ethanol	78.5	1.22	-117.3	1.99
water	100.0	0.512	0.0	1.86

*at 1 atm.

Determining the boiling point elevation and freezing point depression of a solution

PROBLEM: You add 1.00 kg of ethylene glycol antifreeze ($C_2H_6O_2$) to your car radiator, which contains 4450 g of water. What are the boiling and freezing points of the resulting solution?

PLAN: Find the number of mols of ethylene glycol and *m* of the solution; multiply by the boiling or freezing point constant; add or subtract, respectively, the changes from the boiling point and freezing point of water.

SOLUTION:

$$1.00 \times 10^3 \text{ g } C_2H_6O_2 \times \frac{\text{mol } C_2H_6O_2}{62.07 \text{ g } C_2H_6O_2} = 16.1 \text{ mol } C_2H_6O_2$$

$$\frac{16.1 \text{ mol } C_2H_6O_2}{4.450 \text{ kg } H_2O} = 3.62\ m\ C_2H_6O_2$$

$$\Delta T_b = 0.512\ ^{o}C/m \times 3.62\ m = 1.85\ ^{o}C$$

$$\Delta T_f = 1.86\ ^{o}C/m \times 3.62\ m$$

BP = 101.85 °C

FP = -6.73 °C

What is the freezing point of a solution containing 478 g of ethylene glycol (antifreeze) in 3202 g of water? The molar mass of ethylene glycol is 62.01 g/mol.

$\Delta T_f = K_f\, m$ $\quad$ K_f water = 1.86 ^{0}C/m

$$\boldsymbol{m} = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}} = \frac{478\text{ g} \times \dfrac{1\text{ mol}}{62.01\text{ g}}}{3.202\text{ kg solvent}} = 2.41\, m$$

$$\Delta T_f = K_f\, m = 1.86\ ^0\text{C}/m \times 2.41\, m = 4.48\ ^0\text{C}$$

$$\Delta T_f = T_f^0 - T_f$$

$$T_f = T_f^0 - \Delta T_f = 0.00\ ^0\text{C} - 4.48\ ^0\text{C} = -4.48\ ^0\text{C}$$

Osmotic Pressure

- ❐ Applies only to aqueous solutions!

- ❐ Two solutions of different concentrations are separated by a semi-permeable membrane (allows water but not solute to pass through)

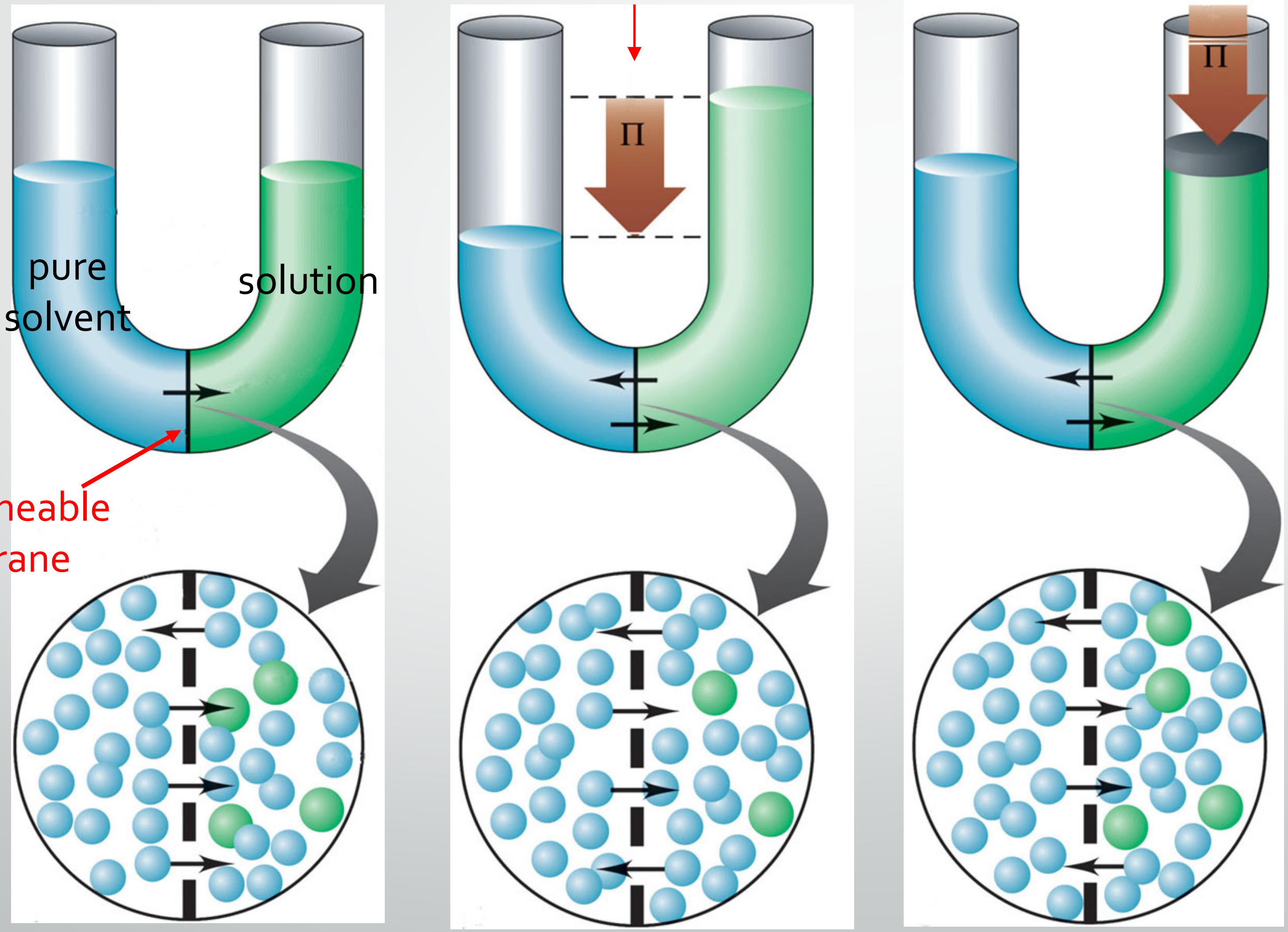
Osmotic Pressure (π)

Osmosis is the selective passage of solvent molecules through a porous membrane from a dilute solution to a more concentrated one.

A ***semipermeable membrane*** allows the passage of solvent molecules but blocks the passage of solute molecules.

Osmotic pressure (π) is the pressure required to stop osmosis.

The development of osmotic pressure



osmotic pressure

pure solvent

solution

semipermeable membrane

applied pressure needed to prevent volume increase; <u>equal to the osmotic pressure</u>

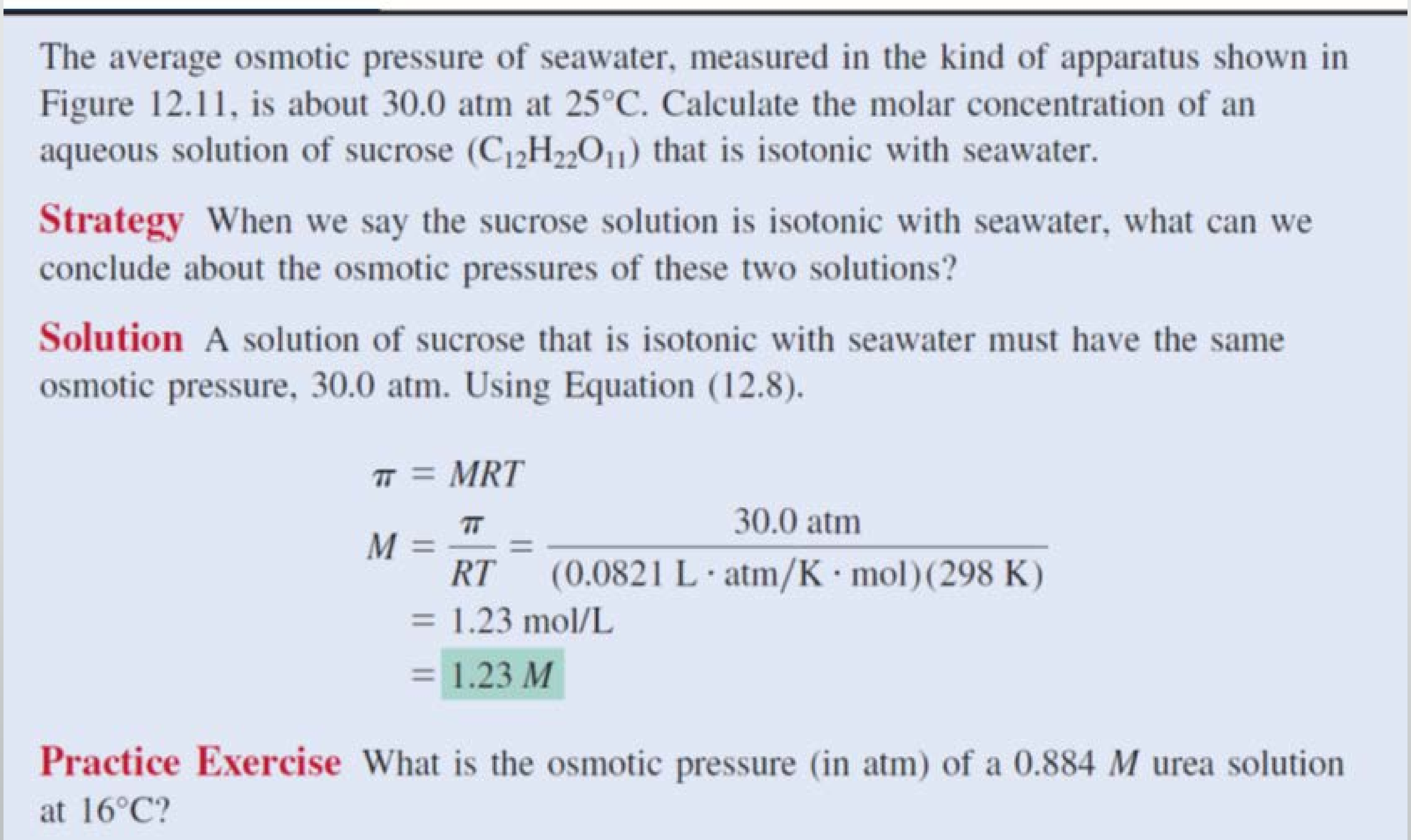
OP is proportional to the number of solute particles in a given *volume* of solution (to *M*).

$$\boxed{\pi = MRT}$$

M is the molarity of the solution

R is the gas constant (0.0821 L . atm/K . mol)

T is the temperature (in K)



The average osmotic pressure of seawater, measured in the kind of apparatus shown in Figure 12.11, is about 30.0 atm at 25°C. Calculate the molar concentration of an aqueous solution of sucrose ($C_{12}H_{22}O_{11}$) that is isotonic with seawater.

Strategy When we say the sucrose solution is isotonic with seawater, what can we conclude about the osmotic pressures of these two solutions?

Solution A solution of sucrose that is isotonic with seawater must have the same osmotic pressure, 30.0 atm. Using Equation (12.8).

$$\begin{aligned} \pi &= MRT \\ M &= \frac{\pi}{RT} = \frac{30.0\ \text{atm}}{(0.0821\ \text{L}\cdot\text{atm/K}\cdot\text{mol})(298\ \text{K})} \\ &= 1.23\ \text{mol/L} \\ &= \boxed{1.23\ M} \end{aligned}$$

Practice Exercise What is the osmotic pressure (in atm) of a 0.884 *M* urea solution at 16°C?

Colligative Properties of Electrolyte Solutions

0.1 *m* NaCl solution $\longrightarrow$ 0.1 *m* Na^+ ions & 0.1 *m* Cl^- ions

Colligative properties are properties that depend only on the **number** of solute particles in solution and not on the **nature** of the solute particles.

0.1 *m* NaCl solution $\longrightarrow$ 0.2 *m* ions in solution

$$\textbf{Van't Hoff factor (i)} = \frac{\text{actual number of particles in soln after dissociation}}{\text{number of formula units initially dissolved in soln}}$$

	i should be
	1
NaCl	2
$CaCl_2$	3

Colligative Properties of Electrolyte Solutions for the other properties

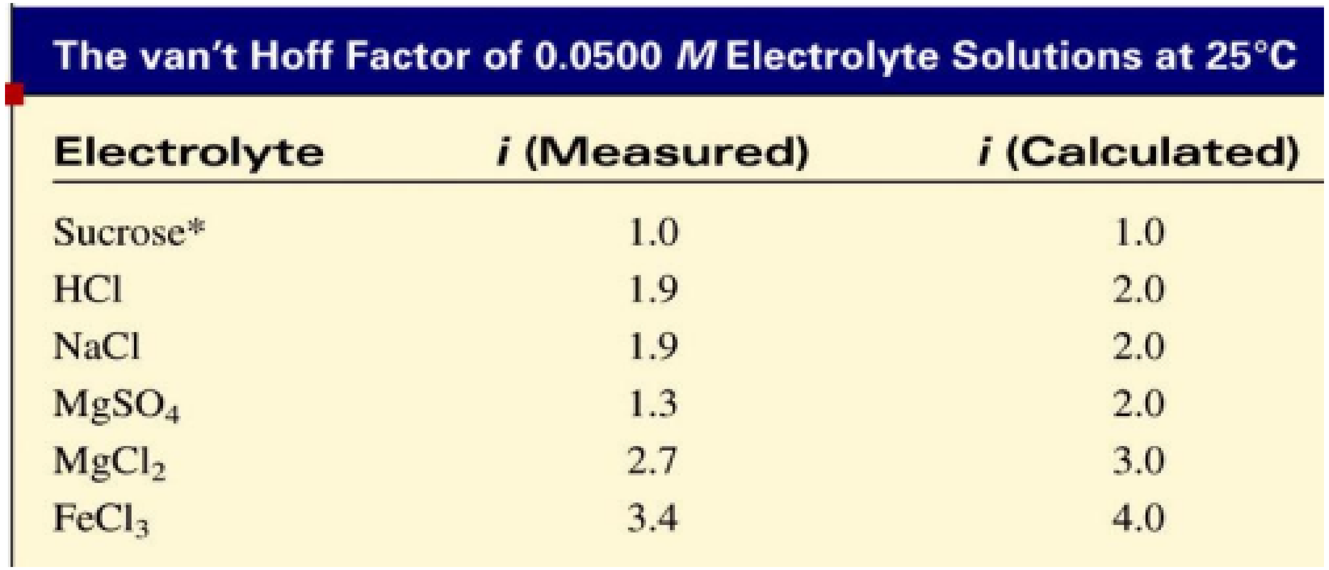
Boiling-Point Elevation $\Delta T_b = i\,K_b\,m$

Freezing-Point Depression $\Delta T_f = i\,K_f\,m$

Osmotic Pressure (p) $\pi = iMRT$

The solution of a rarely dissociate completely ionic compound usually contains fewer particles than the formula of the compound would suggest the actual extent of dissociation can be expressed as a Van't Hoff factor (i)

Van't Hoff factor (i) = moles of particles in solution / moles of solute dissolved.



The van't Hoff Factor of 0.0500 *M* Electrolyte Solutions at 25°C

Electrolyte	*i* (Measured)	*i* (Calculated)
Sucrose*	1.0	1.0
HCl	1.9	2.0
NaCl	1.9	2.0
$MgSO_4$	1.3	2.0
$MgCl_2$	2.7	3.0
$FeCl_3$	3.4	4.0

DIFFERENT EXAPLES

Freezing Point Depression of electrplyte solution

At what temperature will a 5.4 molal solution of NaCl freeze?

$K_f = 1.86\ ^{o}C/molal$

Solution

$$\Delta T_{FP} = K_f \cdot m \cdot i$$

$$\Delta T_{FP} = (1.86\ ^{o}C/molal) \cdot 5.4\ m \cdot 2$$

$$\Delta T_{FP} = 20.1\ ^{o}C$$

$$FP = 0 – 20.1 = -20.1\ ^{o}C$$

Osmotic pressure of electrplyte solution

Calculate the osmotic pressure of 0.05 M MgSO4 at 25^0C if you know that van`t Hoff factor is 1.3

R=0.0821 L/atm.K

$$\pi = iMRT$$

$$\pi = 1.3 \times 0.05 \times 0.082 \times 298$$

$$= 1.6\ atm$$