Ch13

Solution

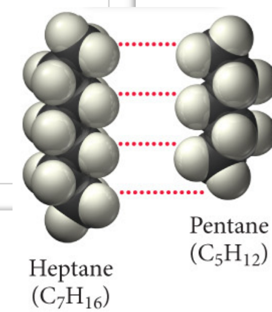
How and why homogenous matter forms.

The give and take between potential energy and entropy.

version 1.6

Ch13

Solution



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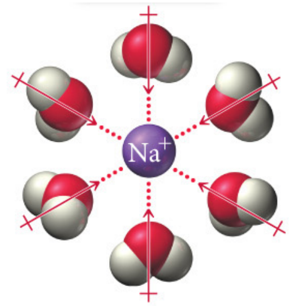

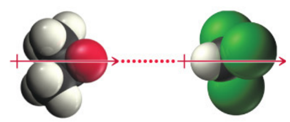

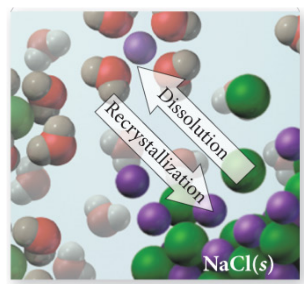

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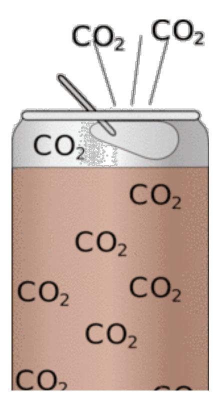

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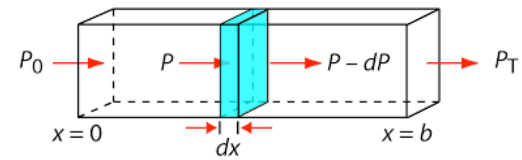

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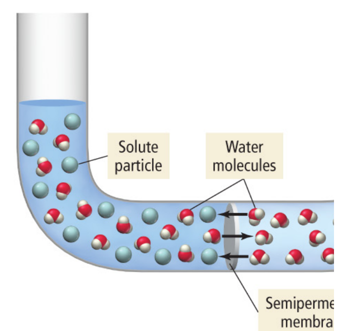
$P_A = \chi_A \bullet P^\circ$



Solutions

- Solutions are homogenous mixtures.
- Mixtures can be liquids, gas, or solid.
- We're going to discuss the structure of mixtures.
- How substances come into mixtures and how substances can be driven out of mixtures.
- How substances in mixtures interact.
  - … and how that interaction facilitates chemical reaction between the mixtures components.

A solution is a homogenous mixture.

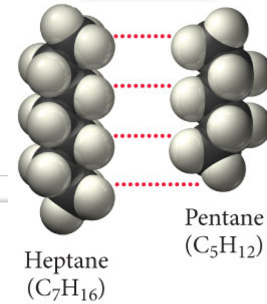
A solvent is the largest component of the mixture.

A solute is a smaller components of the mixture.

Types of Solution

Solution Phase	Solute Phase	Solvent Phase	Example
Gaseous solution	Gas	Gas	Air (mainly oxygen and nitrogen)
Liquid solution	Gas	Liquid	Club soda (CO_2 and water)
	Liquid	Liquid	Vodka (ethanol and water)
	Solid	Liquid	Seawater (salt and water)
Solid solution	Solid	Solid	Brass (copper and zinc) and other alloys

Solution



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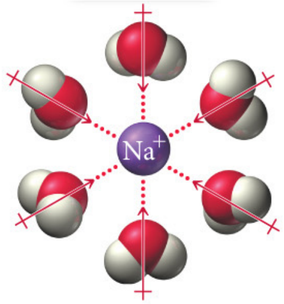

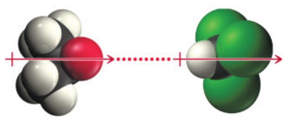

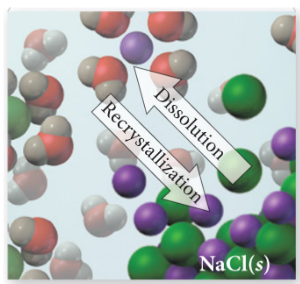

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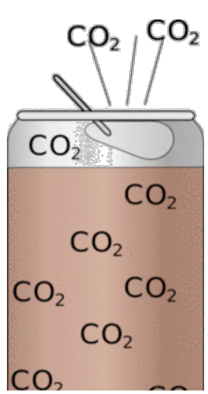

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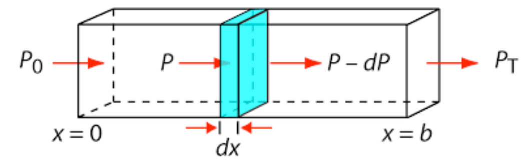

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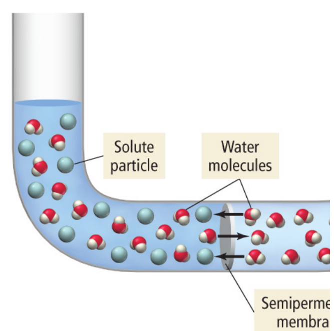
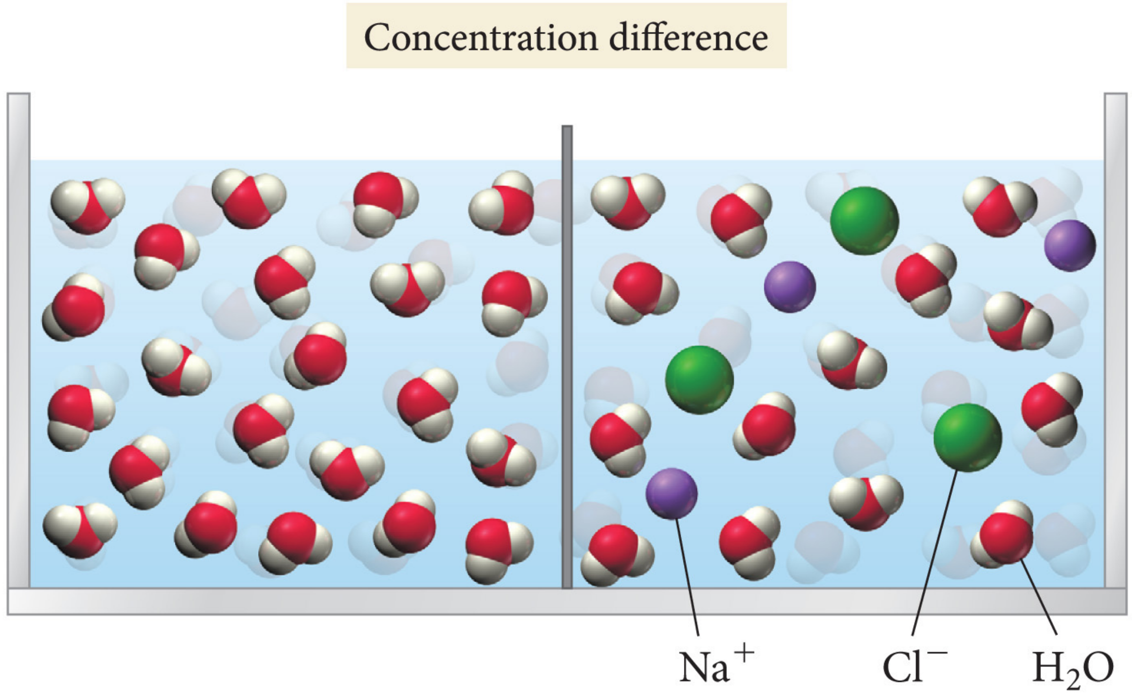
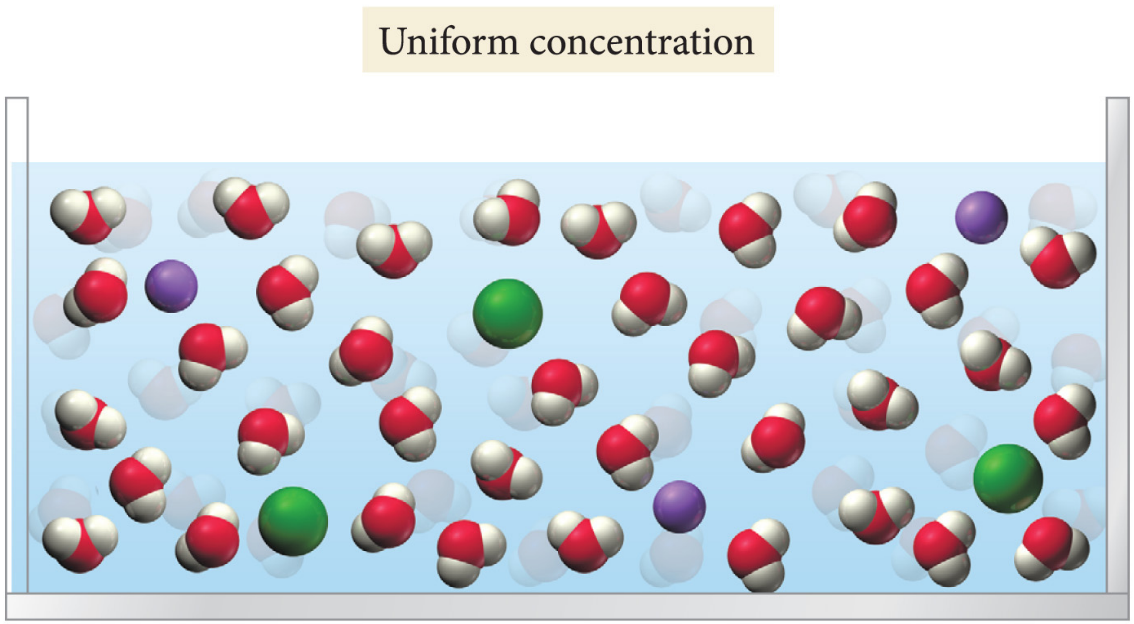


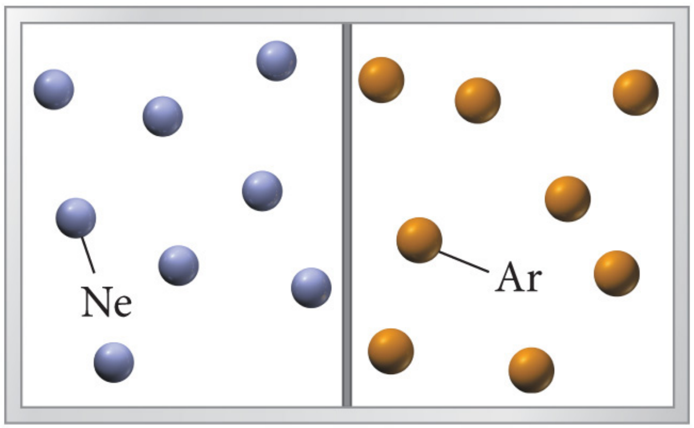

Concentration difference
Na^+
Cl^-
H_2O

(a)

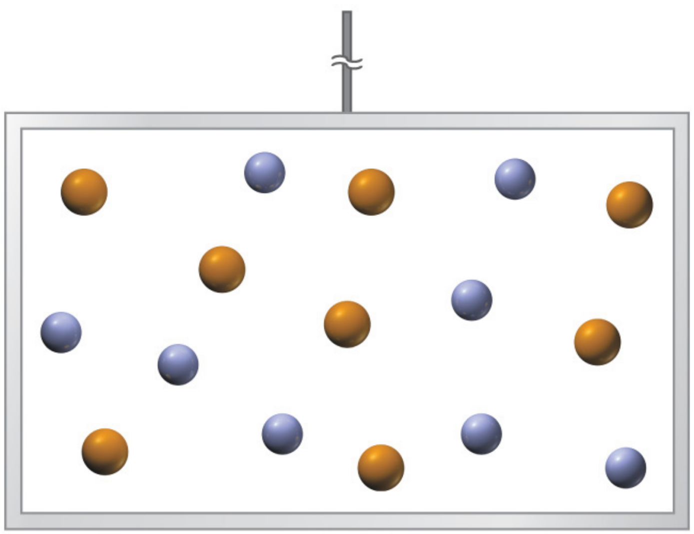

Uniform concentration

(b)

Why does an ideal gas mix?



(a)



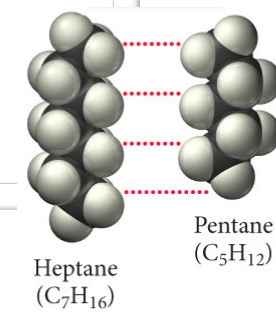
(b)

An ideal gas assumes completely elastic collisions.

Mixing results in no change of energy.

Ch13

Solution



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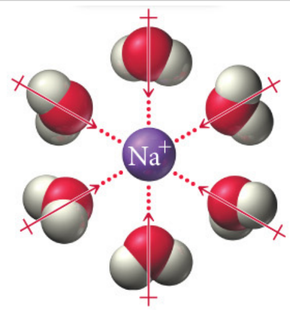

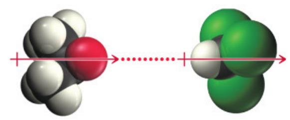

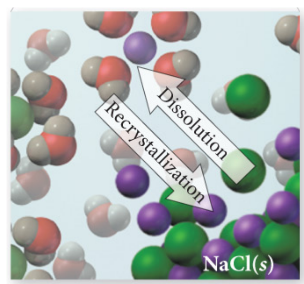

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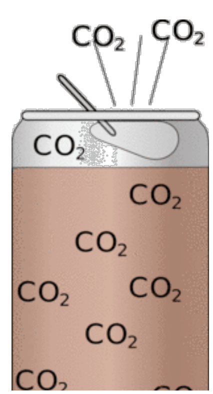

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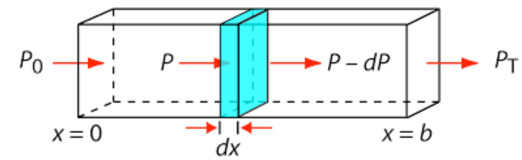

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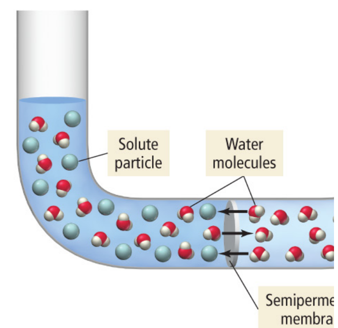
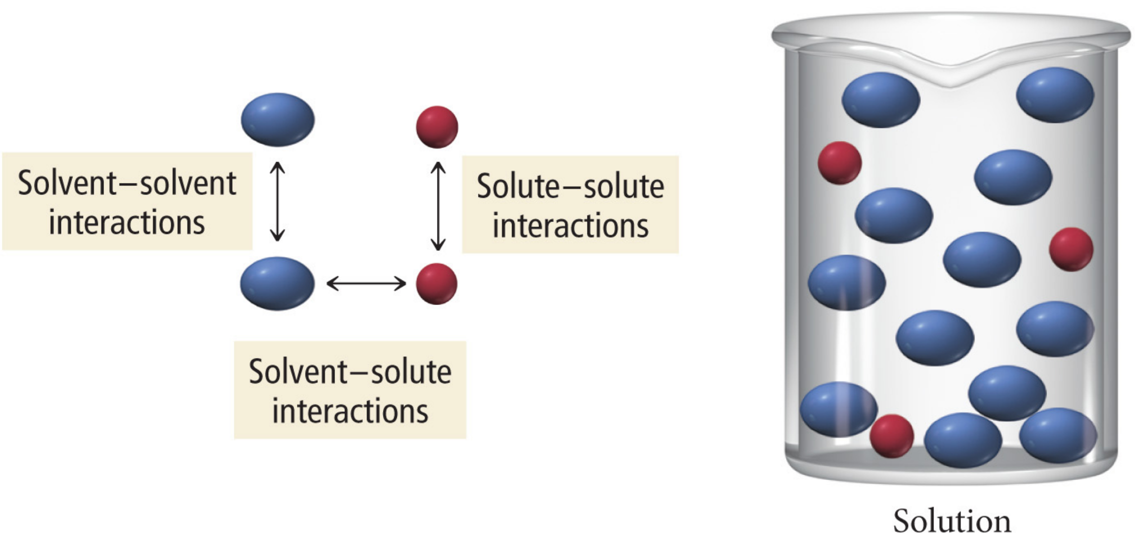
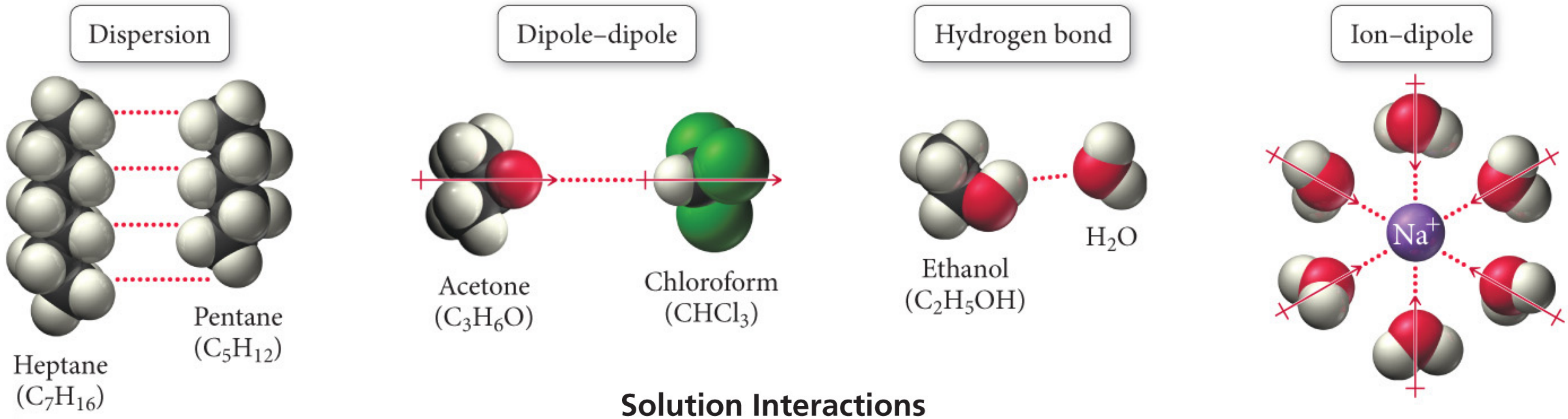


Solution Interactions



Intermolecular Forces

These forces may contribute to or oppose the formation of a solution.



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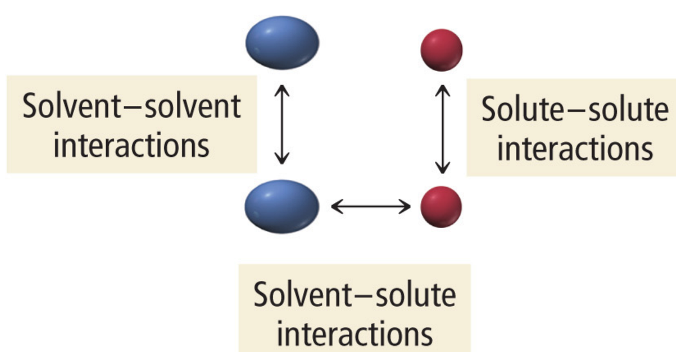
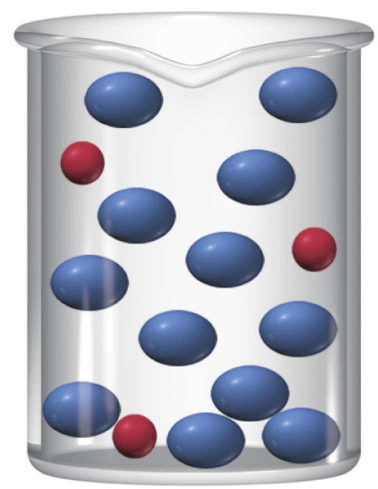
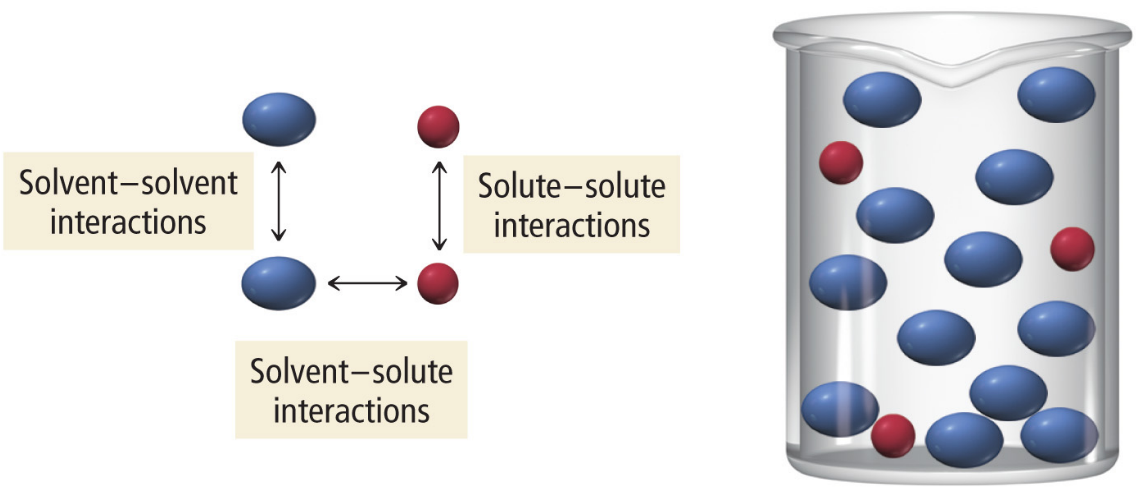




Solution

Solvent–solute interactions	>	Solvent–solvent and solute–solute interactions	Solution forms
Solvent–solute interactions	=	Solvent–solvent and solute–solute interactions	Solution forms
Solvent–solute interactions	<	Solvent–solvent and solute–solute interactions	Solution may or may not form, depending on relative disparity

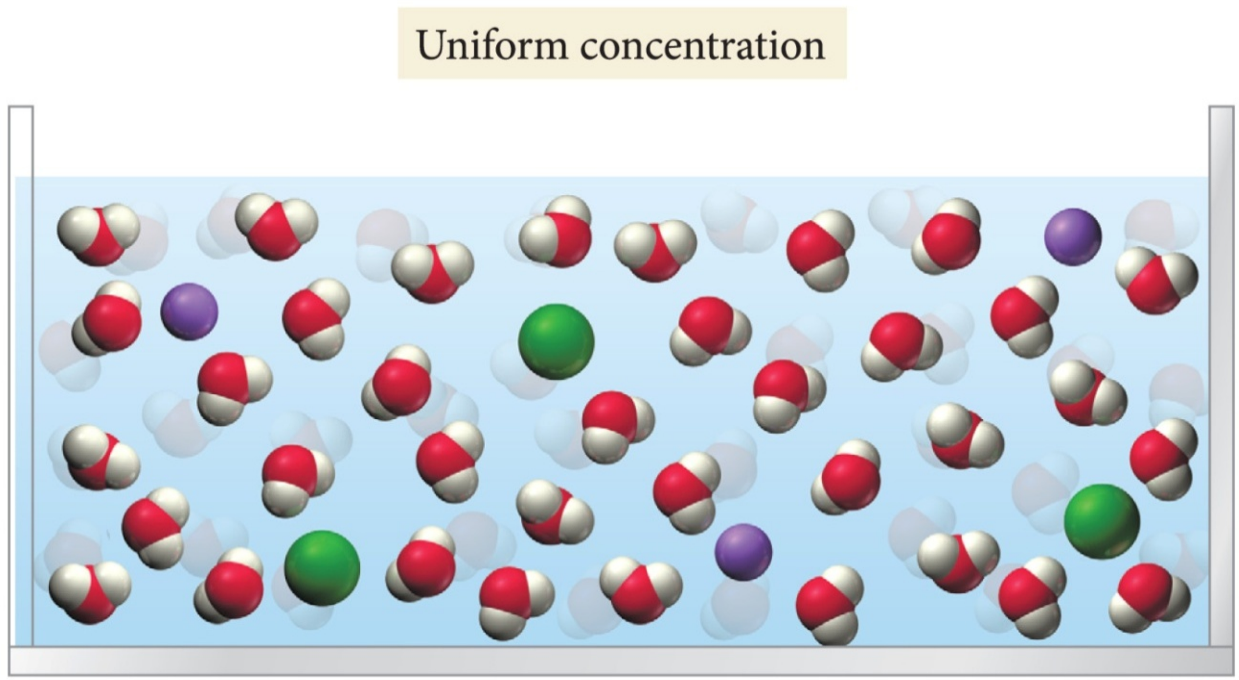
Solution Interactions



Solution

Which of the following pairs is likely to form a homogeneous mixture?

a) LiI and Hg
b) C_6H_{14} and Br_2
c) CH_3OH and C_6H_6
d) CCl_4 and H_2O
e) All pairs above are miscible.



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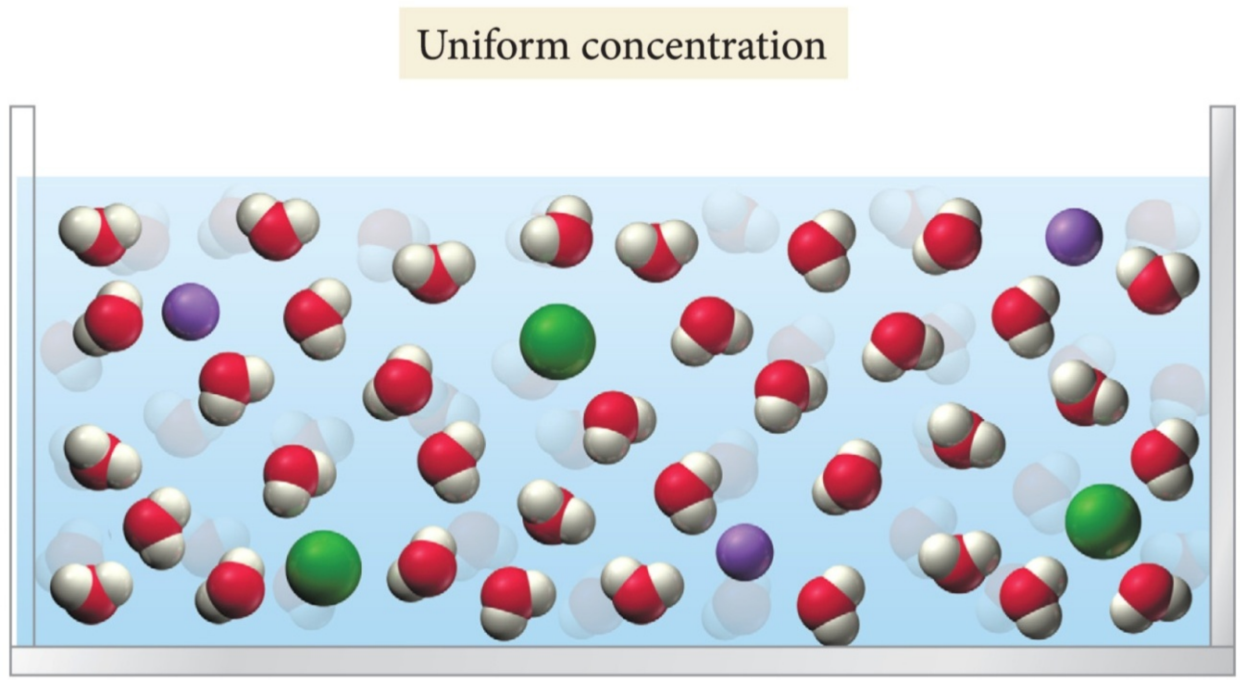
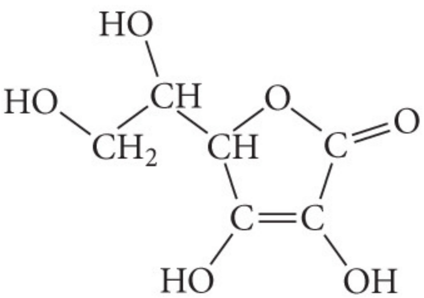
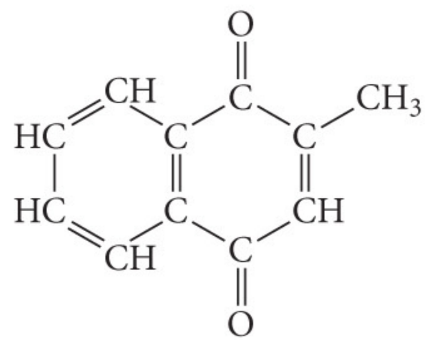


Which would be water soluble?

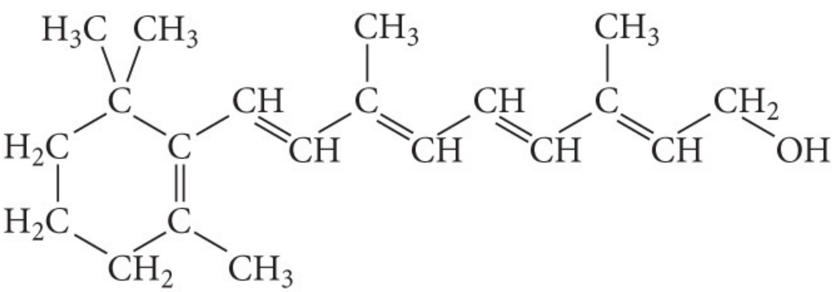
(a) Vitamin C



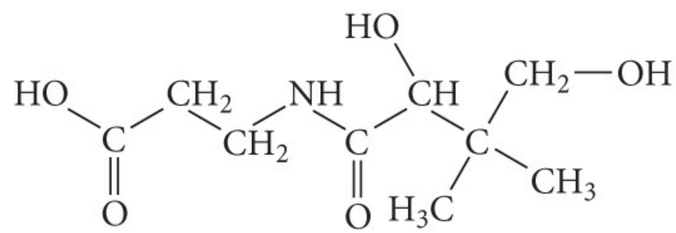
(b) Vitamin K_3



(c) Vitamin A



(d) Vitamin B_5



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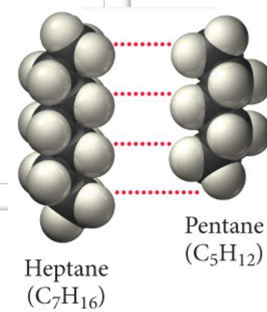

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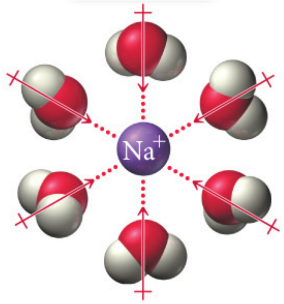

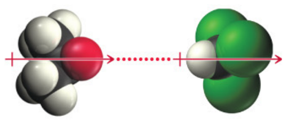

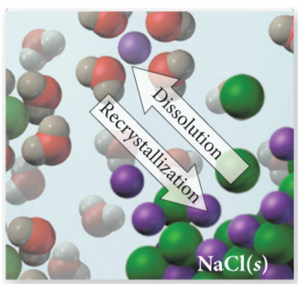

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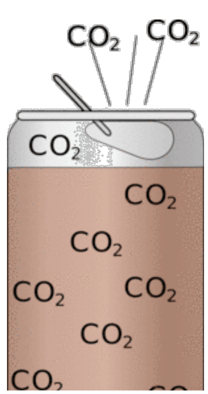

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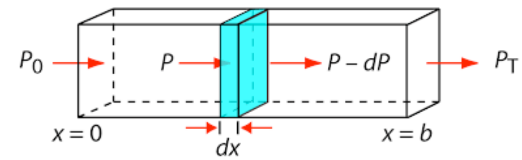

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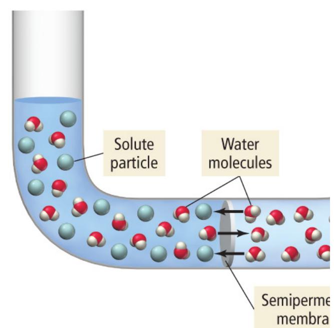
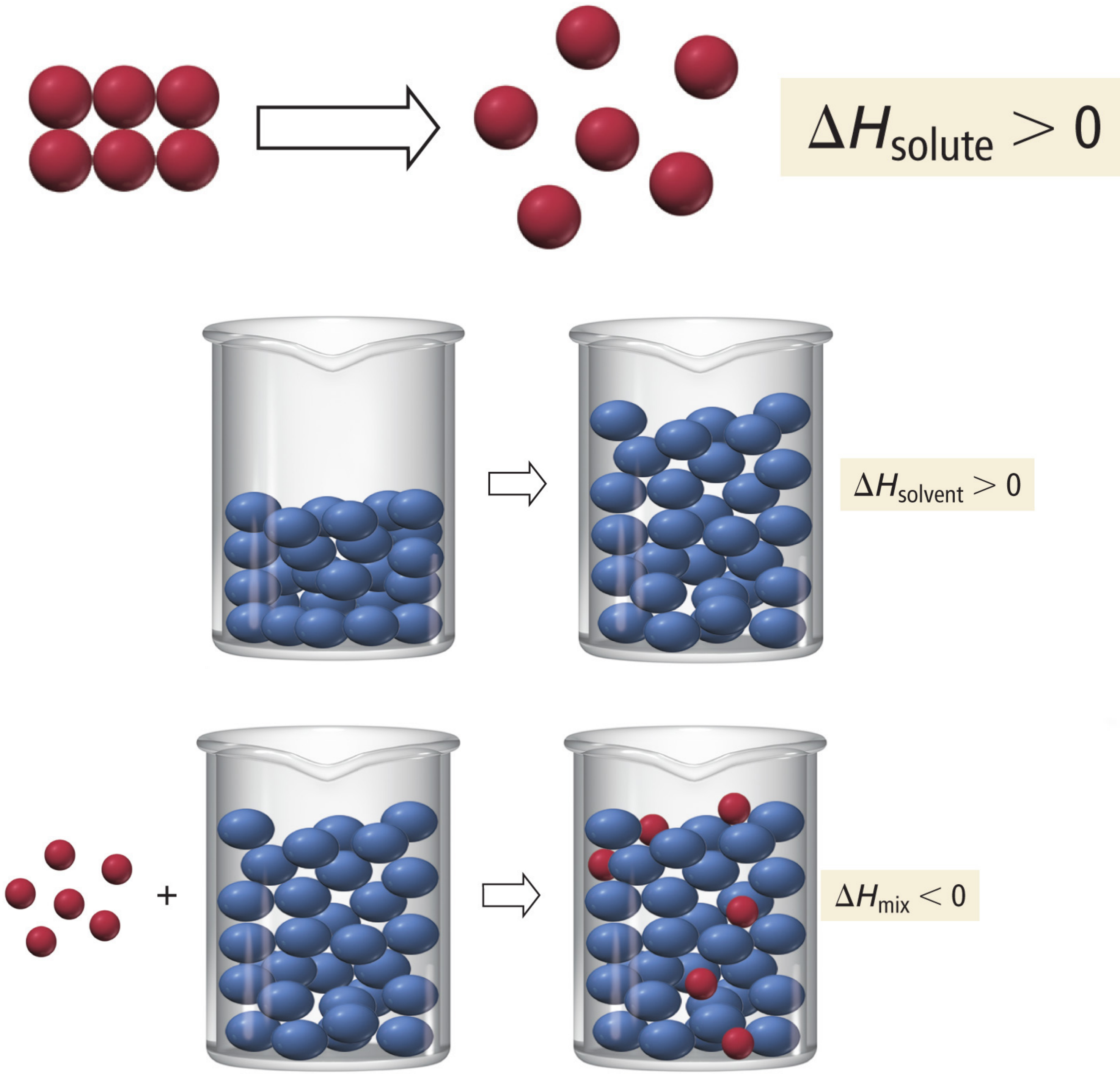
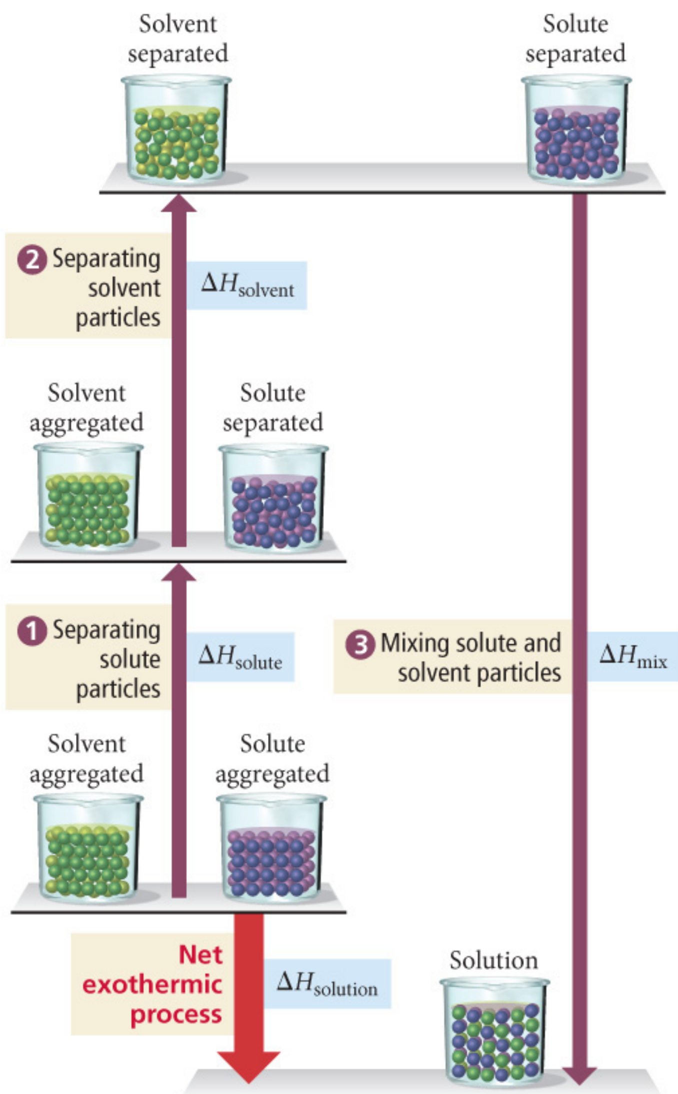



$\Delta H_{solute} > 0$
$\Delta H_{solvent} > 0$
+
$\Delta H_{mix} < 0$

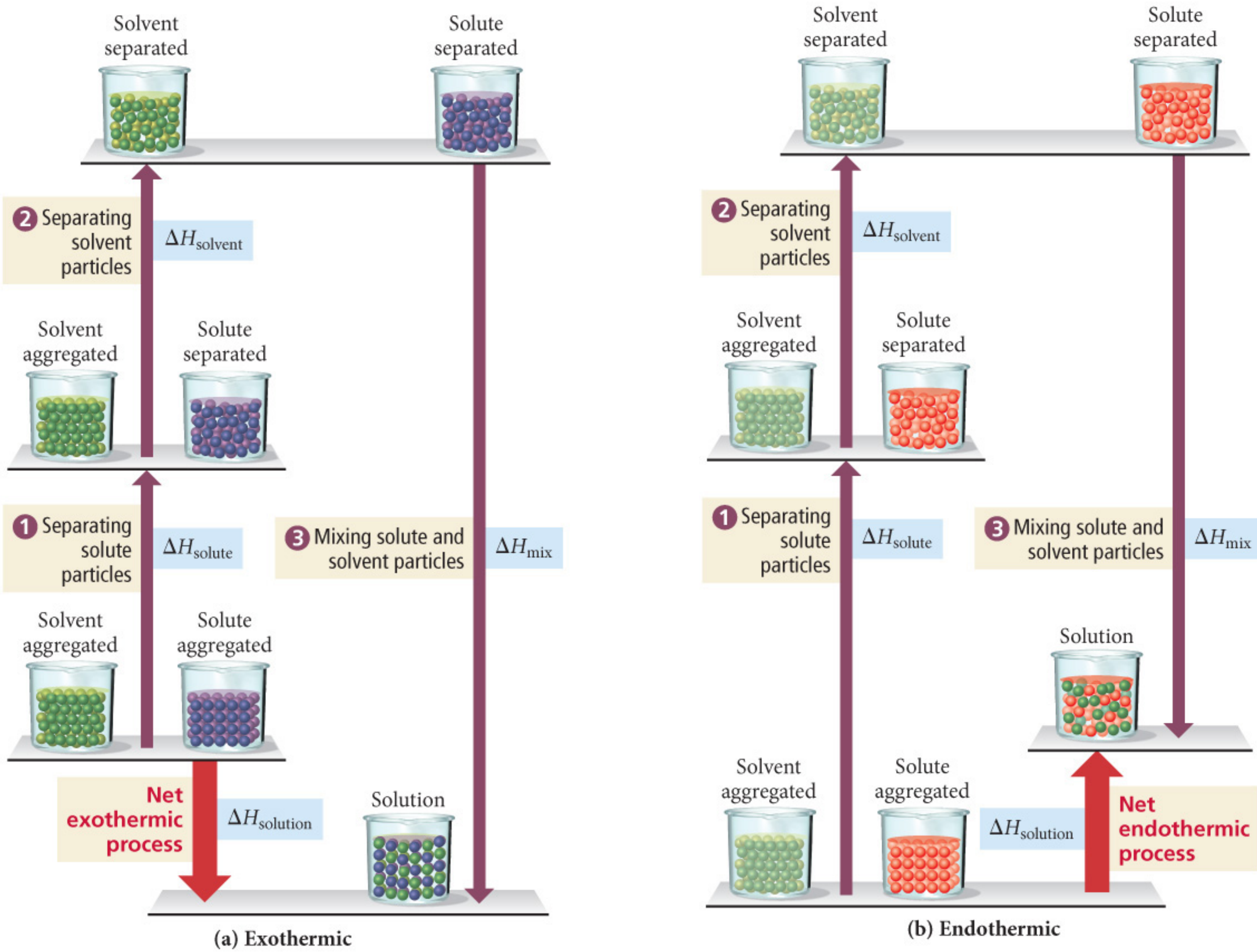
Heat of Solution



(a) Exothermic

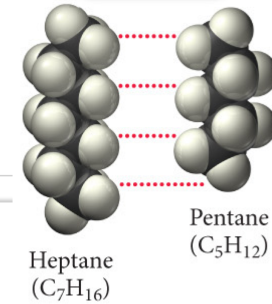
- Combining the heats of solvent, solute and mixing produces the next change in energy.
- Heat of solution, the sum of these heats, is exothermic if the mixing is thermodynamically favored.

Heat of Solution may be Endothermic



Ch13

Solution



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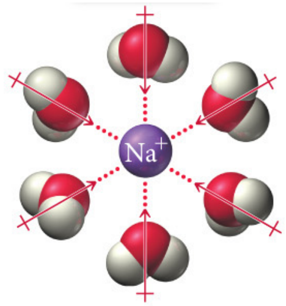

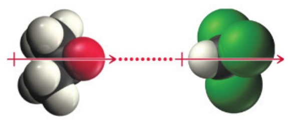

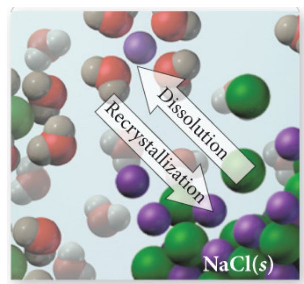

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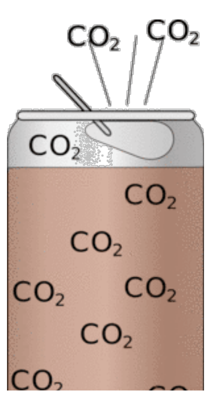

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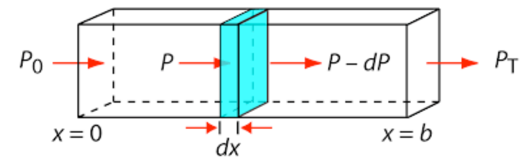

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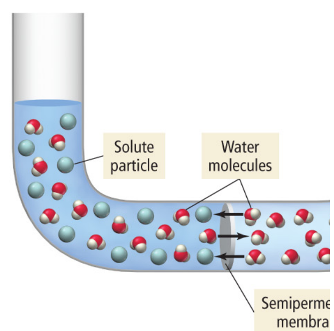
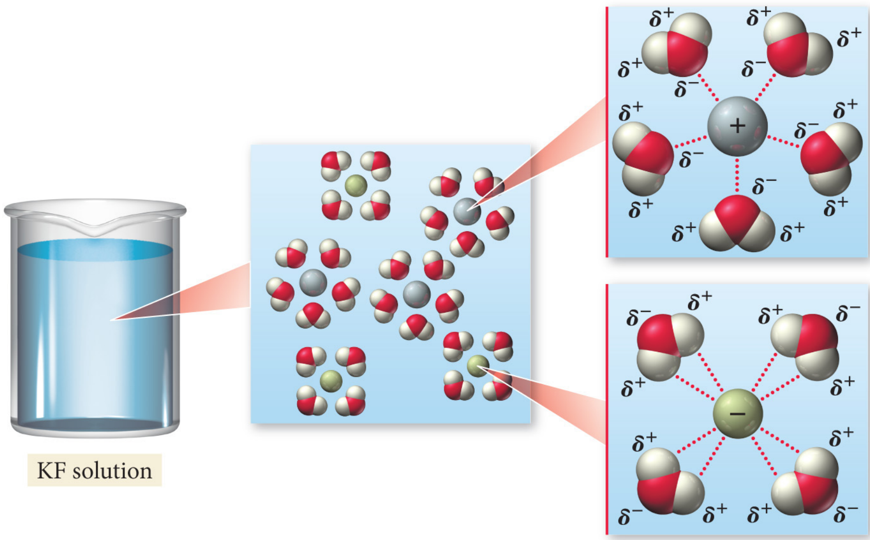
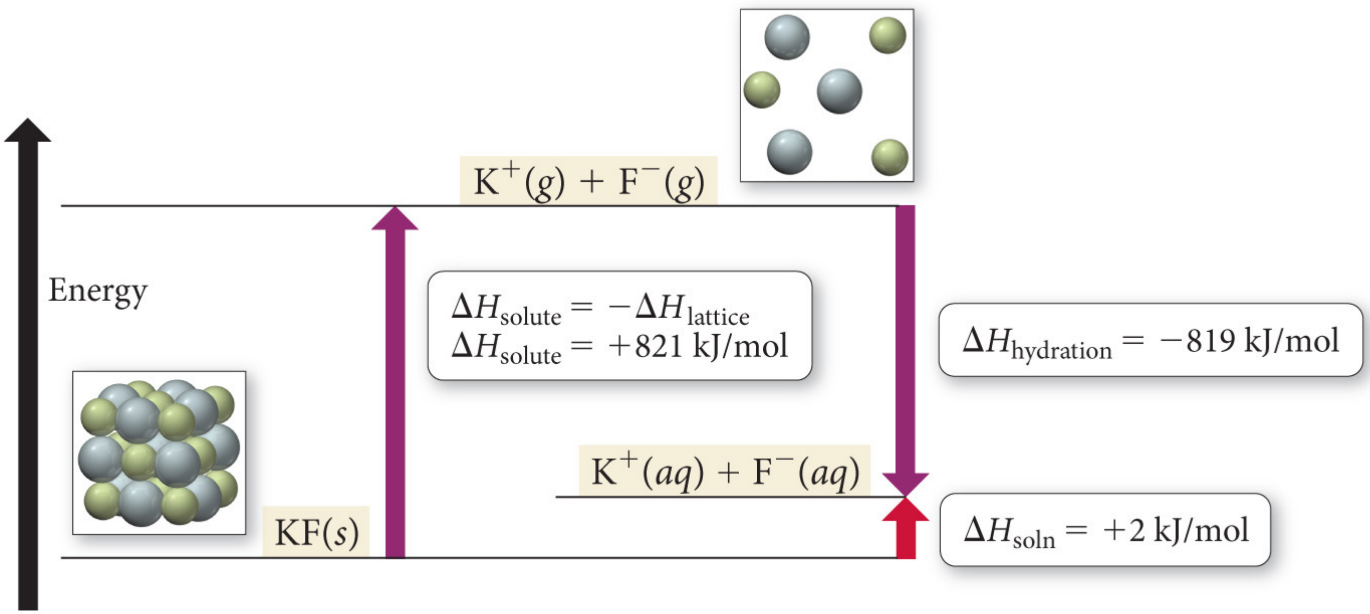


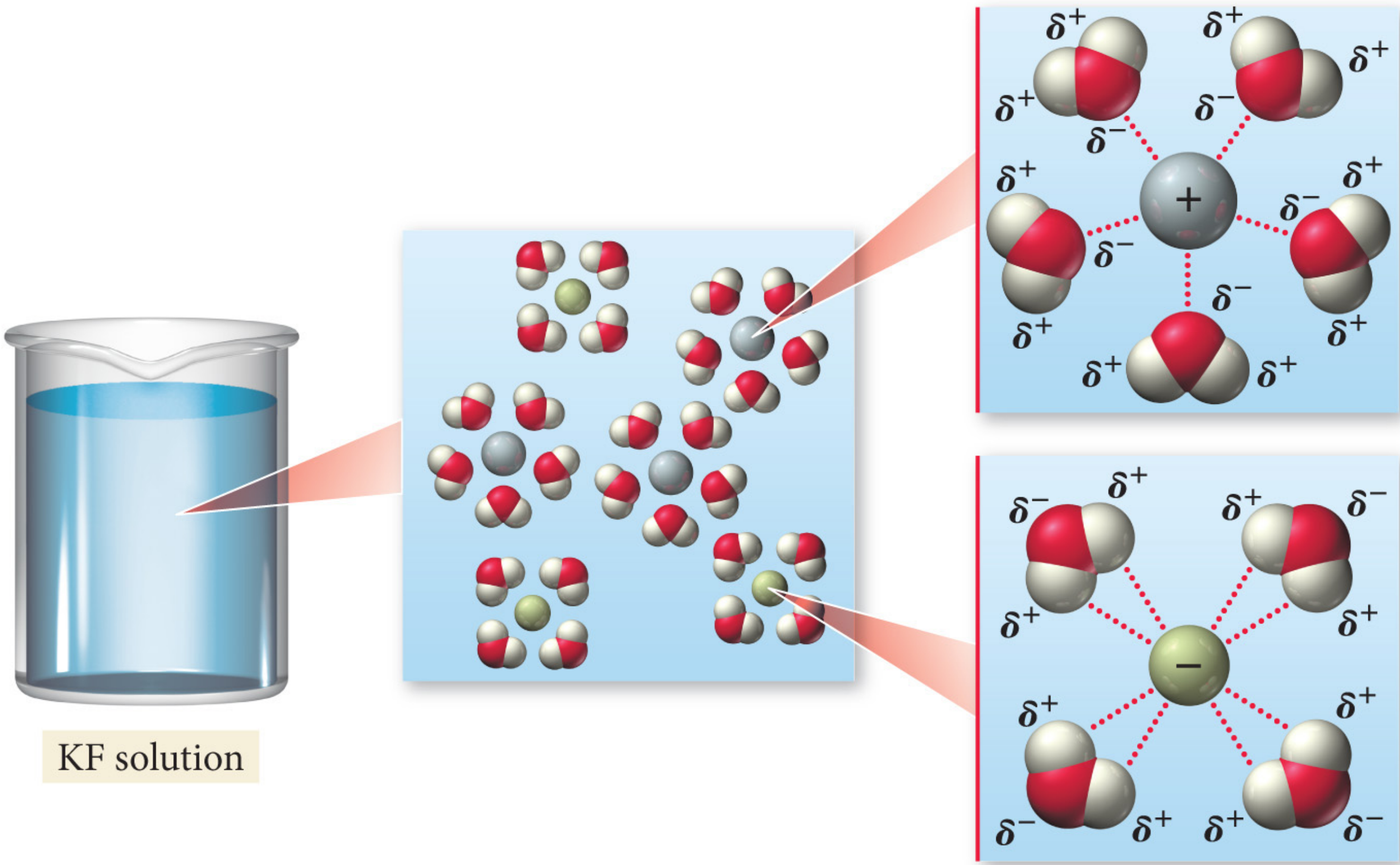
Heat of Hydration

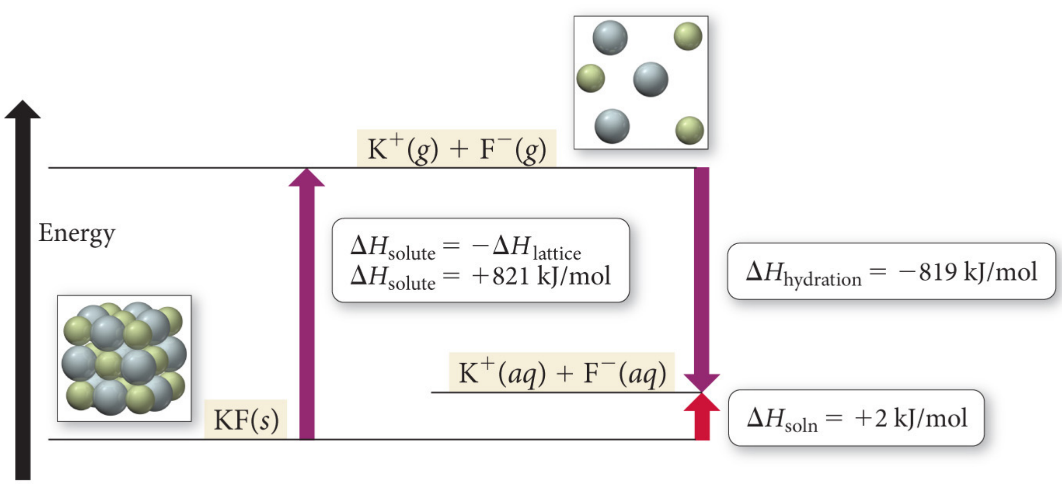


- When the solvent is water, we combine the heats of solvent and mixing into a single term, heat of hydration.
- For ionic compounds dissolved in water, Heat of hydration is usually exothermic.
- The ion-dipole forces produced are much stronger than the hydrogen bonding forces in waters heat of solvent.



Ion–Dipole Interactions





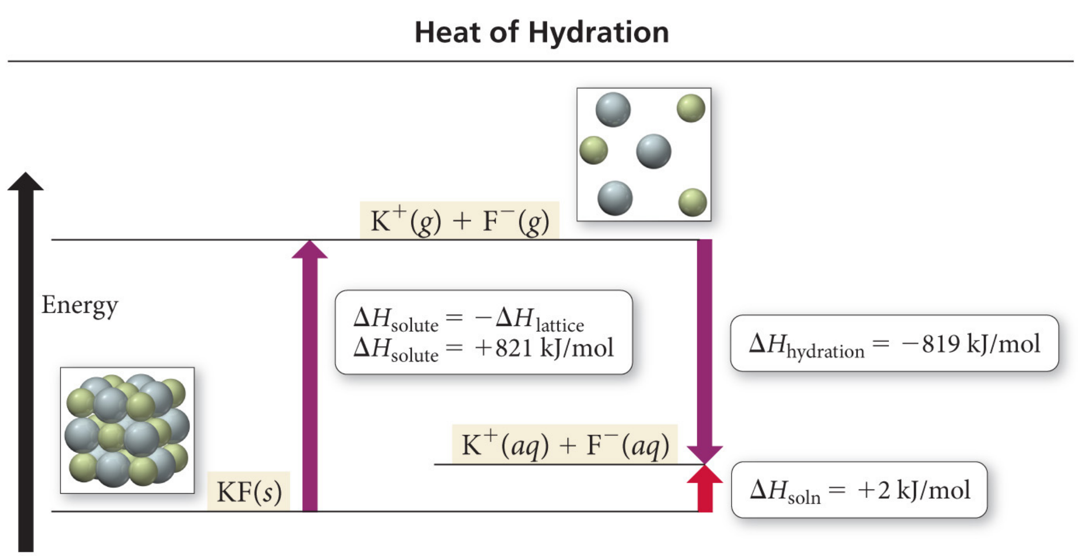
Heat of Hydration

$$\Delta H_{\text{soln}} = \Delta H_{\text{solute}} + \underbrace{\Delta H_{\text{solvent}} + \Delta H_{\text{mix}}}$$

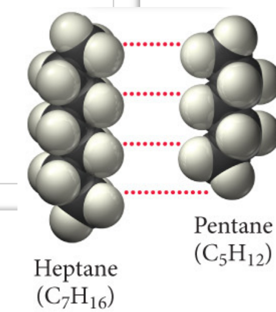
$$\Delta H_{\text{soln}} = \Delta H_{\text{solute}} + \Delta H_{\text{hydration}}$$

ΔH_{solute}: endothermic (positive)

$\Delta H_{\text{hydration}}$: exothermic (negative)



Solution



- Solution Structure
  - Definition & Types
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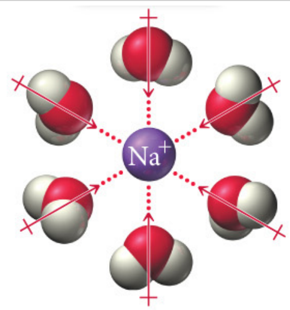
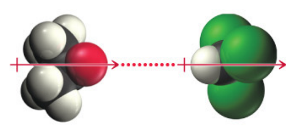
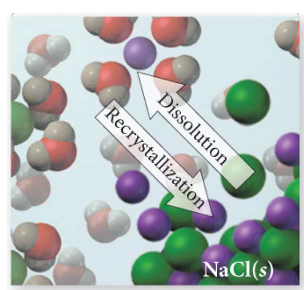
  - Entropy (drives it)
  - IMFs (allow it)
  - Energetics (enthalpy changes)
    - Aqueous Solutions ($\Delta H_{\text{hydration}}$)

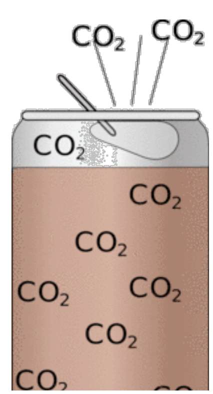




- Solubility Equilibrium
  - Dynamic Equilibrium
  - Saturation
    - Super Saturation
  - Control Factors
    - Temperature
    - Pressure



Henry's Law
$[A] = k_A P_A$



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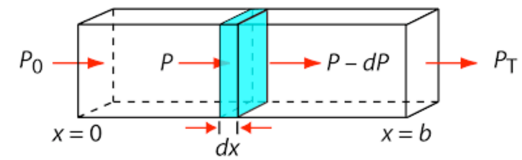
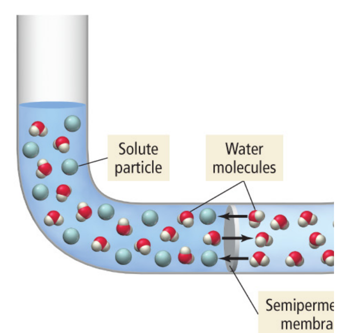


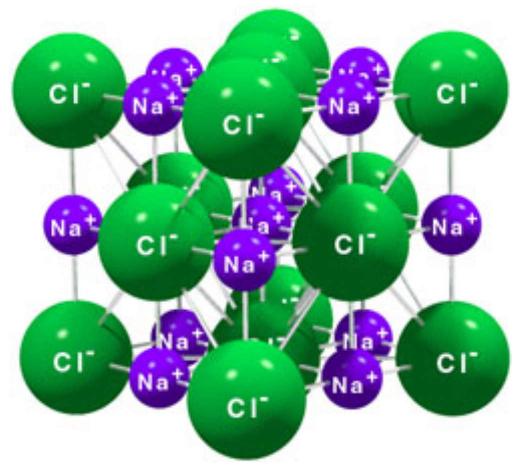
Beer's Law
$A = BC$
$A = \varepsilon \, l \, C$

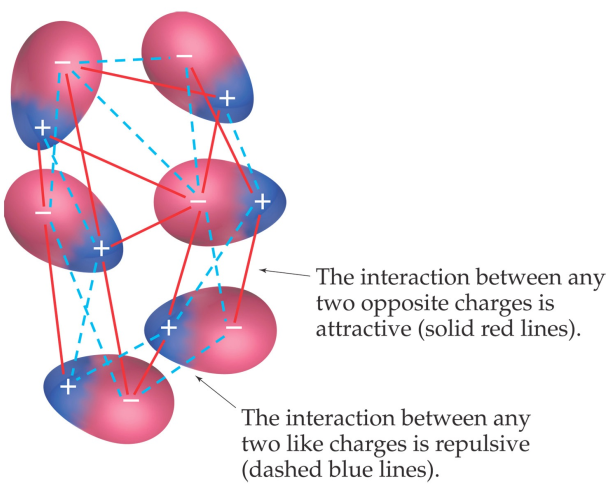
Raoult's Law
$P_A = \chi_A \bullet P^\circ$

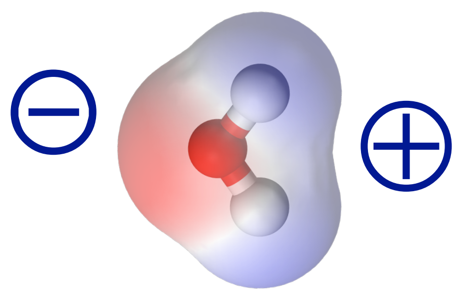


Why solids are solid.

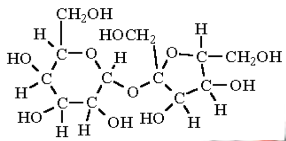
- Intermolecular forces hold solids together.
  - It's usually about plus being attracted to minus (electrostatic attraction).
  - Molecular Solids are held together by many types of intermolecular forces.
    - The quick story is molecules have a negative end and a positive end.
    - The negative end of one molecule sticks to the positive end of another.
    - We'll discuss the rest in Chapter 11.
  - Ionic Solids are held together by one type of intermolecular force.
    - It's a simpler story.
    - The cations stick to a bunch of anions.
    - Those anions stick to more cations.
    - The result is a big clump of particles.

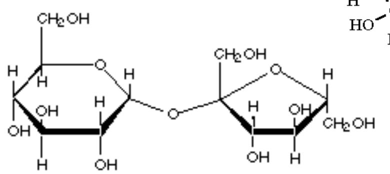




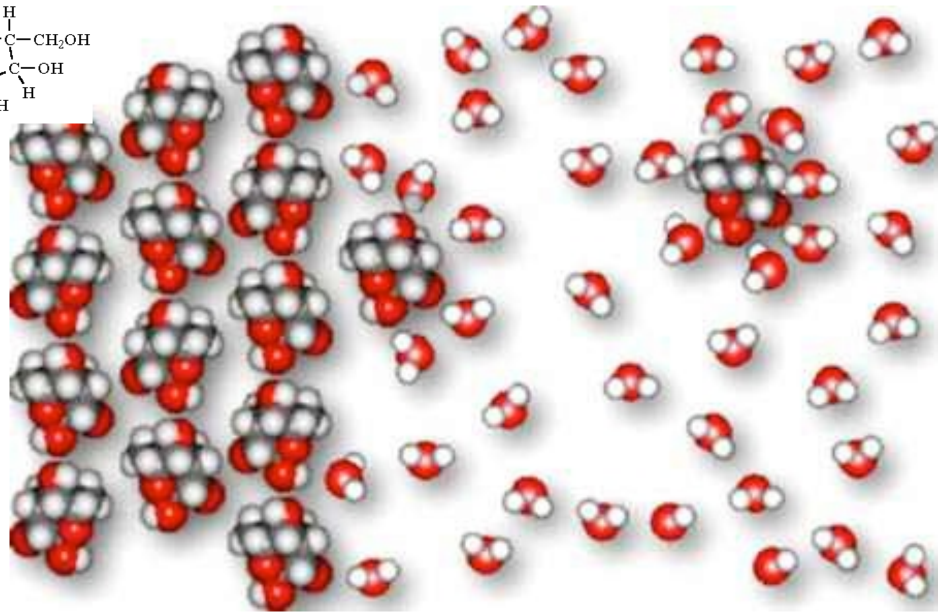


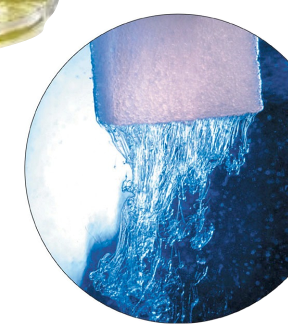
Molecular Solids Dissolve in Water





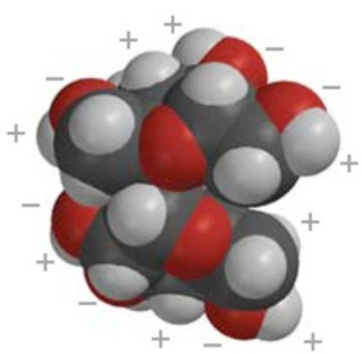
Sucrose
(glucose (α1-->2) fructose)



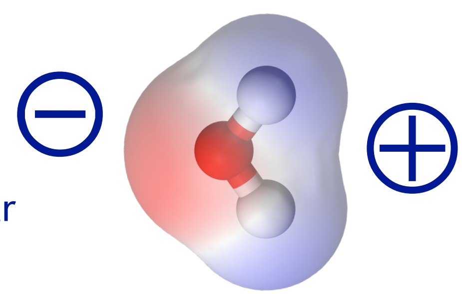


$$C_{12}H_{22}O_{11}\ (s) \xrightarrow{H_2O} C_{12}H_{22}O_{11}\ (aq)$$

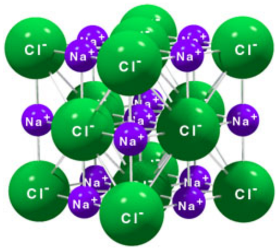
Sugar dissolving in water

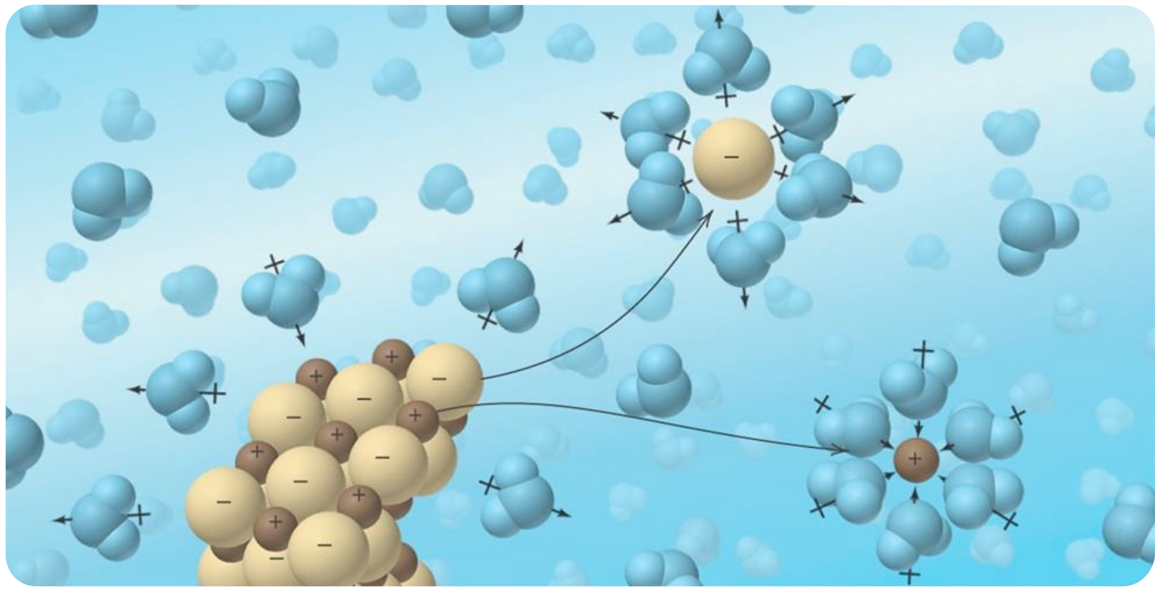


- Sugar dissolves in water.
- The molecules remain intact.
- Water molecules get in between sugar molecules.
- The result is a mixture of sugar and water.
- Mostly water.



Ionic Solids Dissolve in Water

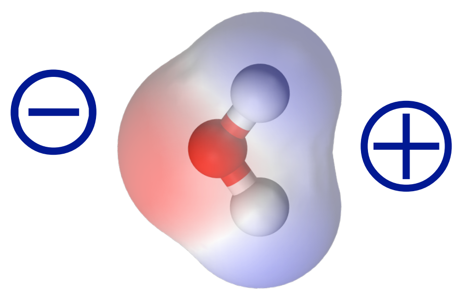




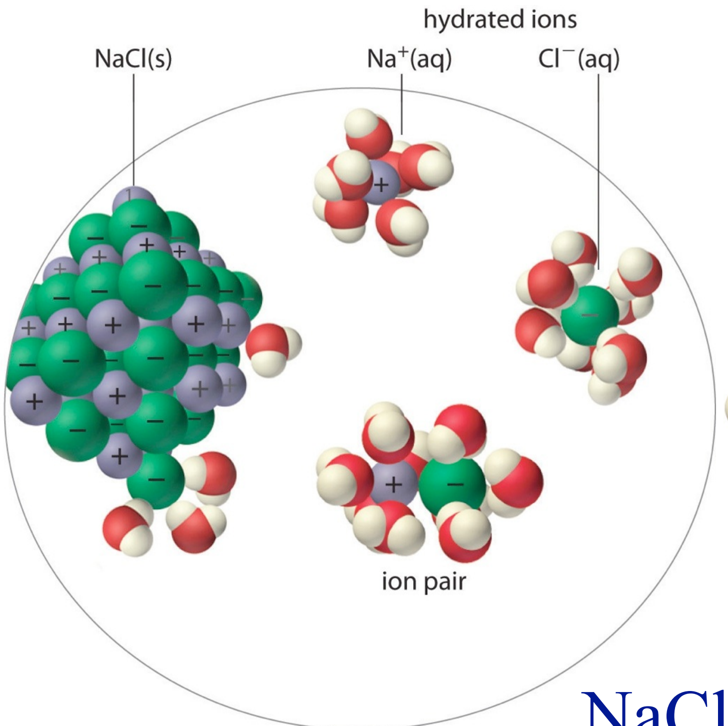
$$NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$$

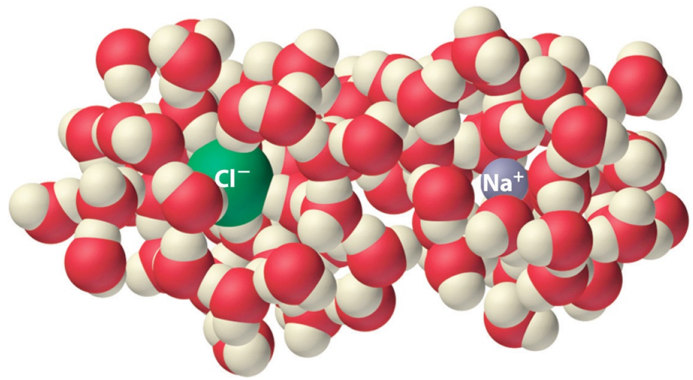
Dissociation of sodium chloride in water

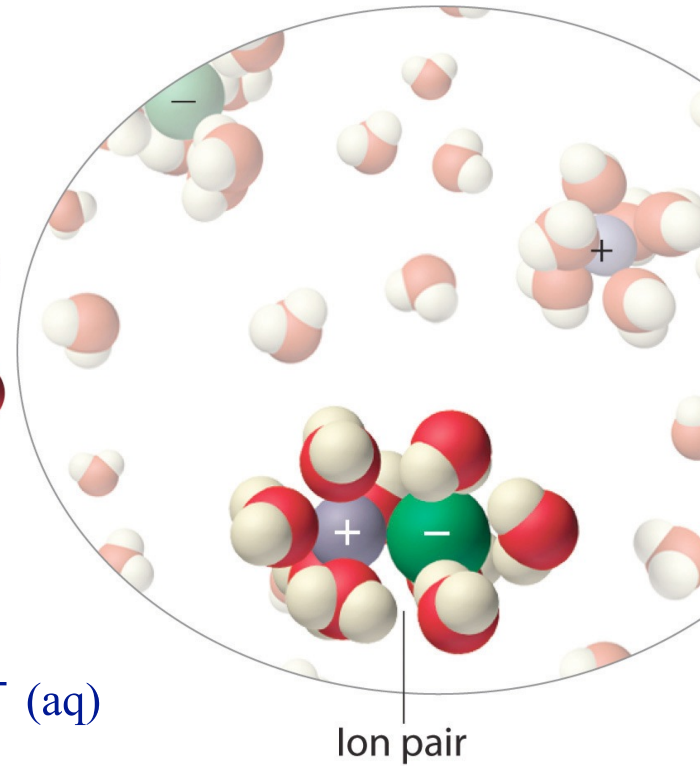
- Salt dissolves in water.
- The the ions separate.
- Water molecules get in between the ions.
- The result is a mixture of ions and water.
- Mostly water.
- Ions separating in solution is a process called dissociation.



Dissociation is often Reversible







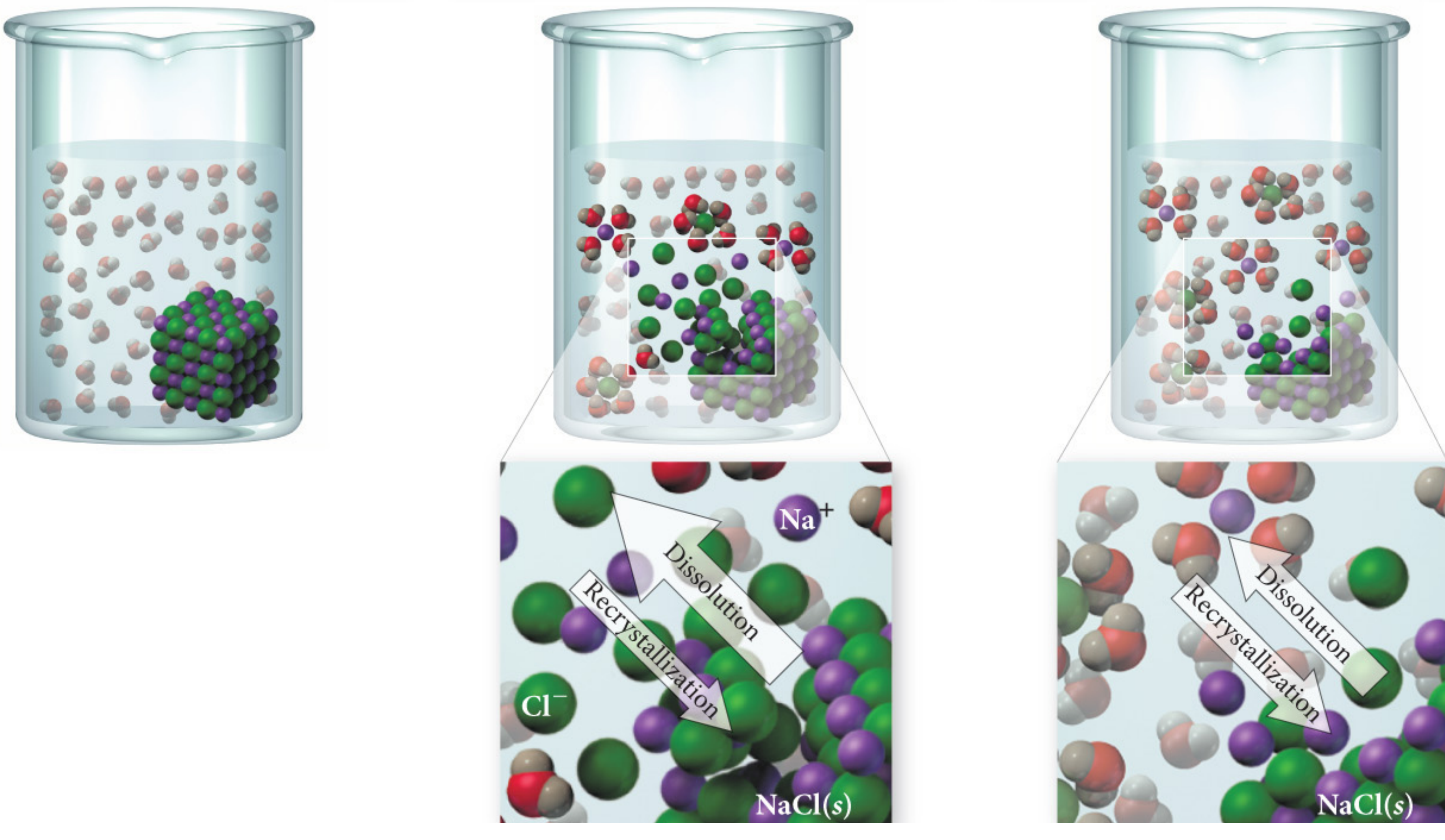
$$NaCl_{(aq)} \xrightleftharpoons{H_2O} Na^{+}_{(aq)} + Cl^{-}_{(aq)}$$

Dissociation of sodium chloride in water

- Dissolved ions in solution can find other dissolved ions.
- If the attraction between those ions is strong, they can re-associate.
- These dissolved ions form ion pairs.
- The ion pair is not a solid, it's still dissolved in solution.
- Ions that dissociate and re-associate in solution are a kind of reversible reaction.

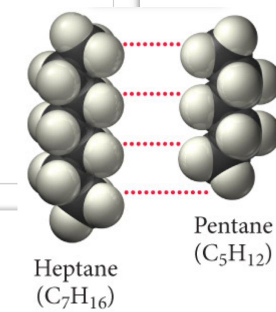
Dynamic Equilibrium

- Crystallization and dissolution are competing processes.
- When the two processes occur just as quickly, the solution is saturated and dynamic equilibrium is achieved.



Ch13

Solution



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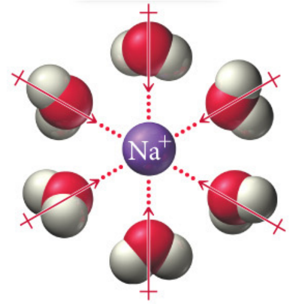

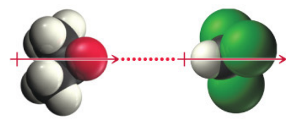

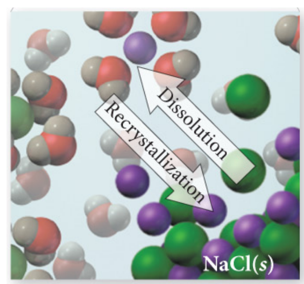

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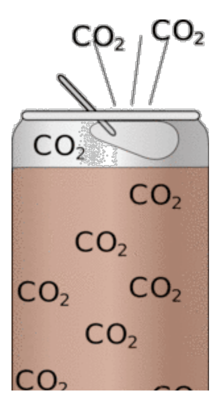

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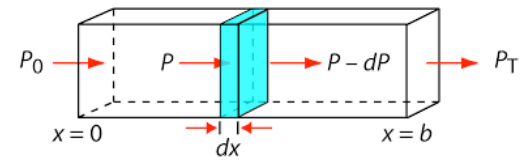

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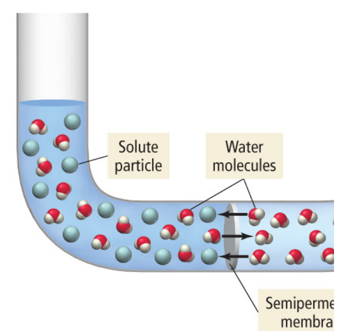
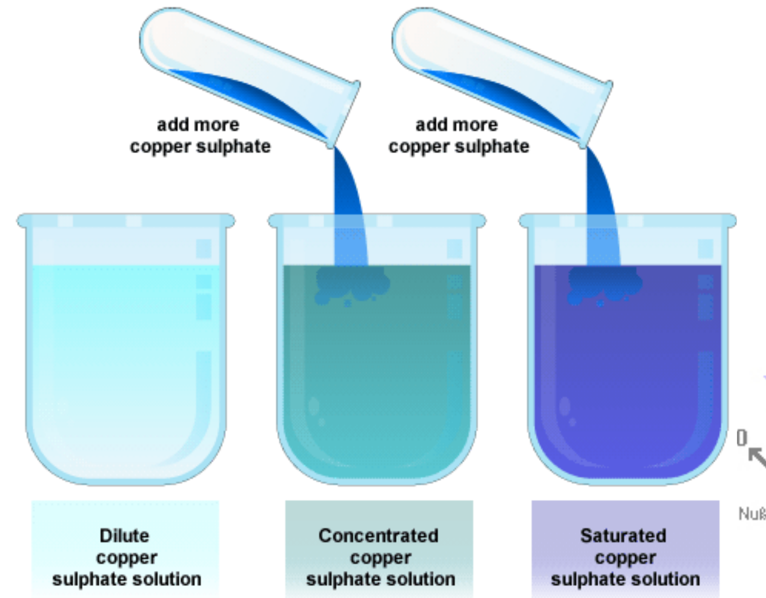


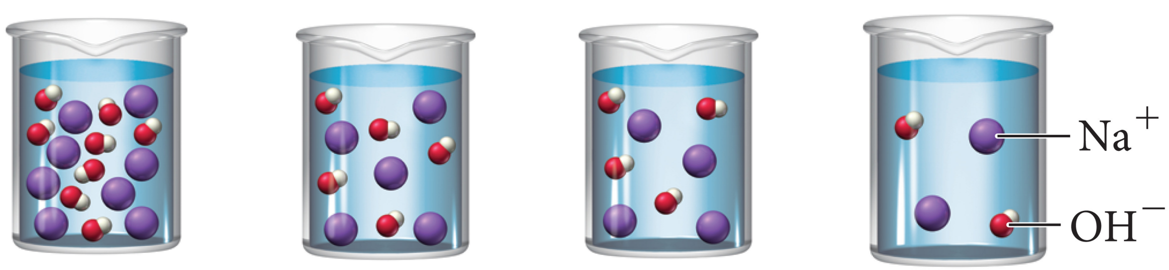
Solutions & Concentration

- Solutions are homogeneous mixtures.
- We know mixtures have tunable properties.
- The properties vary with the ratio of the pure substances that make up that mixture.

We describe that ratio as concentration.

- Concentration is the relationship between amount of a minor component of the mixture (a solute) to the major component of the mixture (the solvent).
- Concentration is how “crowded” the mixture is in a substance.
- Concentration is the amount of a solute in a given quantity of solvent.
- Solutions that contain greater amounts of solute are said to be more concentrated.
- Solutions that contain lesser amounts of solute are said to be more dilute.
- Solutions that contain the maximum amount of solute a solution can hold are said to be saturated.





A solution is a homogenous mixture.

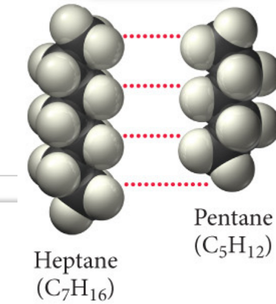
A solvent is the largest component of the mixture.

A solute is a smaller components of the mixture.

Dynamic Equilibrium

- Crystallization and dissolution are competing processes.
- When the two processes occur just as quickly, the solution is saturated and dynamic equilibrium is achieved.
- It is possible to hinder crystallization and in so doing create a super saturated solution—but these solutions can spontaneously crystallize without warning.

Solution



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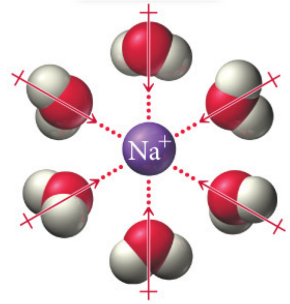

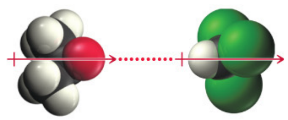

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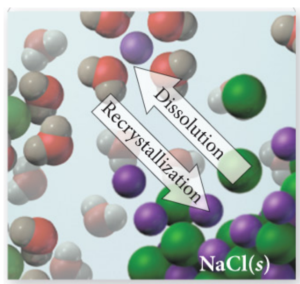

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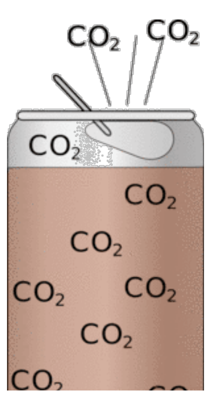

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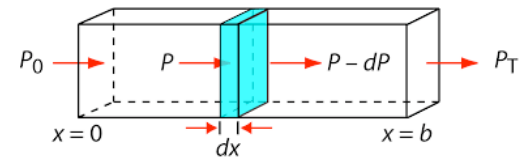

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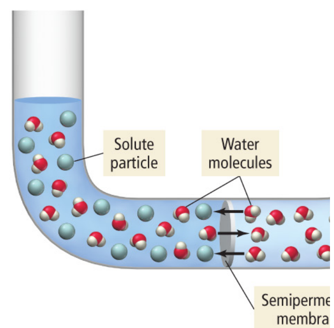
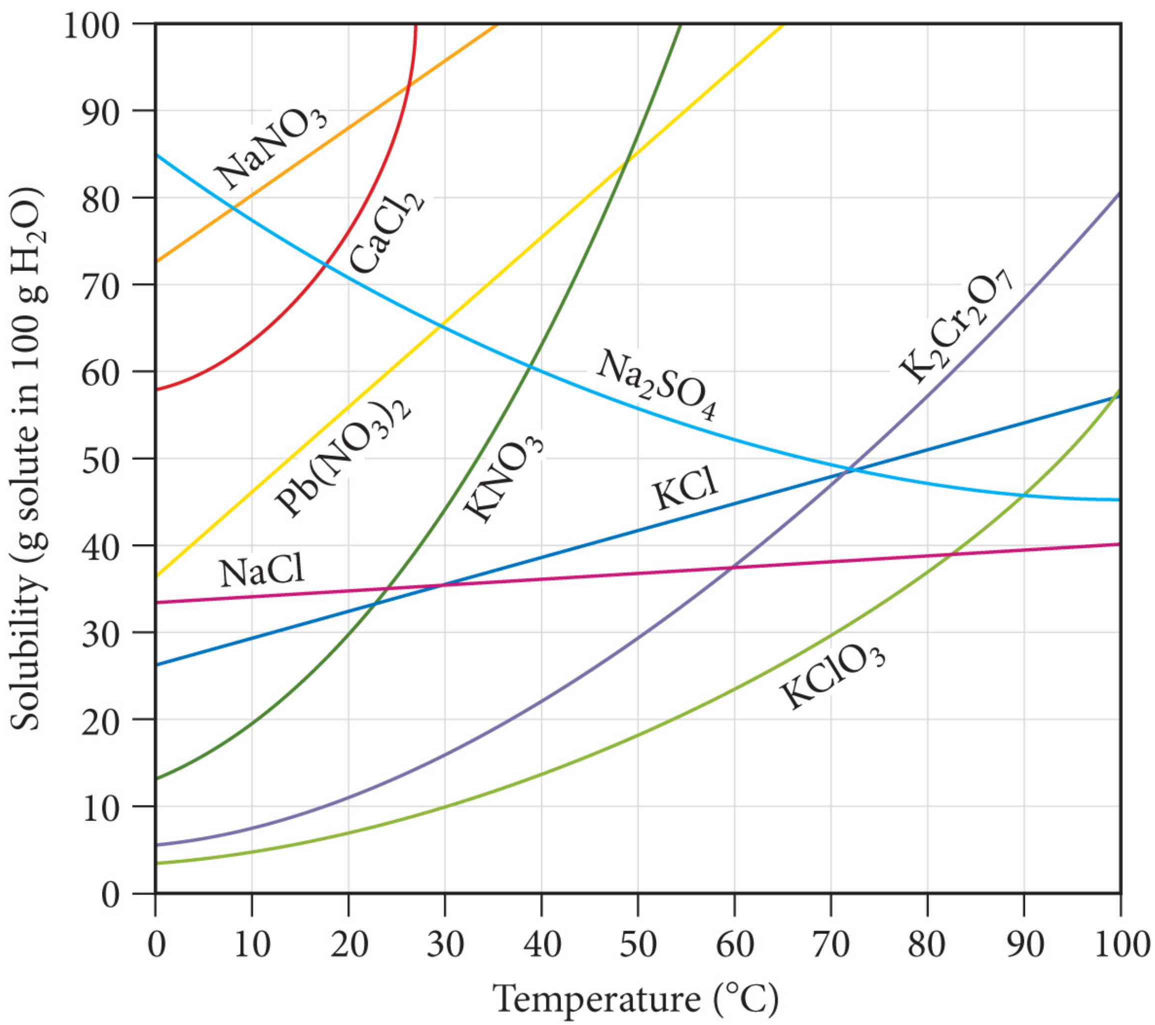


Temperature on Solubility of Solids

- The maximum solubility (saturation) of a solution is temperature dependent.
- For solids, solubility generally **increases** with increasing temperature.



Which of the following statements is NOT true?

a) All soluble ionic compounds become more soluble with increasing temperature.
b) Additional solvent cannot dissolve in a saturated solution.
c) Some ionic compounds become less soluble with increasing temperature.
d) Supersaturated solutions are very stable.

Which of the following statements is NOT true?

a) All soluble ionic compounds become more soluble with increasing temperature.
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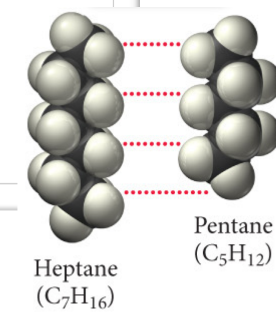
Solution



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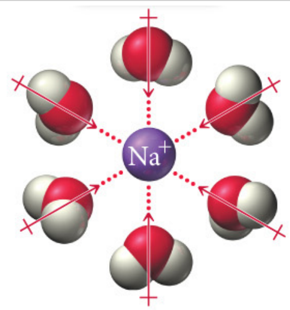

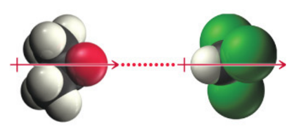

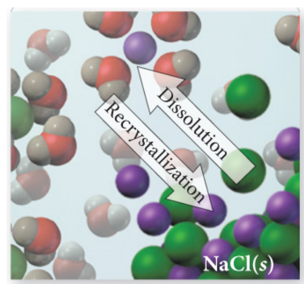

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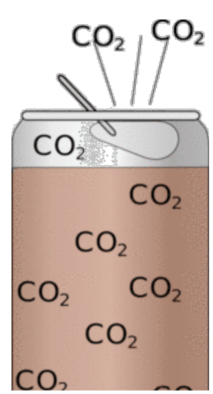

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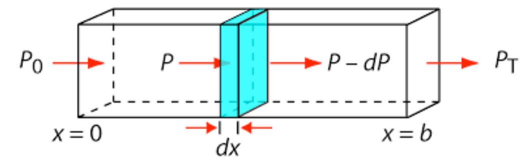

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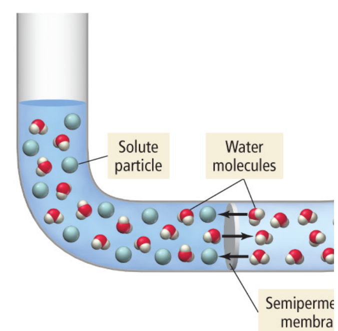


Temperature on Solubility of Gases

- The maximum solubility (saturation) of a solution is temperature dependent.
- For gases, solubility generally **decreases** with increasing temperature.

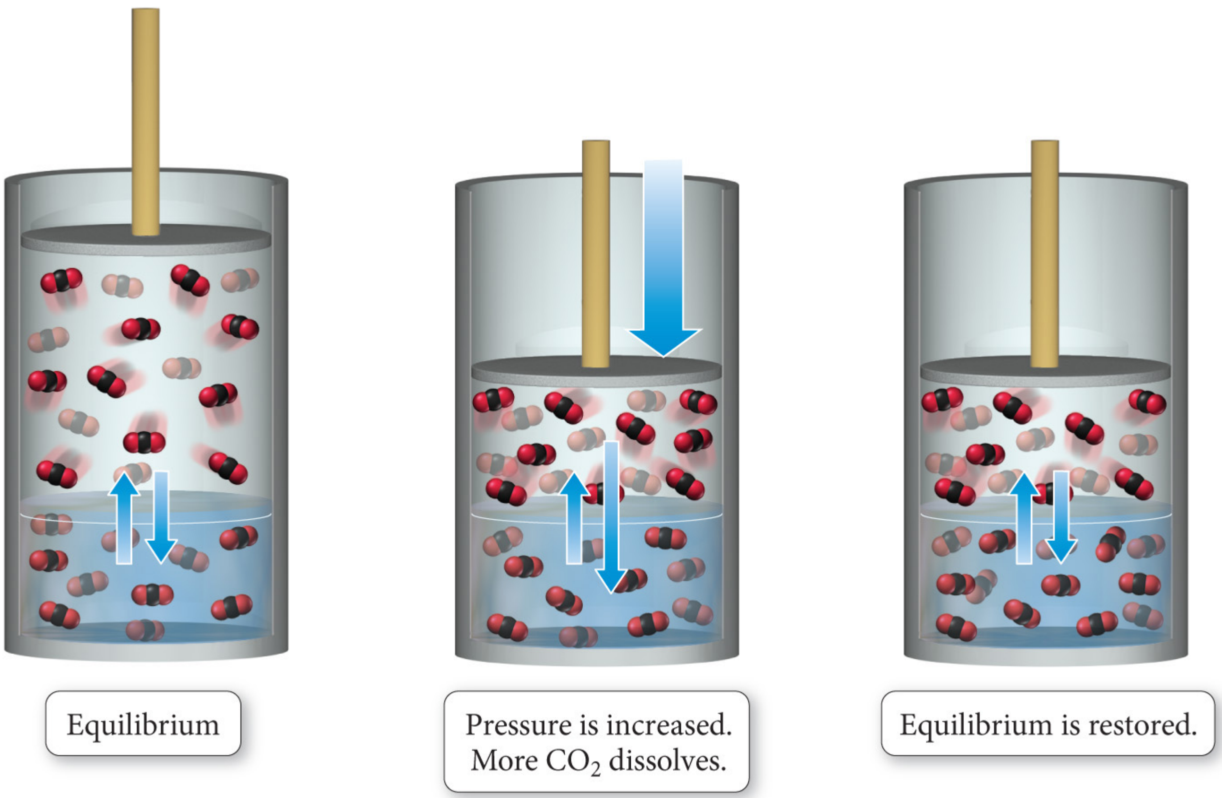
Cold soda pop | Warm soda pop

Henry’s Law

- The partial pressure of the solute also effects solubility. Increasing the pressure of gas linearly increases is maximum solubility in solution.
- Henry’s Law "At a constant temperature, the amount of a given gas that dissolves in a given type and volume of liquid is directly proportional to the partial pressure of that gas in equilibrium with that liquid.”

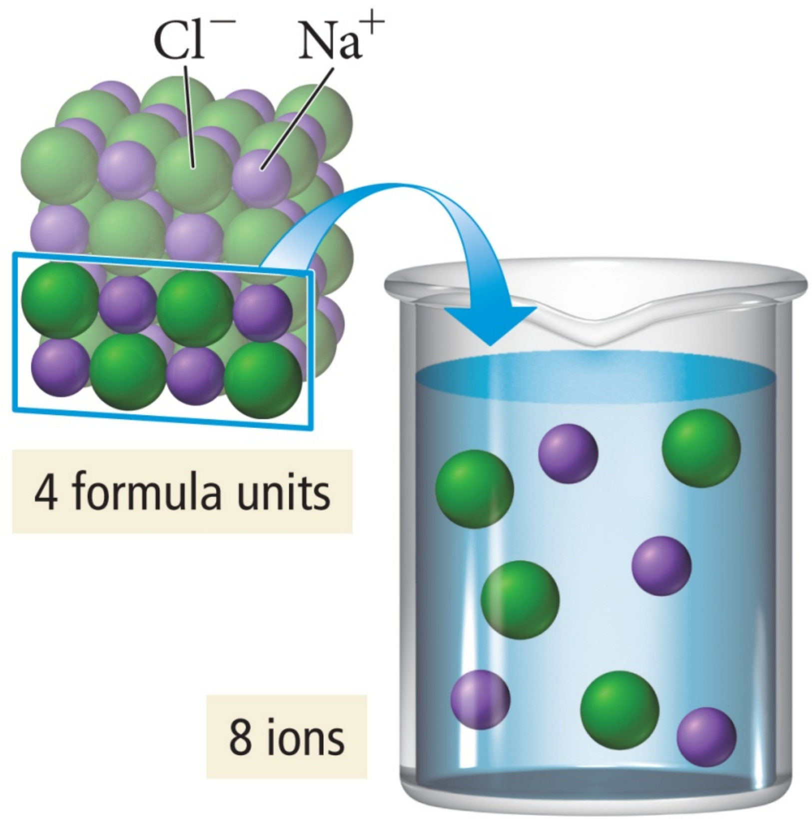
Henry’s Law

$$[A] = k_A \, P_A$$



Which of the following solutions has its solubility most strongly dependent on pressure?

a) 0.05 *m* NaCl
b) 0.05 *m* $C_6H_{12}O_6$
c) 0.02 *m* $Al(NO_3)_3$
d) 0.005 *m* CO_2
e) 0.02 *m* NH_4Cl



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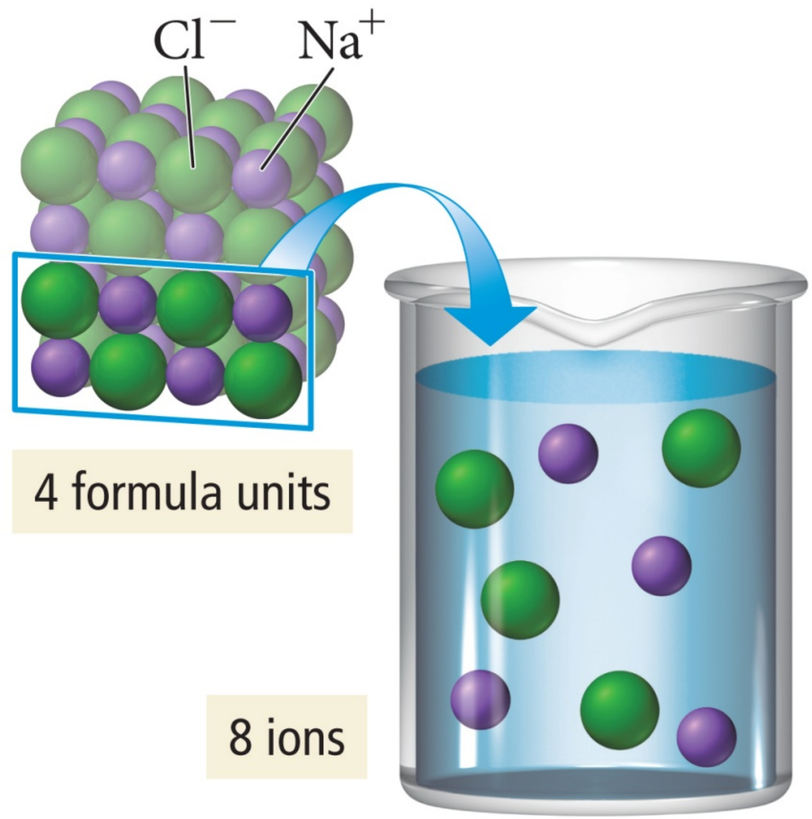


Henry’s Law

- "At a constant temperature, the amount of a given gas that dissolves in a given type and volume of liquid is directly proportional to the partial pressure of that gas in equilibrium with that liquid.”

$$S_{CO_2} \longrightarrow P_{CO_2}$$

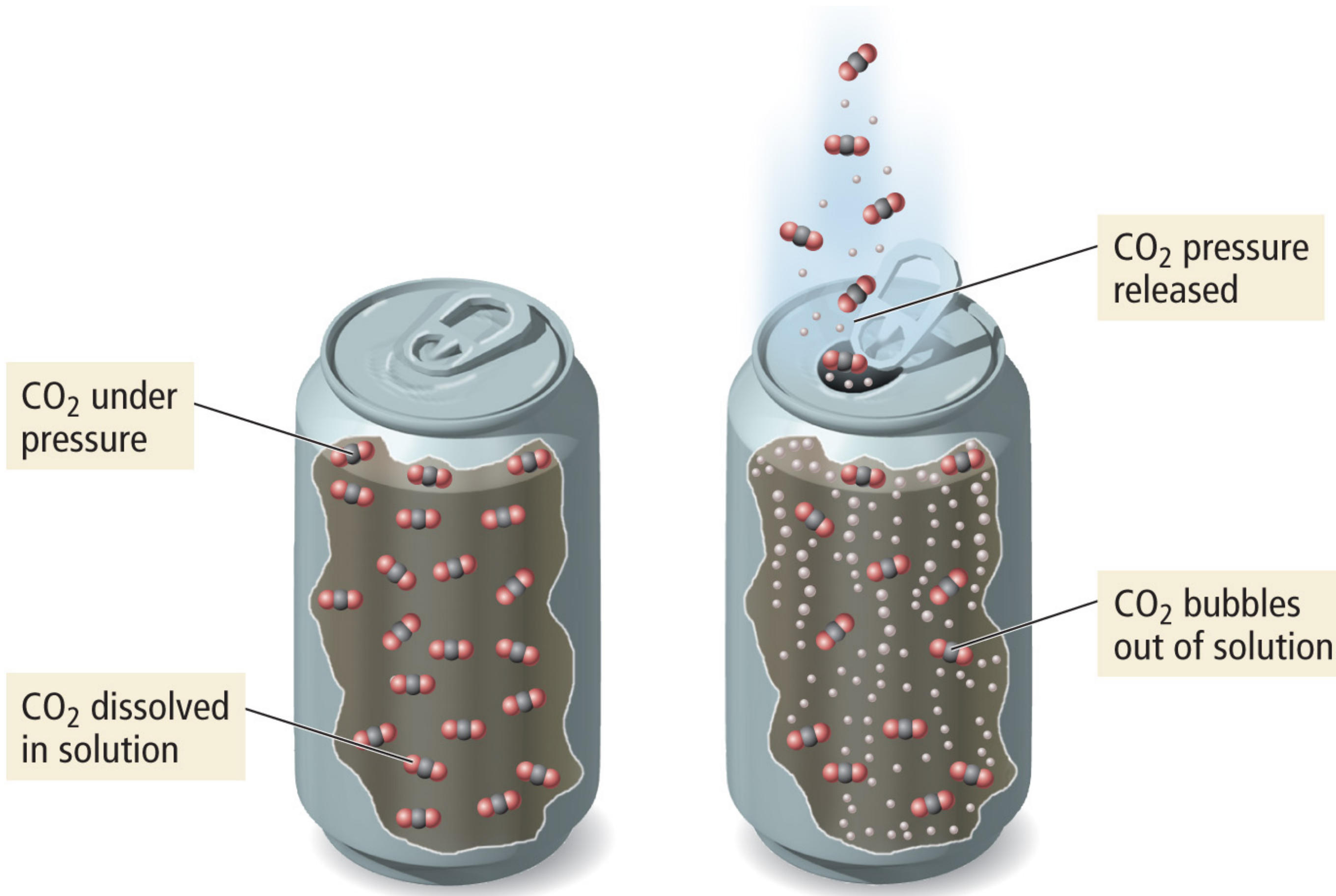
$$S_{CO_2} = k_{H,CO_2} P_{CO_2}$$

$$[A] = k_A \, P_A$$

Gas	k_H (M/atm)
O_2	1.3×10^{-3}
N_2	6.1×10^{-4}
CO_2	3.4×10^{-2}
NH_3	5.8×10^{1}
He	3.7×10^{4}

A pressure of 3.5 atm in CO_2 is required to maintain a 0.12 M CO_2 concentration in soda. Calculate the Henry's Law constant for CO_2. (at that temperature)

a) 29.2
b) 0.42
c) 2.9×10^{1}
d) 3.4×10^{-2}
e) 4.2



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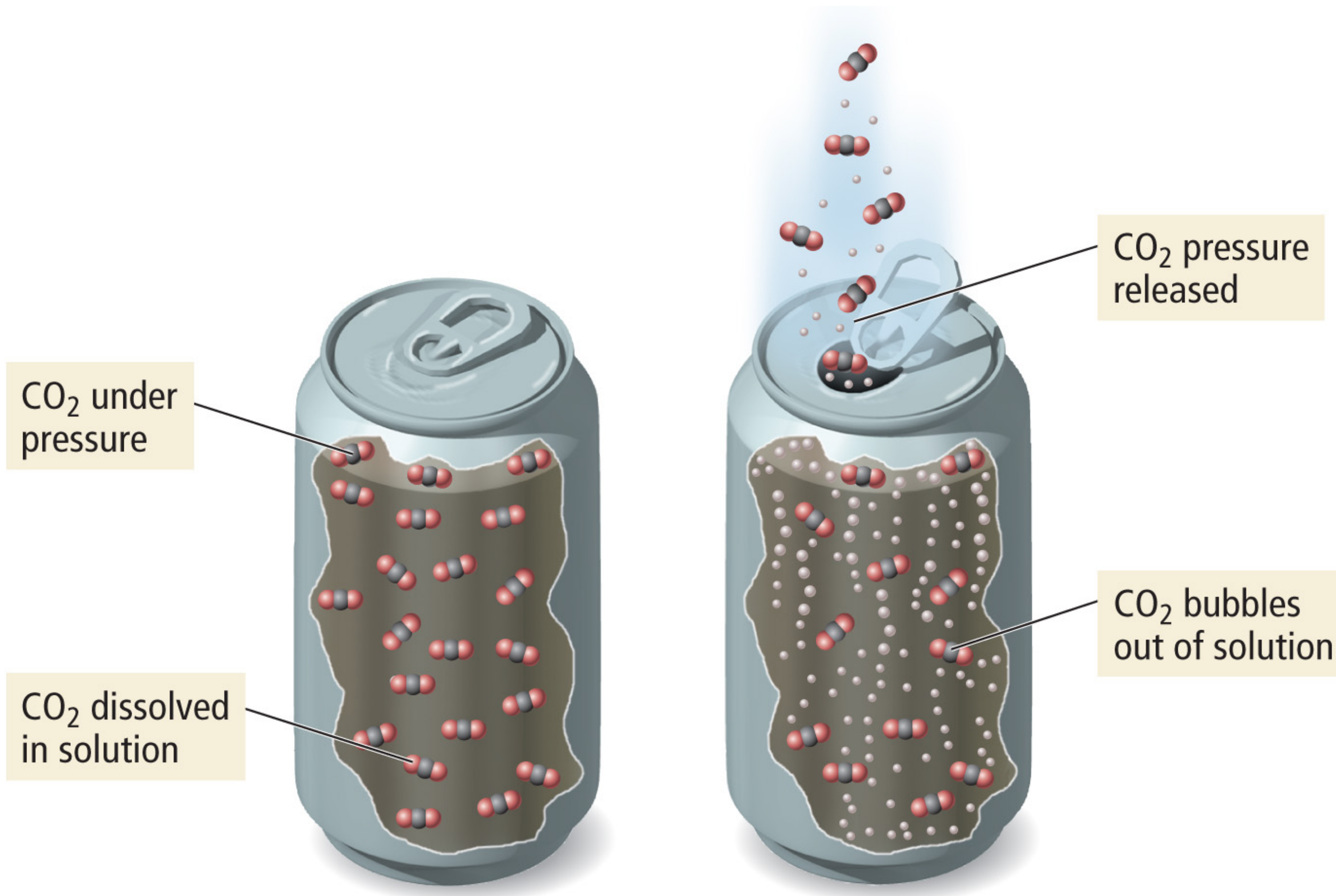
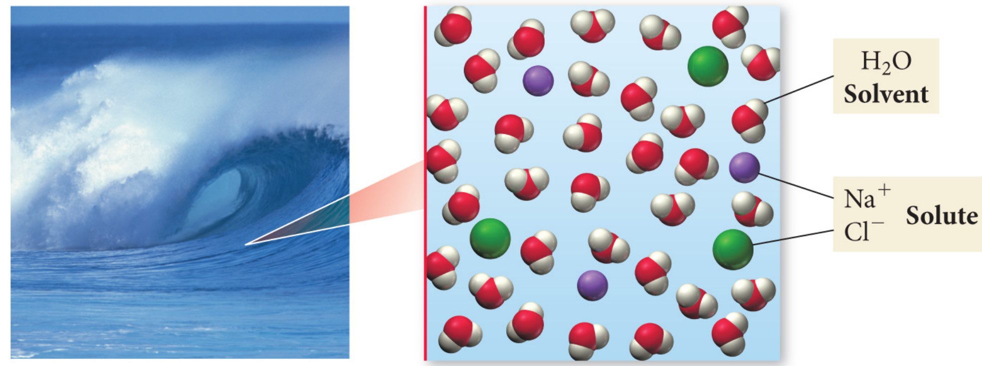


On a clear day at sea level, the partial pressure of N_2 in air is 0.78 atm at 25°C. Under these conditions, the concentration of N_2 in water is 5.3×10^{-4} M. What is the partial pressure of N_2 when the concentration in water is 1.1×10^{-3} M?



a) 0.63 atm
b) 0.78 atm
c) 1.0 atm
d) 2.1 atm
e) 1.6 atm

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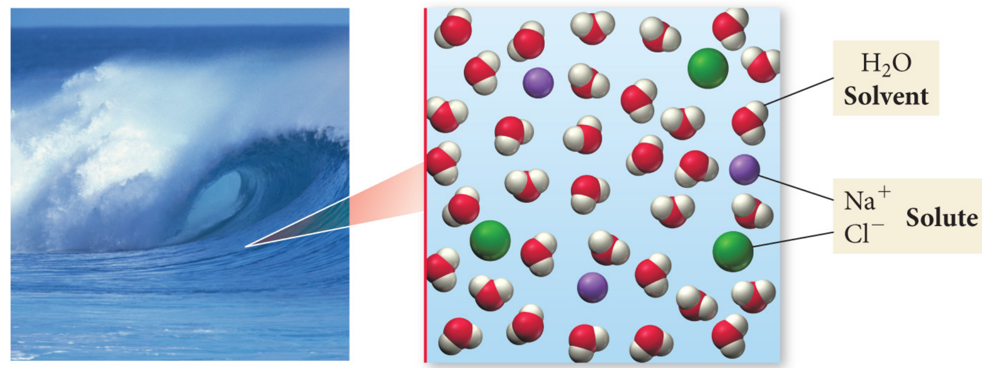
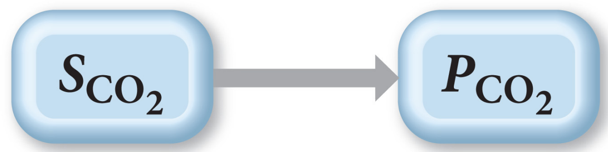




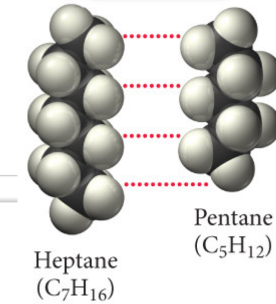
$$S_{CO_2} = k_{H,CO_2} P_{CO_2}$$

Factors Effecting Solubility

	Temp ↑	Pressure ↑
solid in liquid solubility	↑ (usually)	N/C
gas in liquid solubility	↓	↑

Ch13

Solution



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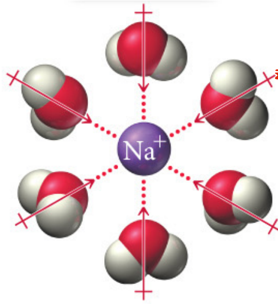

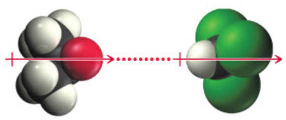

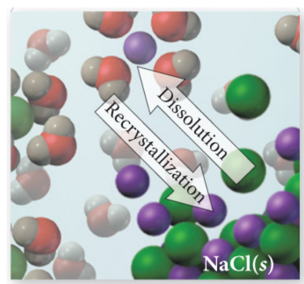

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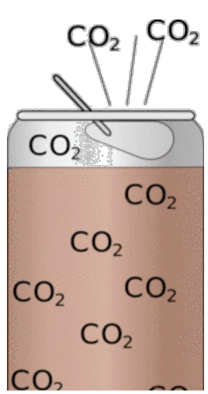
$[A] = k_A P_A$



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- Concentration
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    - Van't Hoff Factor
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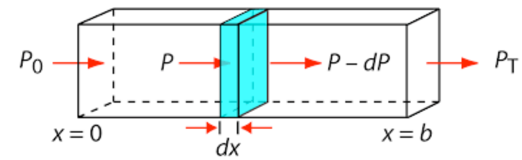

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$A = BC$
$A = \varepsilon\, l\, C$

Raoult's Law
$P_A = \chi_A \bullet P^\circ$

- Dispersions
  - Colloids
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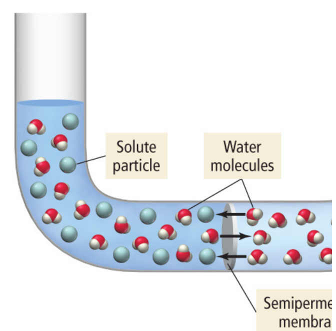
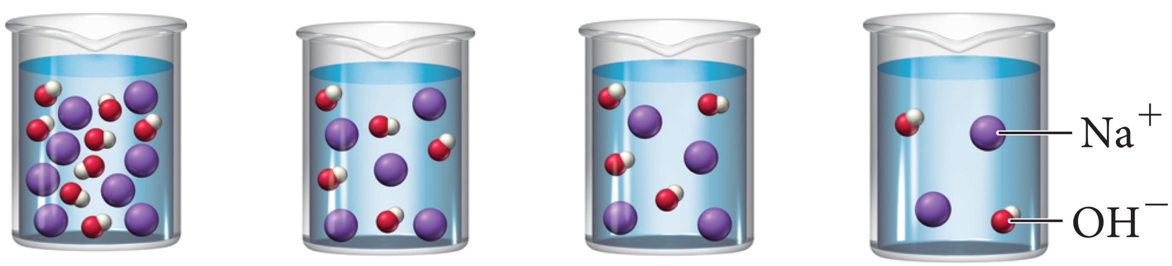


Measures of Concentration

- Concentration of a substance is represented at the substances chemical formula in brackets.

[NaCl] means concentration of sodium chloride.

Measurement of [NaCl] can be given in many different units. There are competitive advantages to each measure.

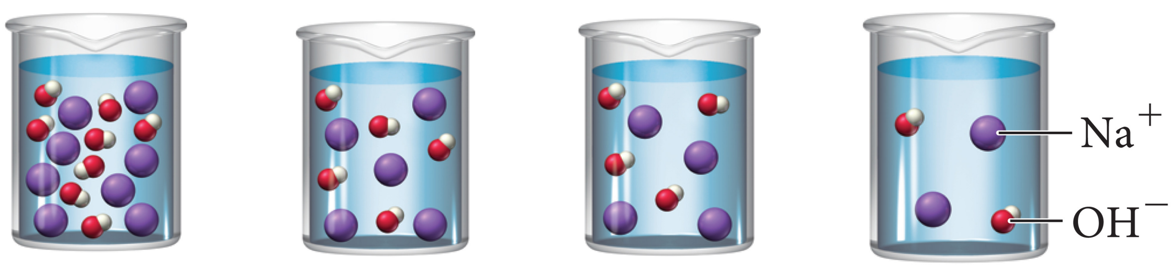


Measures of Concentration

- There are many units of measure for concentration.

Unit	Definition	Units
Molarity (M)	$\frac{\text{amount solute (in mol)}}{\text{volume solution (in L)}}$	$\frac{\text{mol}}{\text{L}}$
Molality (m)	$\frac{\text{amount solute (in mol)}}{\text{mass solvent (in kg)}}$	$\frac{\text{mol}}{\text{kg}}$
Mole fraction (χ)	$\frac{\text{amount solute (in mol)}}{\text{total amount of solute and solvent (in mol)}}$	None
Mole percent (mol %)	$\frac{\text{amount solute (in mol)}}{\text{total amount of solute and solvent (in mol)}} \times 100\%$	%
Parts by mass	$\frac{\text{mass solute}}{\text{mass solution}} \times$ multiplication factor	
Percent by mass (%)	Multiplication factor $= 100$	%
Parts per million by mass (ppm)	Multiplication factor $= 10^6$	ppm
Parts per billion by mass (ppb)	Multiplication factor $= 10^9$	ppb
Parts by volume (%, ppm, ppb)	$\frac{\text{volume solute}}{\text{volume solution}} \times$ multiplication factor *	

*Multiplication factors for parts by volume are identical to those for parts by mass.



Measures of Concentration

- There are many units of measure for concentration.
- Three common measures are:
  - Mole Fraction (X)
    - Moles of solute per mole of solution.
    - We'll use this when we discuss gases, it's less useful for liquids.
  - Molality (m)
    - Moles of solute, per kg of **solvent**.
    - Useful for liquids when temperature changes.
  - Molarity (M)
    - Moles of solute per liter of solution.
    - Most generally useful measure for liquids

$$X = \frac{\text{moles of solute}}{\text{moles of solution}}$$

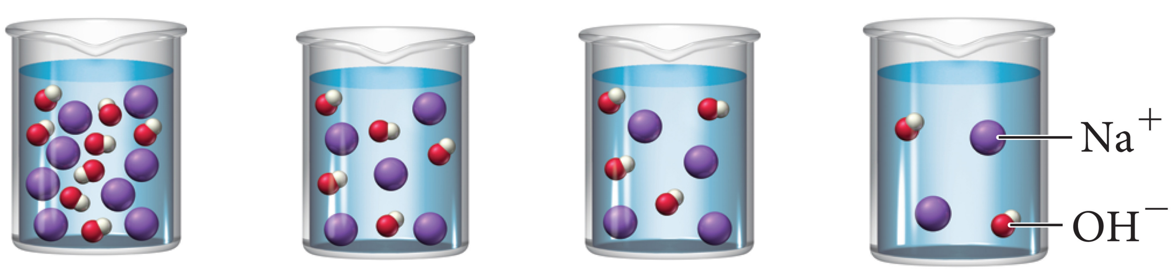
mole per mole

$$m = \frac{\text{moles of solute}}{\text{kilogram of solvent}}$$

mole per mass

$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

mole per volume



Measures of Concentration

- Percent mass and “parts” are also useful.

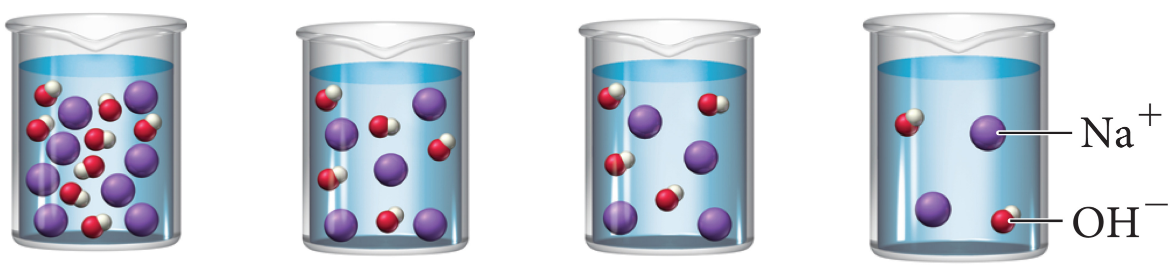
parts per million (ppm) 1.0×10^{-6}

-- commonly used for minerals or contaminants in water supplies.

$$Concentration = \left(\frac{mass\ of\ solute}{volume\ of\ solution} \right)$$

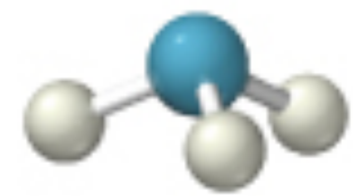
Mass percent - the ratio of mass units of solute to mass units of solution, expressed as a percent

$$Mass\ percent = \left(\frac{mass\ of\ solute}{mass\ of\ solution} \right) x100$$



A solution of ammonia is made by dissolving 35 g of NH_3 in 90.0 g of H_2O. The density of the solution is 0.898 g/mL. What is the molarity of NH_3?

a) 53 M
b) 0.39 M
c) 23 M
d) 15 M
e) not possible to calculate



(hint: concentration is X over Y, easiest way to convert between measures is often to convert X and Y separately and then combine)

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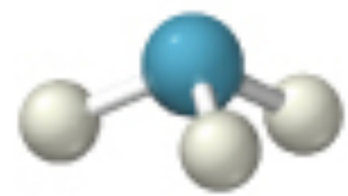
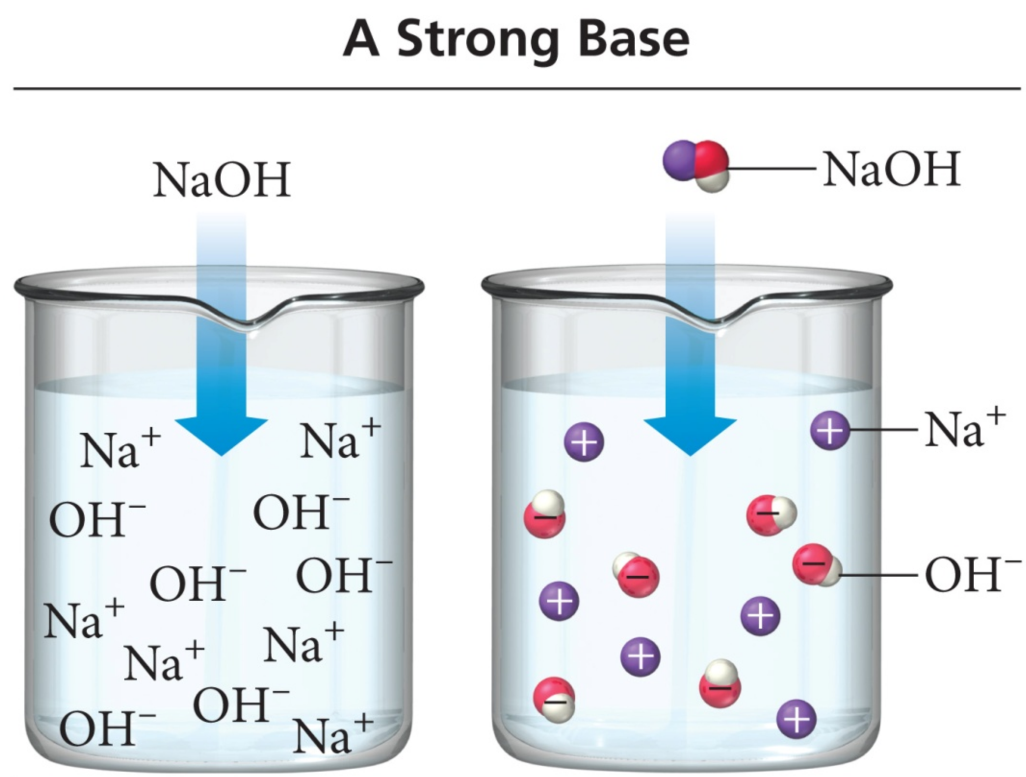


(hint: concentration is X over Y, easiest way to convert between measures is often to convert X and Y separately and then combine)

What is the molality a 4.0 M NaOH solution? The density of the solution is 1.04 g/mL.

a) 0.15 *m*
b) 4.5 *m*
c) 0.18 *m*
d) 4.0 *m*
e) 3.8 *m*



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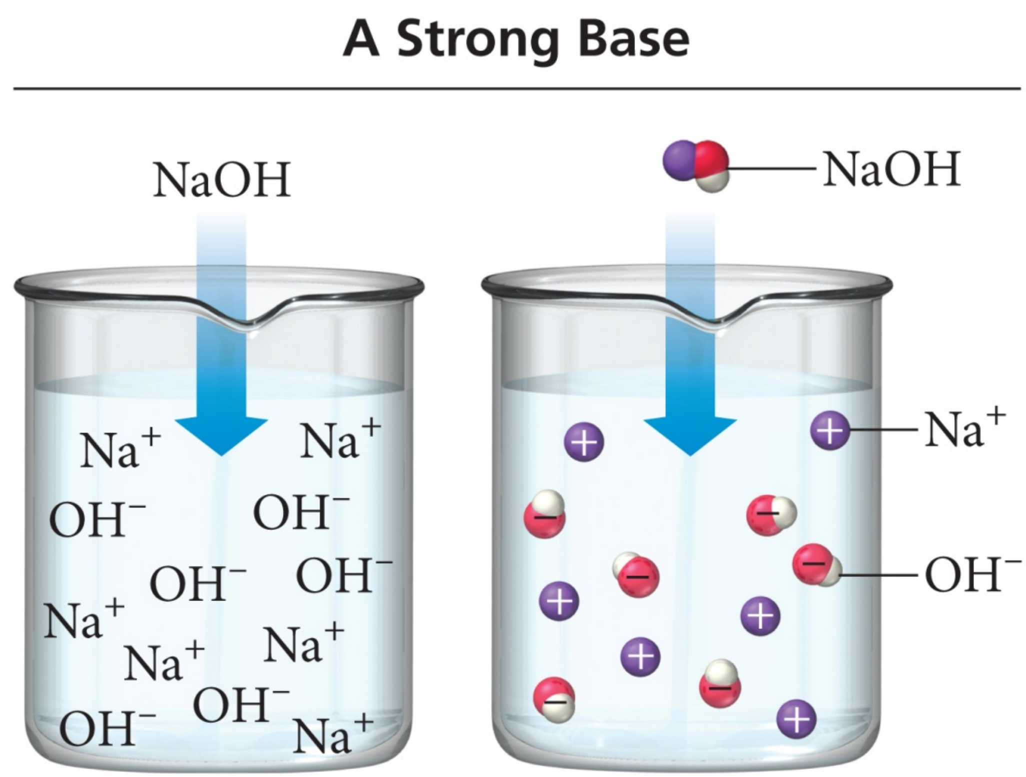
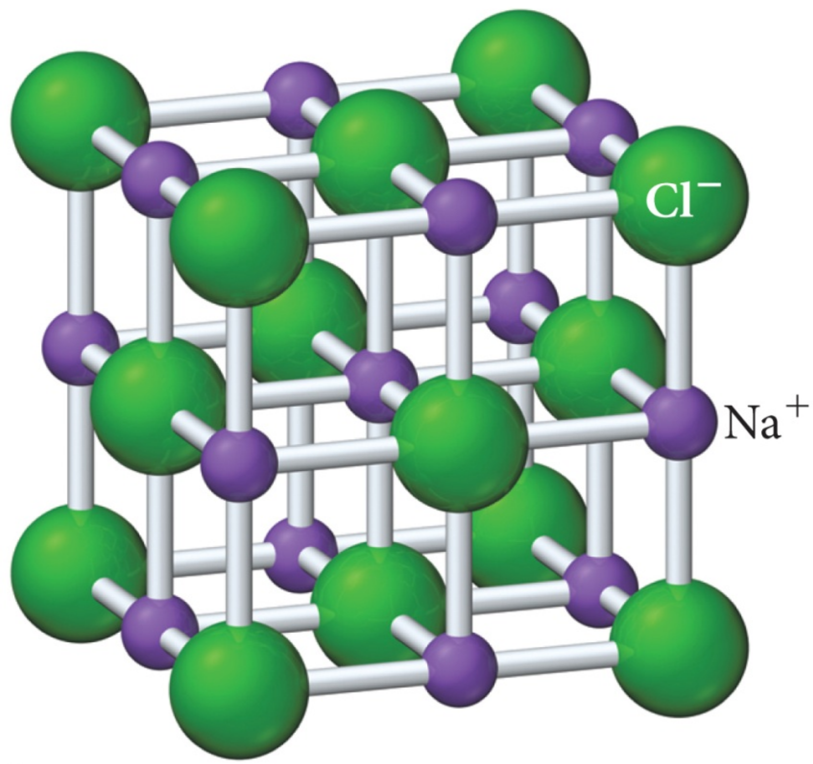


(hint: concentration is X over Y, easiest way to convert between measures is often to convert X and Y separately and then combine)

What is the mass percent of a solution prepared from 15 g of NaCl in 45 g of H_2O?

a) 33%
b) 25%
c) 50%
d) 15%

Sodium chloride (NaCl)



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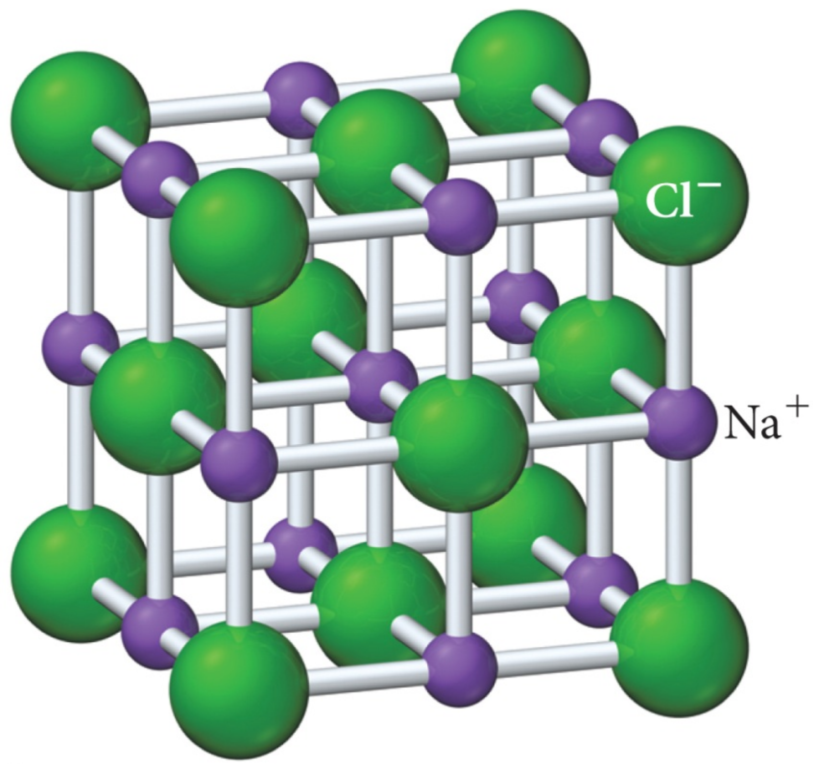
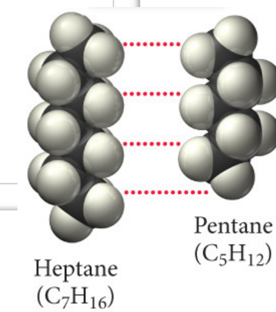


Solution



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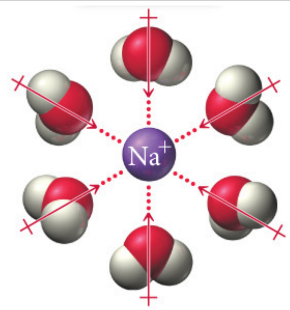

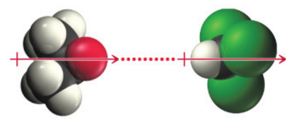

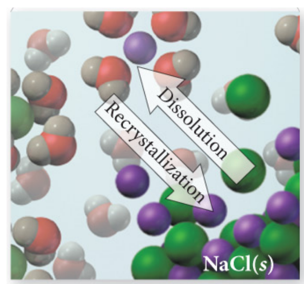

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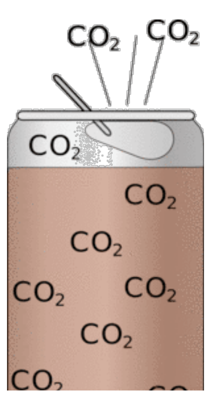

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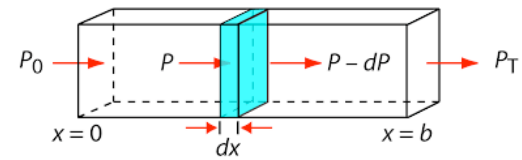

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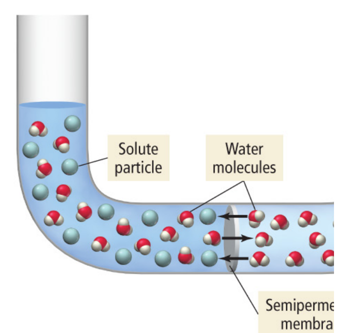


Molarity

- Molarity is the number of moles of a solute divided by the total volume of the solution in L.
- Molarity (M) means moles per liter (mol/L).

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

3.0 M HCl means that solution is 3.0 moles/L in concentration.

Not that a particular sample contains 3.0 moles.

Molarity

- Molarity is the number of moles of a solute divided by the total volume of the solution in L.
- Molarity makes it easy to interconvert between volumes of a solution and mols of solute.
- e.g. if I have 3.0 M H_2SO_4
  - How many mols H_2SO_4 in 0.150 L?

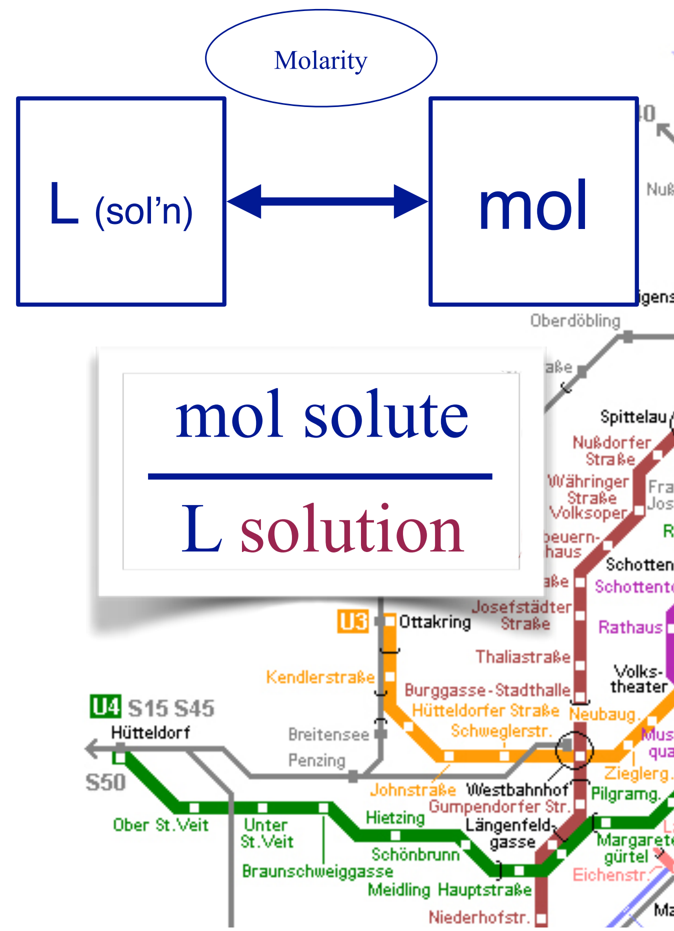
① $L \rightarrow mol$ ① $3.0\,mol = 1\,L$

$$0.150\,L \cdot \frac{3.0\,mol}{1\,L} = 0.45\,mol\ H_2SO_4$$

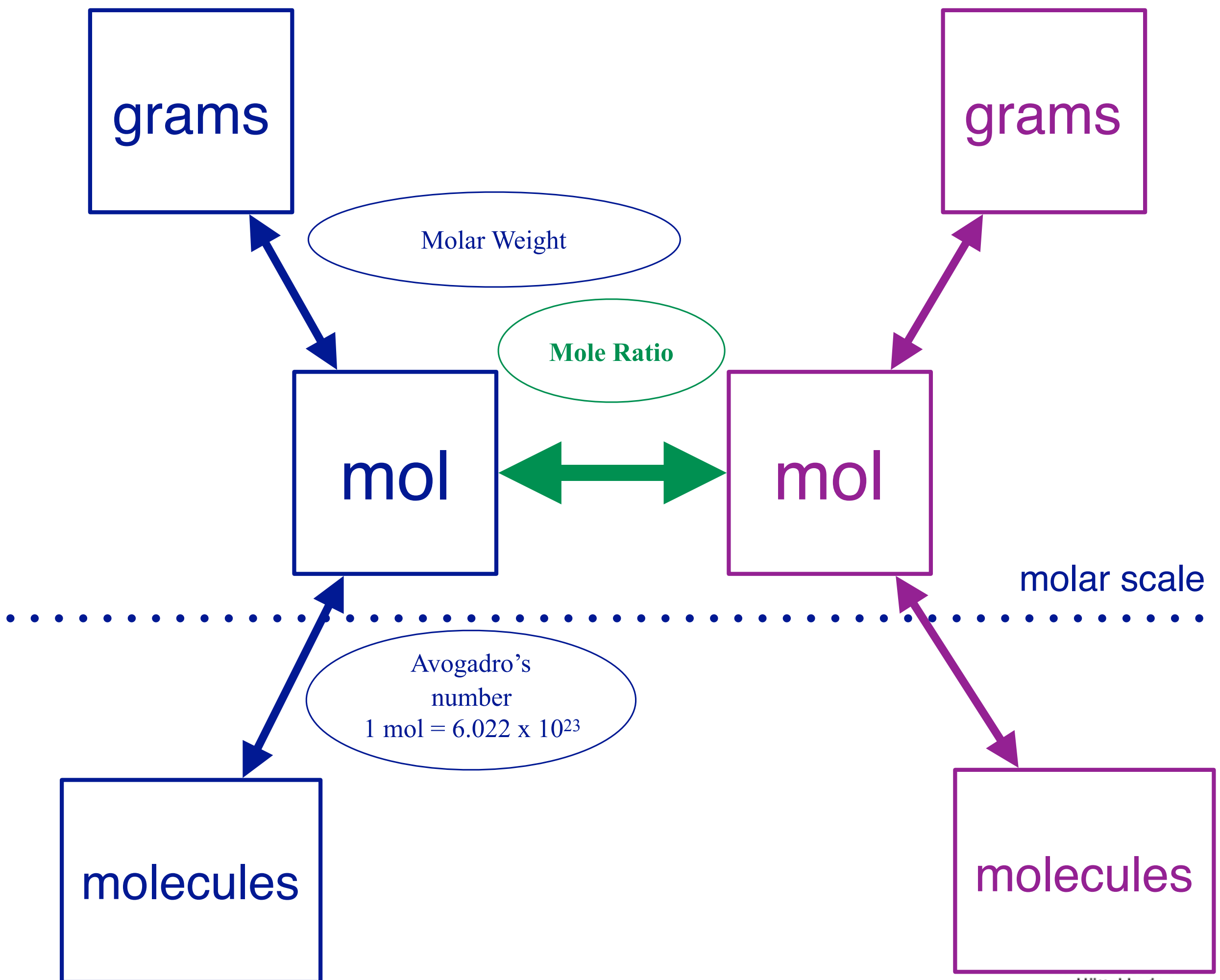
  - What volume do I need to get 0.42 mol?

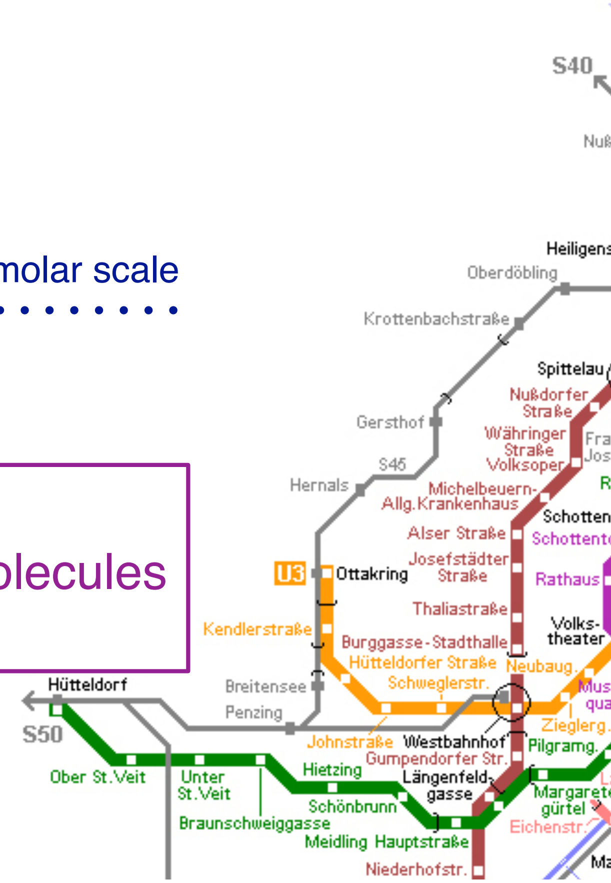
① $mol \rightarrow L$ ① $3.0\,mol = 1\,L$

$$0.42\,mol \cdot \frac{1\,L}{3.0\,mol} = 0.14\,L \quad (140\,mL)$$



The Molar Subway



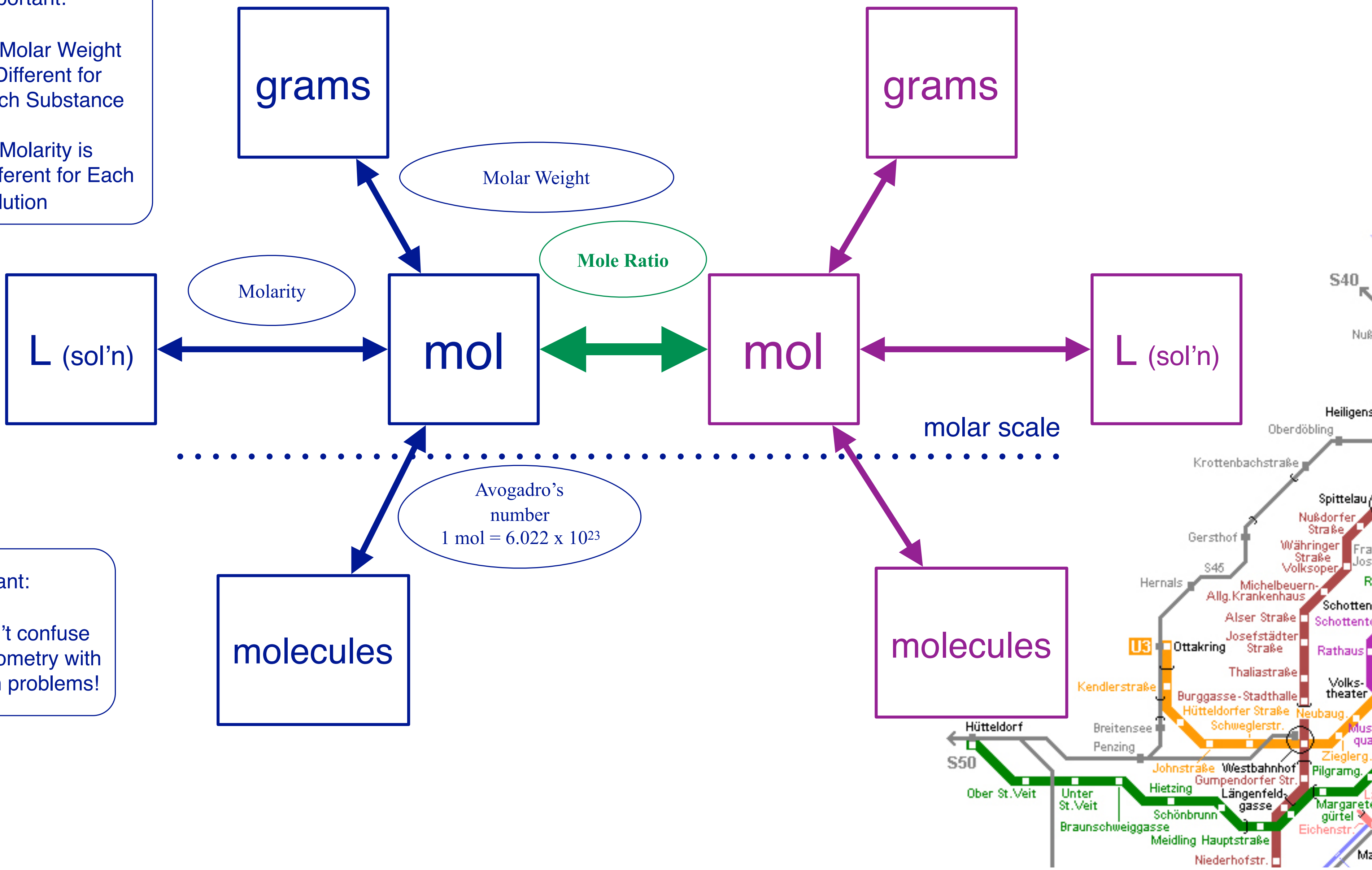


The Molar Subway

Important:

— Molar Weight is Different for Each Substance

— Molarity is Different for Each Solution

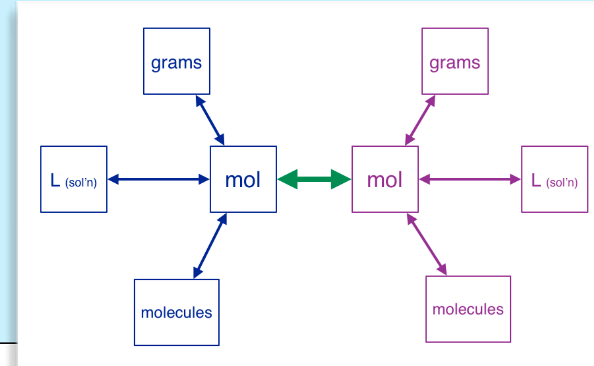


Important:

— Don’t confuse stoichiometry with dilution problems!

Problem:

How many grams of $CaCl_2$ are needed to completely react with 25.0 mL of 0.100 M $AgNO_3$?



Solution

$$CaCl_{2\,(s)} + 2\,AgNO_{3\,(aq)} \rightarrow Ca(NO_3)_{2\,(aq)} + 2\,AgCl_{(s)}$$

$$\underbrace{mL \xrightarrow{①} L \xrightarrow{②} mol}_{AgNO_3} \xrightarrow{③} \underbrace{mol \xrightarrow{④} g}_{CaCl_2}$$

① $1000\ mL = 1\ L$

② $0.100\ mol = 1\ L$

③ $2\ AgNO_3 = 1\ CaCl_2$

④
1 (Ca) 40.078
2 (Cl) 70.906
110.984

$110.984\ g = 1\ mol$

3 s.f. ∞ 3 s.f. ∞ 6 s.f.

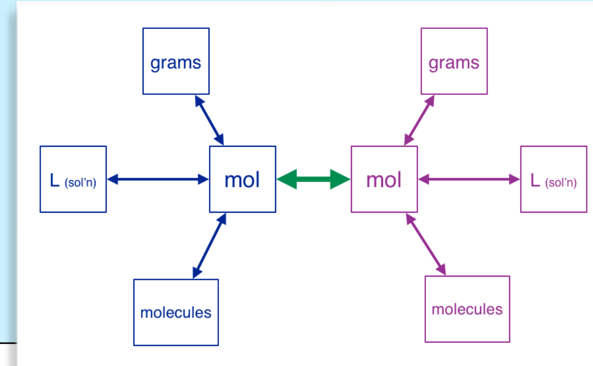
$$25.0\ mL \cdot \frac{1\ L}{1000\ mL} \cdot \frac{0.100\ mol}{1\ L} \cdot \frac{1\ CaCl_2}{2\ AgNO_3} \cdot \frac{110.984\ g}{1\ mol} =$$

$$0.13873\ g$$

$$= \boxed{0.139\ g\ CaCl_2}$$

Problem:

How many mL of 3.0 M HNO_3 are needed to completely consume 2.7 g Mg?



Solution

$$Mg_{(s)} + 2\ HNO_{3\,(aq)} \rightarrow Mg(NO_3)_{2\,(aq)} + H_{2(g)}$$

$$\underbrace{g\,Mg \xrightarrow{1} mol\,Mg}_{Mg} \xrightarrow{2} \underbrace{mol \xrightarrow{3} L \xrightarrow{4} mL}_{HNO_3}$$

1 24.3050 g = 1 mol

2 1 Mg = 2 HNO_3

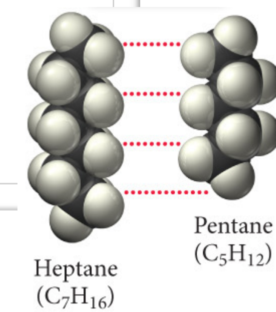
3 3.0 M HNO_3

4 1 L = 1000 mL

$$2.7g \cdot \frac{1\ mol}{24.3050g} \cdot \frac{2\ HNO_3}{1\ Mg} \cdot \frac{1\ L}{3.0\ mol} \cdot \frac{1000\ mL}{1\ L} = \boxed{74\ mL}$$

Ch13

Solution



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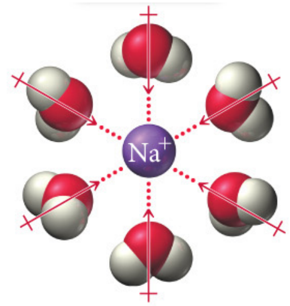

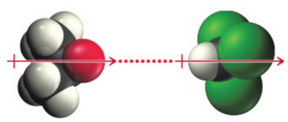

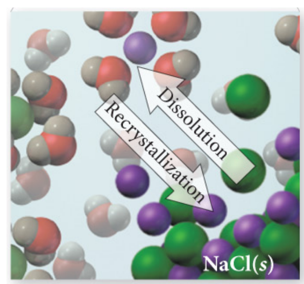

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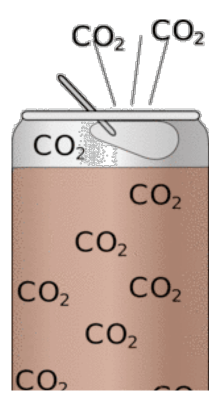

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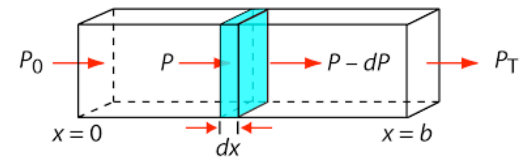

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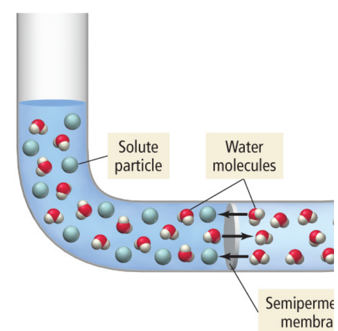
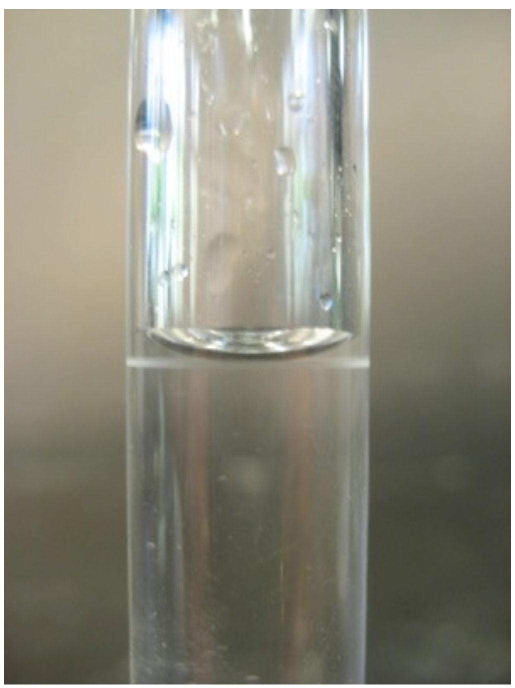
$A = \varepsilon\, l\, C$

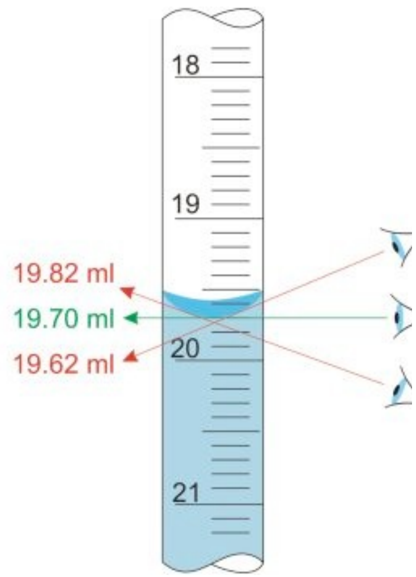
Raoult's Law
$P_A = \chi_A \bullet P^\circ$

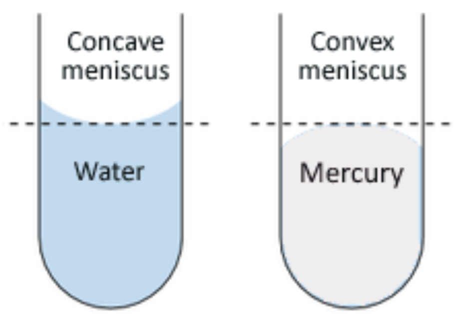


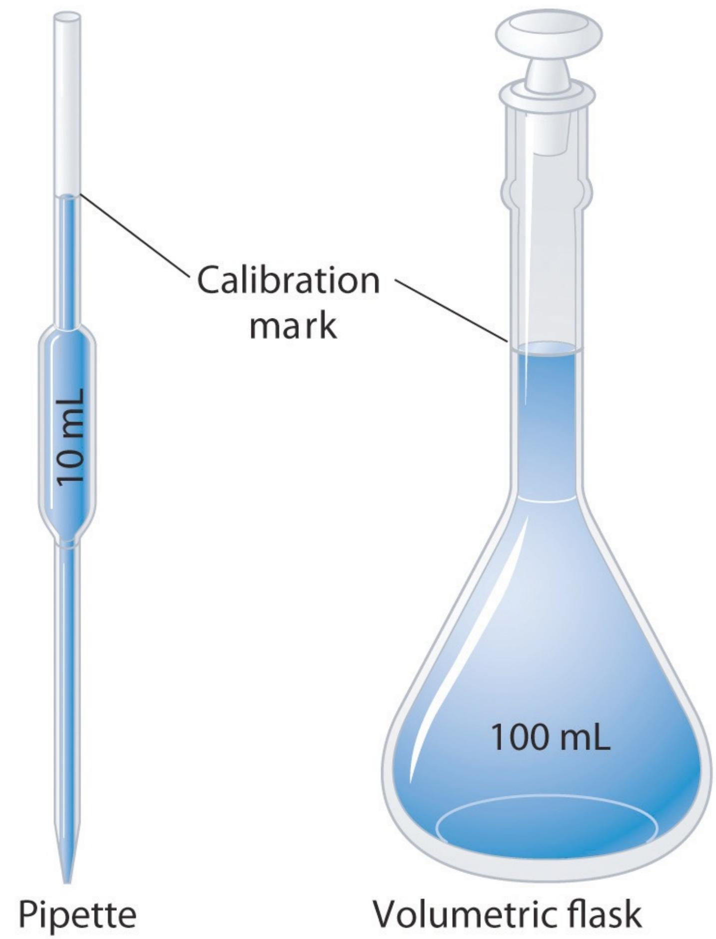
Volumetric Glassware

- Volumetric Pipets and Volumetric Flasks have a long thin neck and with a calibration mark.
- Small changes in volume make big changes in the level of the liquid allowing you to precisely measure the volume for which the device is calibrated.
- The volume is right when the meniscus of the liquid meets the calibration mark.









Dilution

- Stock solutions are solutions of known concentration.
- Most solutions are made by diluting a stock solution to a new molarity.
- Dilution just means adding more solvent.
- Dilution never changes the number of mols dissolved in the solution.
  - just the volume of the solution around them.
- Molarity and volume change with dilution, but because the mols don't change…
  - the ratio of volume to molarity is constant.

moles before = moles after

$$v_{before} M_{before} = v_{after} M_{after}$$

only when diluting!

molarity × volume = moles

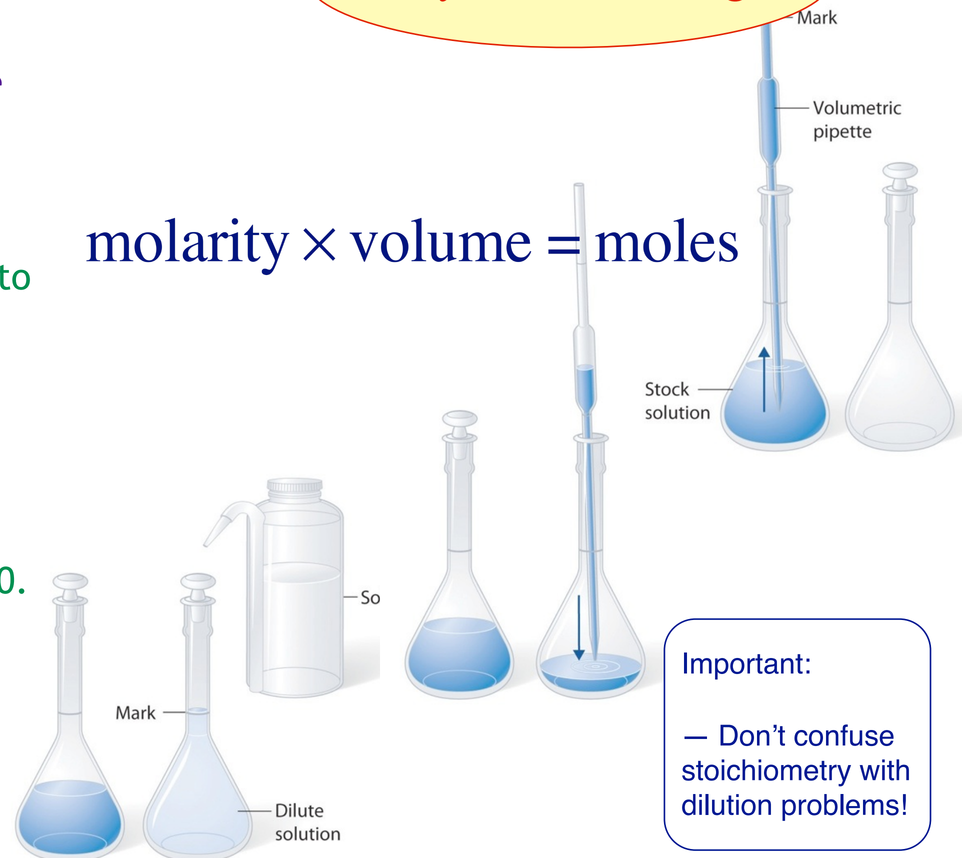
- What volume must you dilute 25 mL of 8.0 M $Ca(NO_3)_2$ to make a 2.0 M solution?

$$V_A = \frac{V_B M_B}{M_A} = \frac{8.0M \cdot 25mL}{2.0M} = 100mL$$

$(1.0 \times 10^2 mL)$

- How many mL of 6.0 M $HCl_{(aq)}$ do you need to make 200. mL of 2.0 M $HCl_{(aq)}$?

$$V_B = \frac{V_A M_A}{M_B} = \frac{200.mL \cdot 2.0M}{6.0M} = 67mL$$

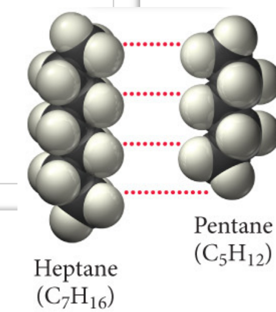


Important:

— Don't confuse stoichiometry with dilution problems!

Ch13

Solution



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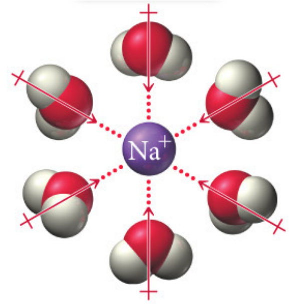

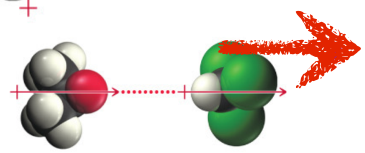

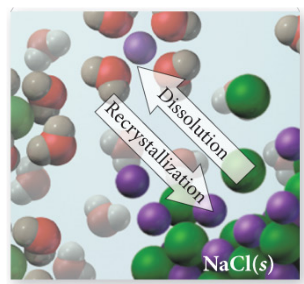

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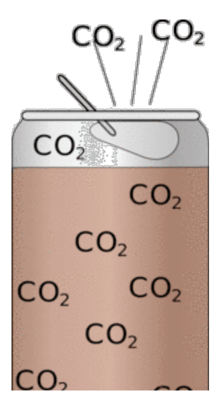

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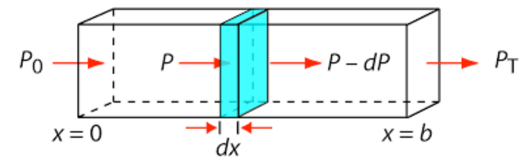

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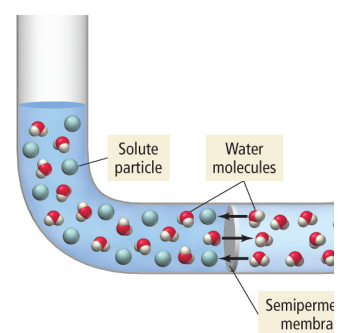
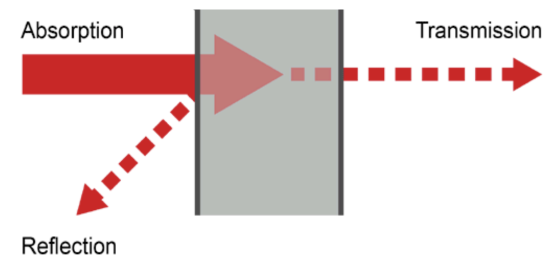
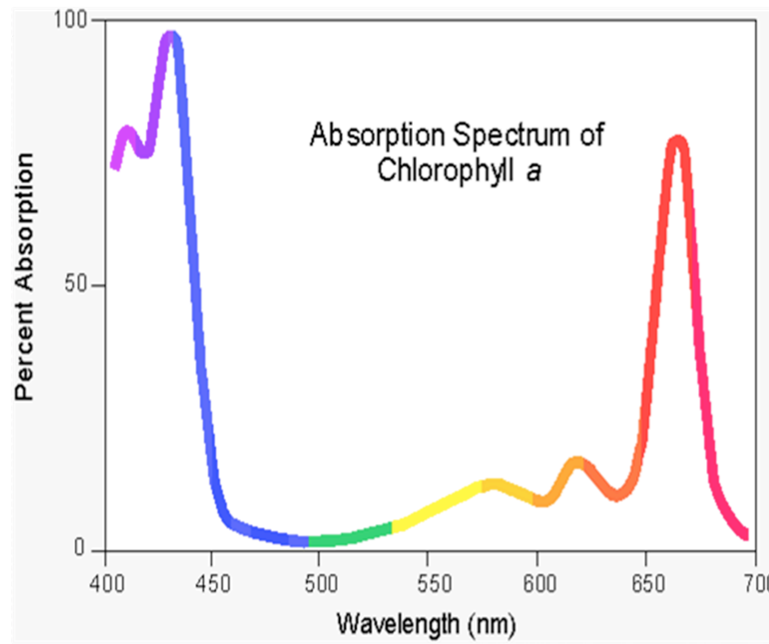
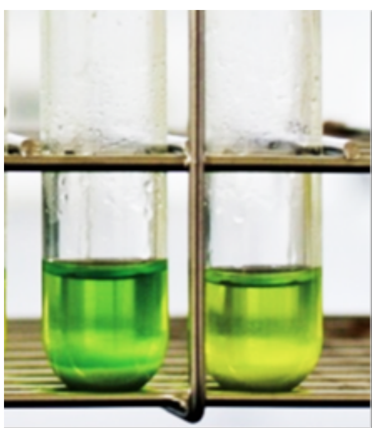


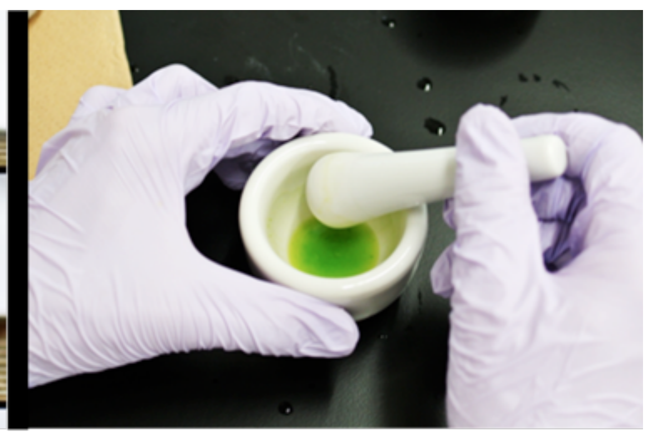
Spectroscopy

- Different substances do a better job at absorbing a particular color light.
- Chlorophyl for example really like to absorb blues and reds.
- It's transmits and reflects greens well.
- That's why plants look green and the light that passes through a chlorophyl solution is green.









Spectroscopy

- If a solute has color…
  - If it absorbs a particular color (wavelength) of light…
- … then there is a relationship between the intensity of that absorbed light and concentration.

- Spectroscopy is the measure of light absorbed or emitted by matter.
- Spectroscopy can be used to determine the concentration of a solution by measuring the light it absorbs.

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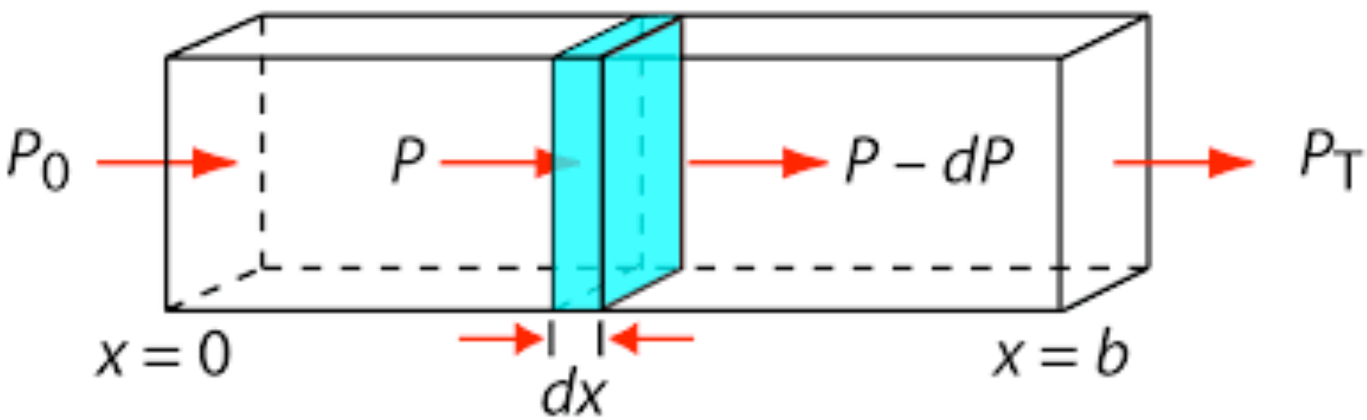

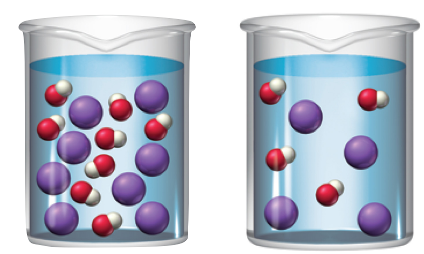

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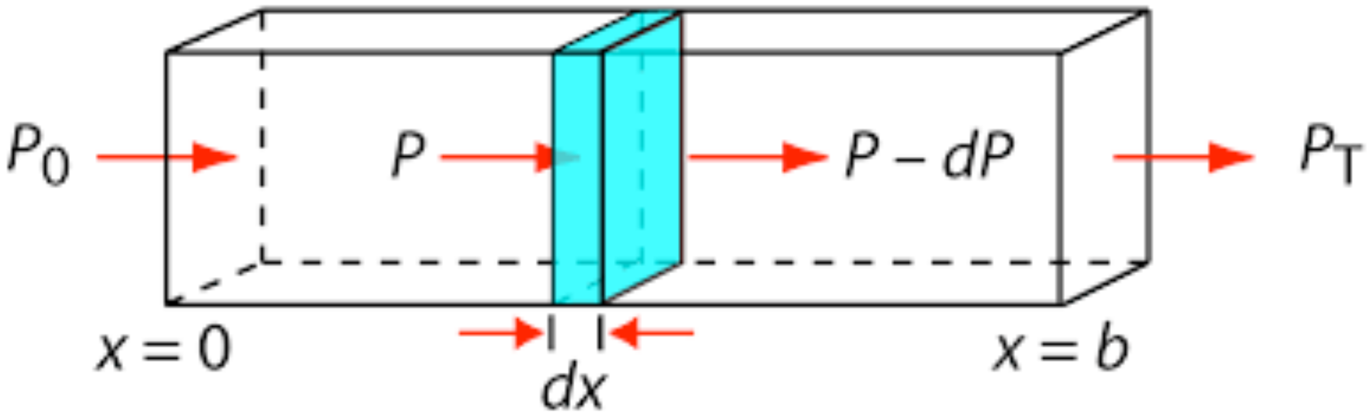

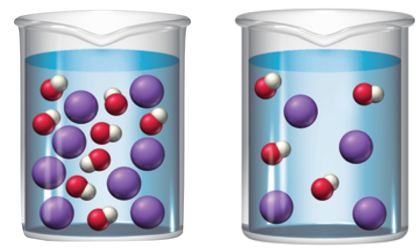

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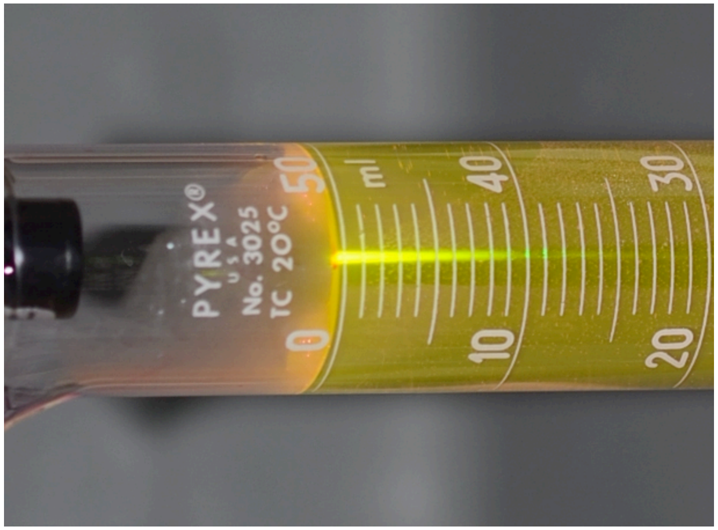
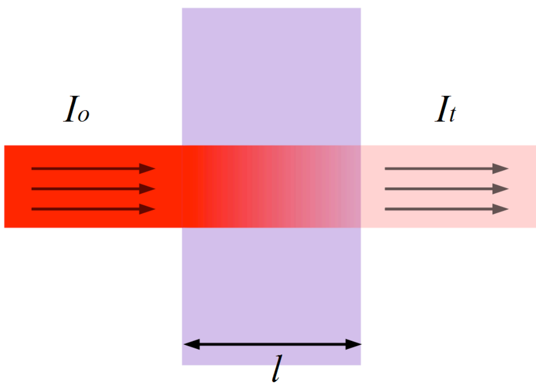
Beer's Law



- The longer the light travels through solution, the more solute it encounters so the more light get's absorbed. Absorption is relative to the path length (l)



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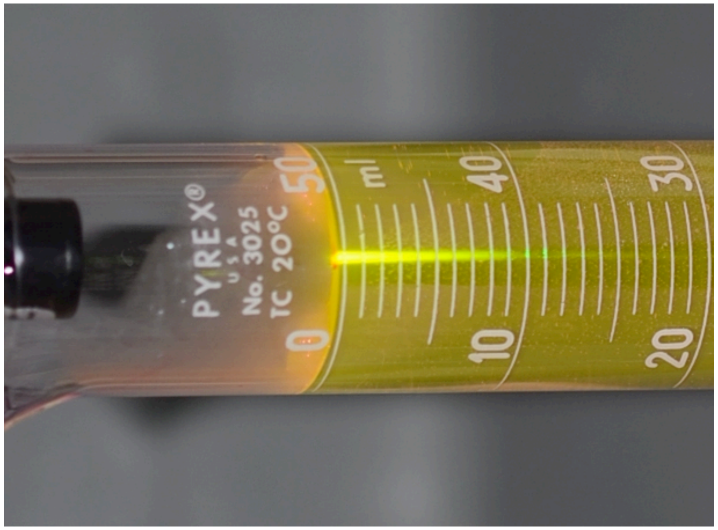

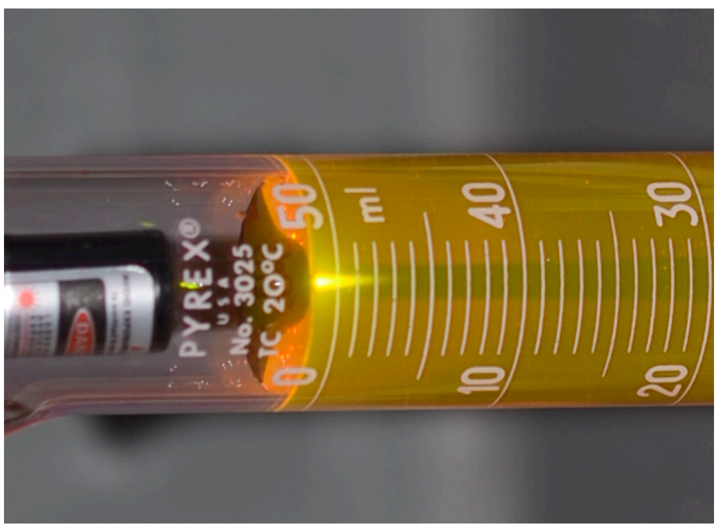

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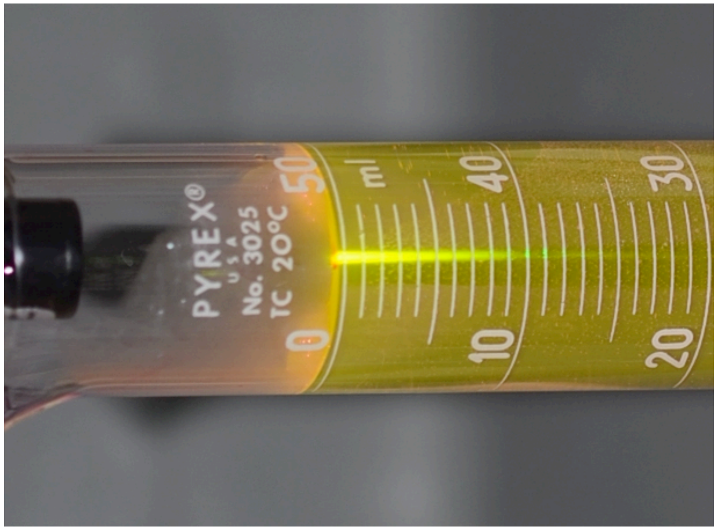

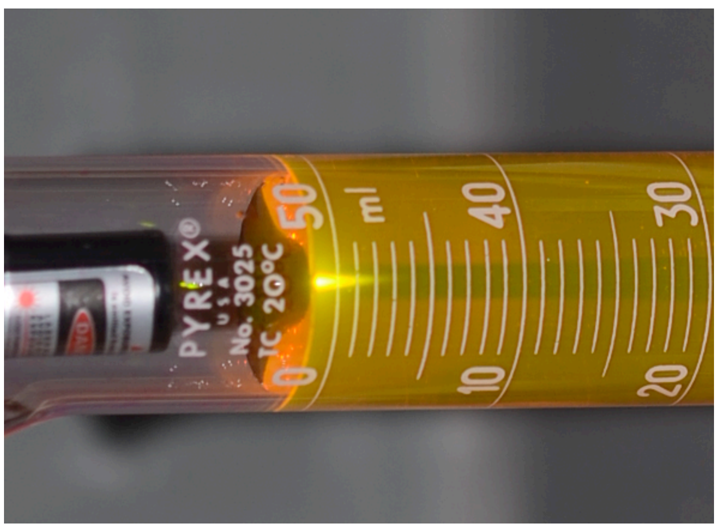

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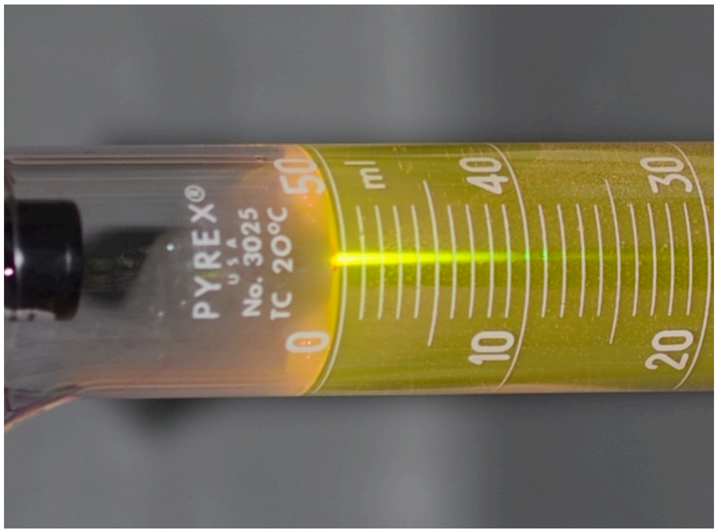
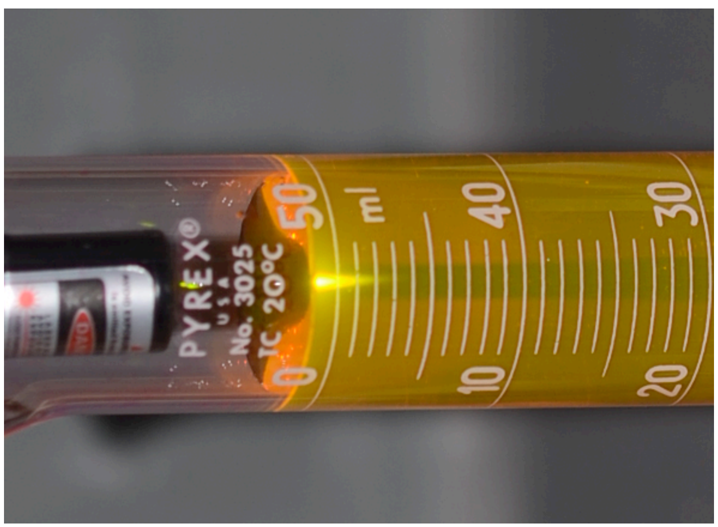
- Different substances can absorb light more or less effectively. The degree to which a substance absorbs light is it's absorptivity (ε).

    (also called molar extinction coefficient)

Beer's Law





- Absorption is relative to the path length (l)
- Absorption is relative to concentration (c)
- Absorption is relative to absorptivity (ε).

- Combining these factors gives us an empirical law that predicts absorption (A) — Beer's law.

$$A = \varepsilon l c$$

A = absorbance

ε = absorptivity

l = pathlength

c = concentration

Beer's Law

- Beer's law is a linear relationship to absorption.

$$A = l\,\varepsilon\;\; c$$

- If we use the same tube for many experiments with the same substance the length of the tube and absorptivity of the substance are constant (k).

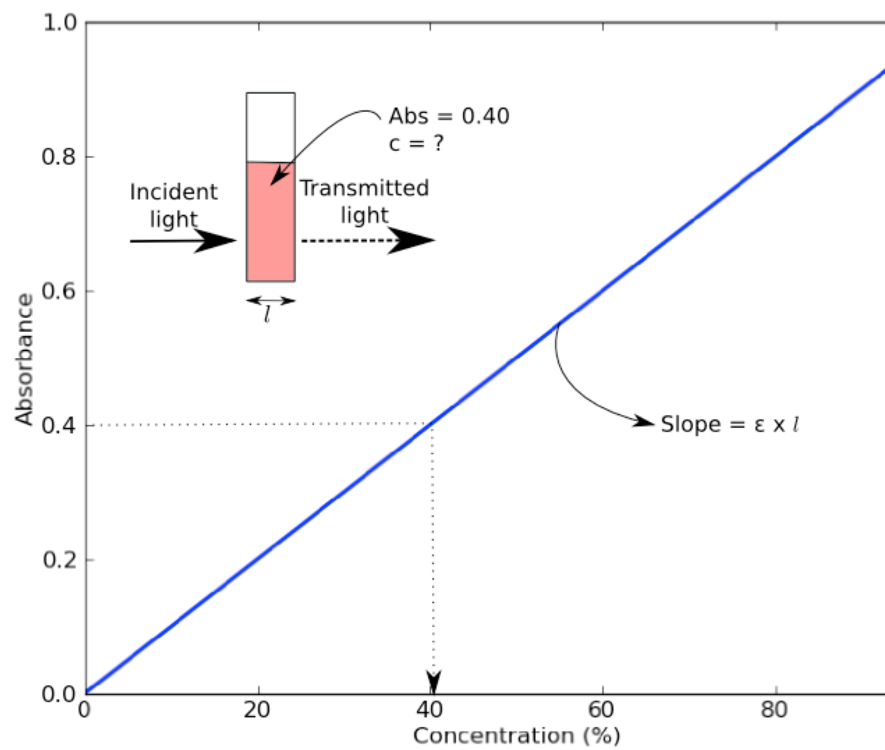
- So we can find a line that relates absorbance to concentration.

$$A = k\;\; c$$

- Absorption on the y axis, concentration on the x axis.

- If we find that line, we can then use it to know the concentration of an unknown solution from it's observed absorption.

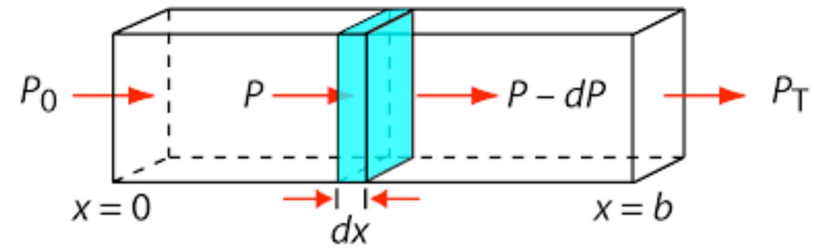
$$y = m\;\; x$$



Spectrometry

- The amount of light absorbed is effected by each slice of our sample.

- If 10% of the light passes through one slice.
- Then 10% of 10% gets though two.
  - … and 10% of 10% of 10% gets through three.



$$Transmitted = \frac{P}{P_0}$$

$$Absorbed = \log\left(\frac{P_0}{P_T}\right) = \log\left(\frac{1}{T}\right)$$

Absorbance (optical density)	Light Transmittance %
0	100%
1	10
2	1
3	0.1
4	0.01
5	0.001
6	0.0001

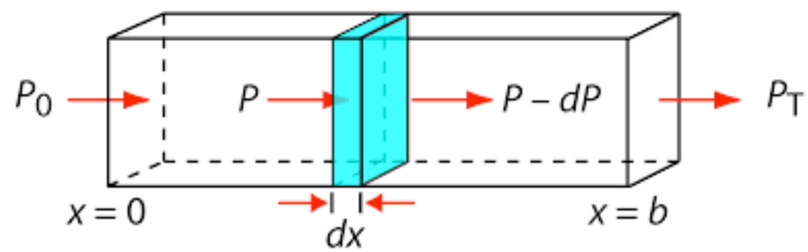
$$10^{Absorbed} = \frac{1}{T}$$

$$10^{Absorbed} = \frac{100}{\%T}$$

Spectrometry

- The mount absorbed is logarithmically related to what we see coming out the other side.

$$10^{\text{Absorbed}} = 100\% \div \%\text{Transmitted}$$

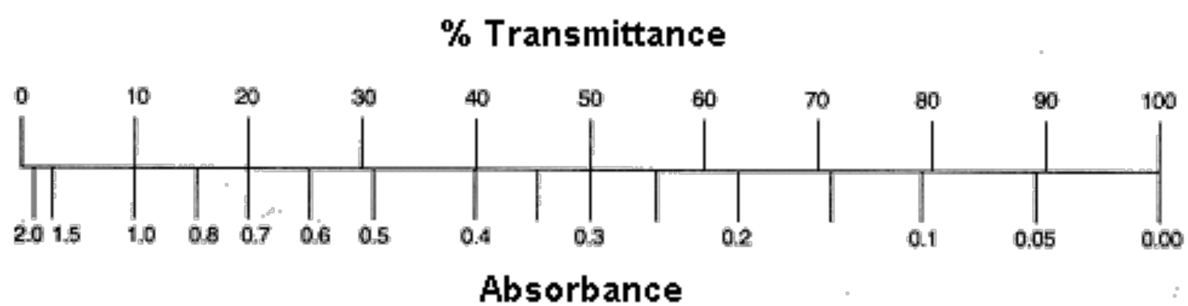


- If we take the log of both sides we get

$$\log\left(10^{\text{Absorbed}}\right) = \log\left(100\% \div \%\text{Transmitted}\right)$$

$$\log\left(10^{\text{Absorbed}}\right) = \log(100) - \log(\%\text{Transmitted})$$

$$\text{Absorbed} = 2 - \log(\%T)$$



Ch13

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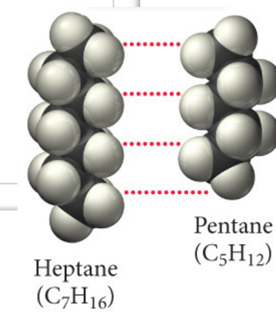

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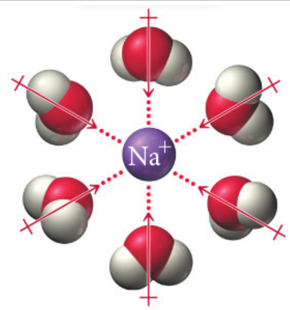

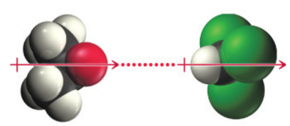

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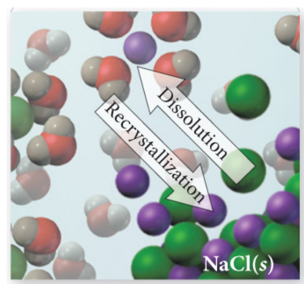

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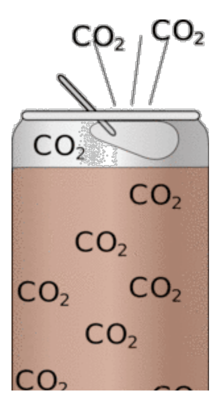

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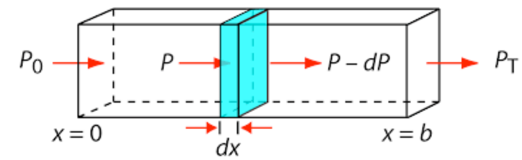

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- Dispersions
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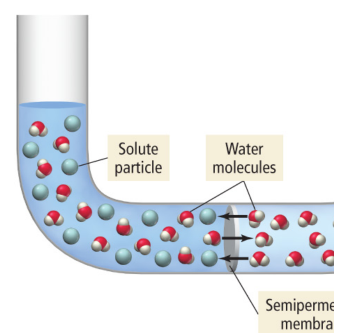
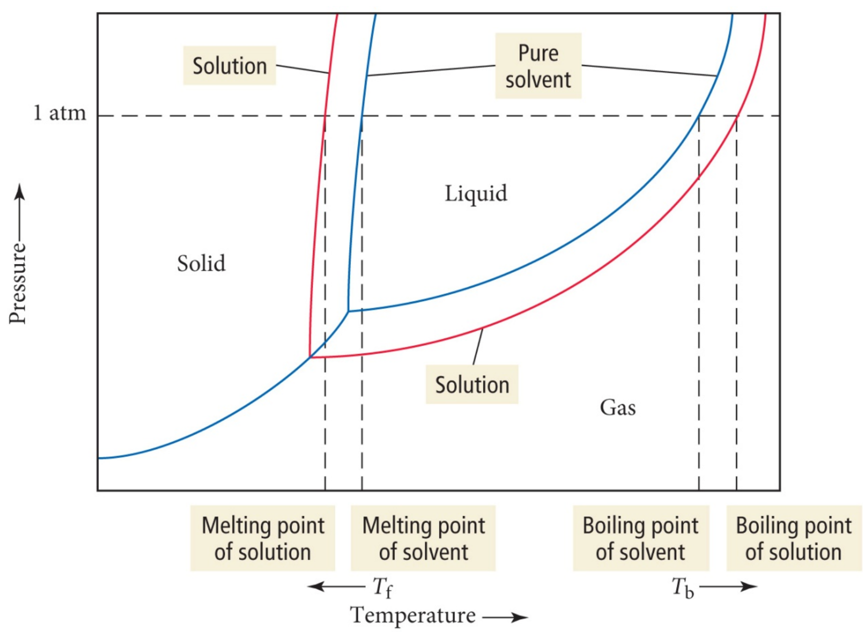


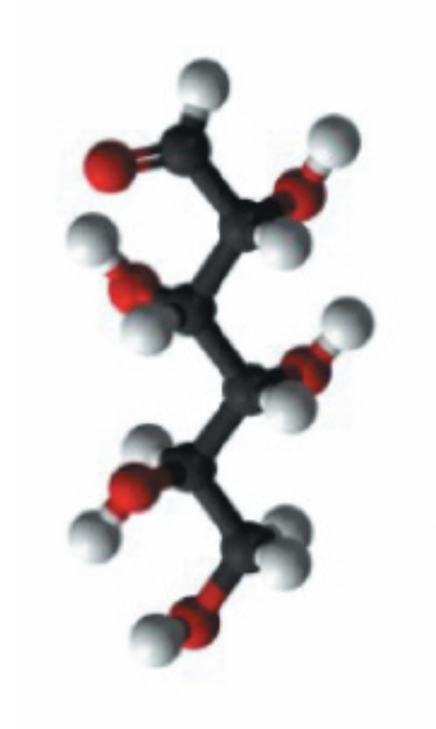
Colligative Properties

- Colligative properties are properties whose value depends only on the number of solute particles, and *not on what substance*.
    - Value of the property depends on the concentration of the solution.

- The difference in the value of the property between the solution and the pure substance is generally related to the different attractive forces and solute particles occupying solvent molecules positions.
- Melting point and boiling point are colligative properties.



Consider two solutions: (1) 10 g of glucose ($C_6H_{12}O_6$) in 1 L of water and (2) 10 g of sucrose ($C_{12}H_{22}O_{11}$) in 1 L of water. Which of the following is true?



a) The glucose solution has the higher vapor pressure.
b) The sucrose solution has the higher vapor pressure.
c) Both solutions have the same vapor pressure.
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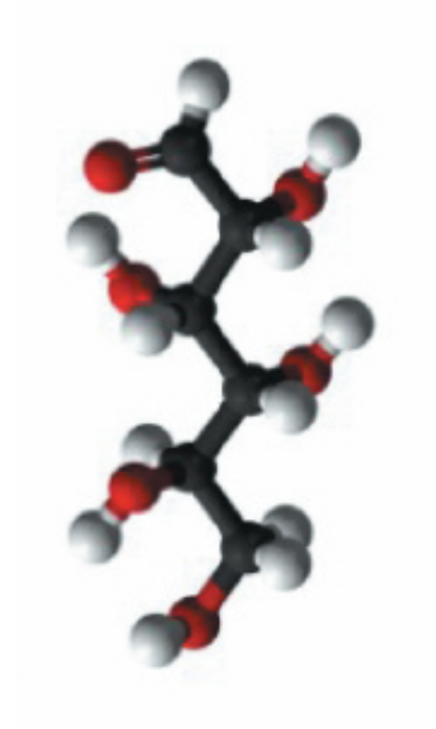
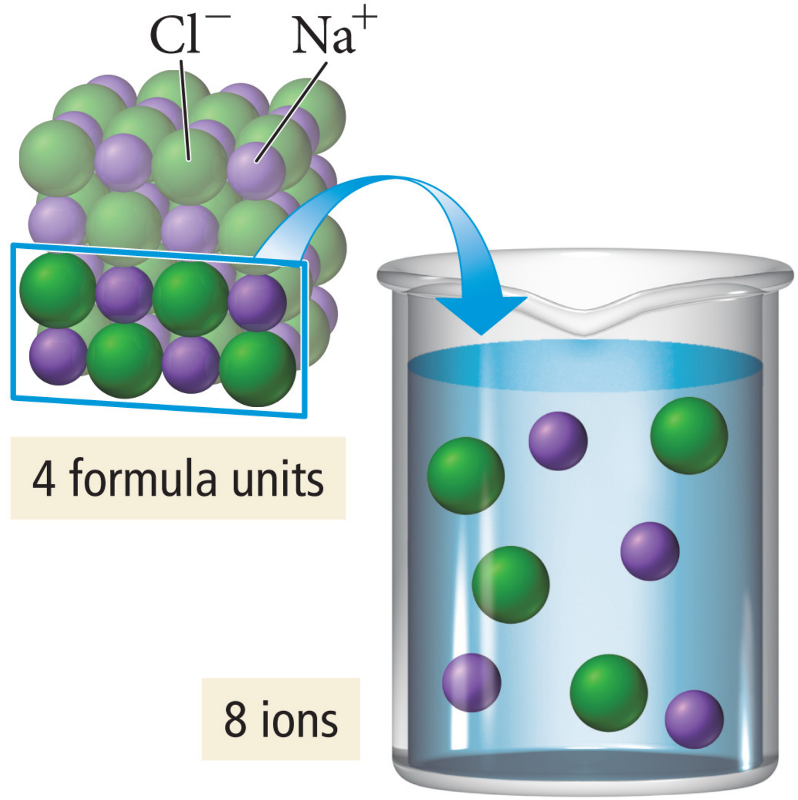
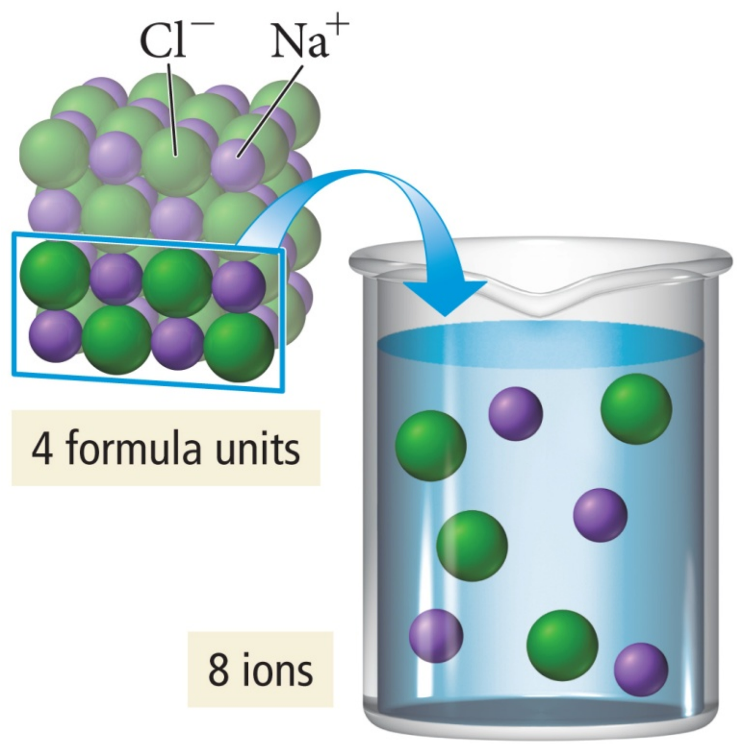


Electrolytic Power in Colligative Properties



Which of the following has the highest boiling point?

a) 0.05 *m* NaCl
b) 0.05 *m* $C_6H_{12}O_6$
c) 0.02 *m* $Al(NO_3)_3$
d) 0.005 *m* CO_2
e) 0.02 *m* NH_4Cl



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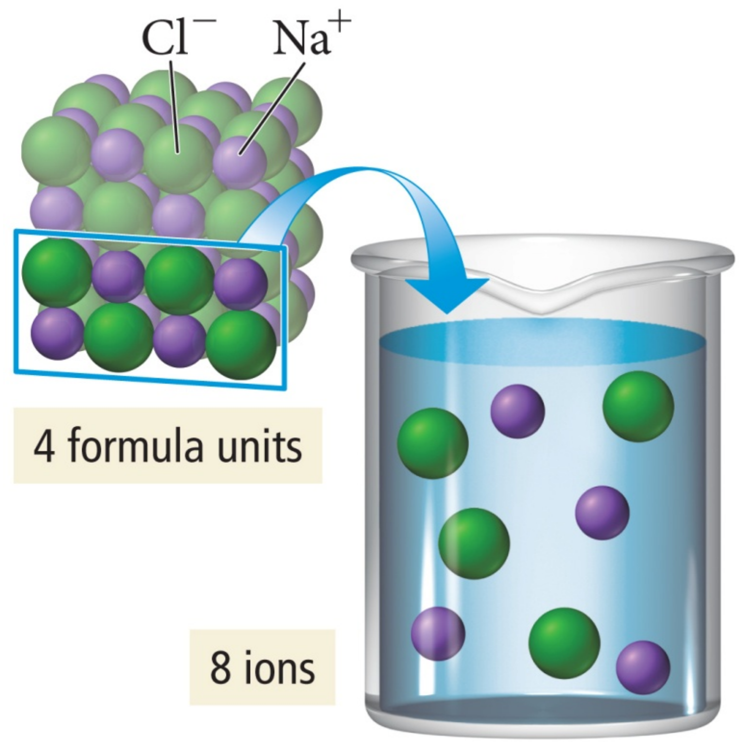

Ch13

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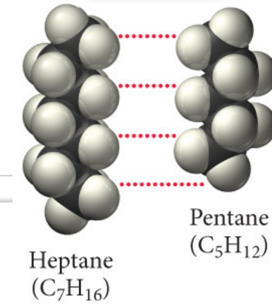

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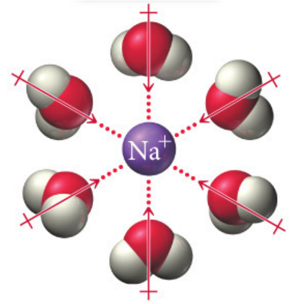

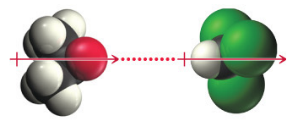

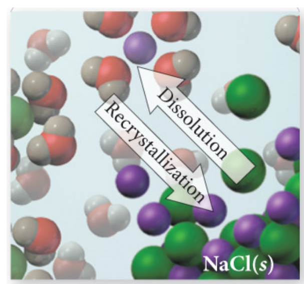

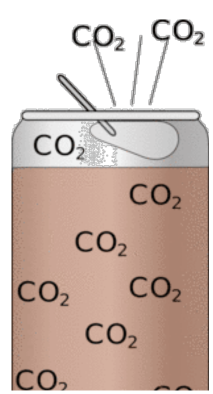

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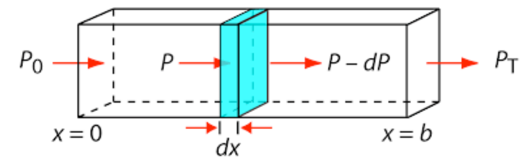
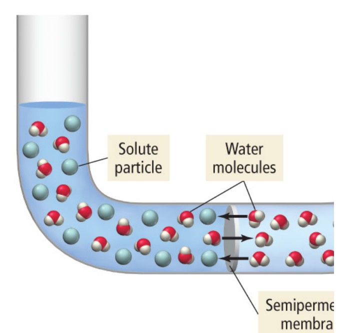
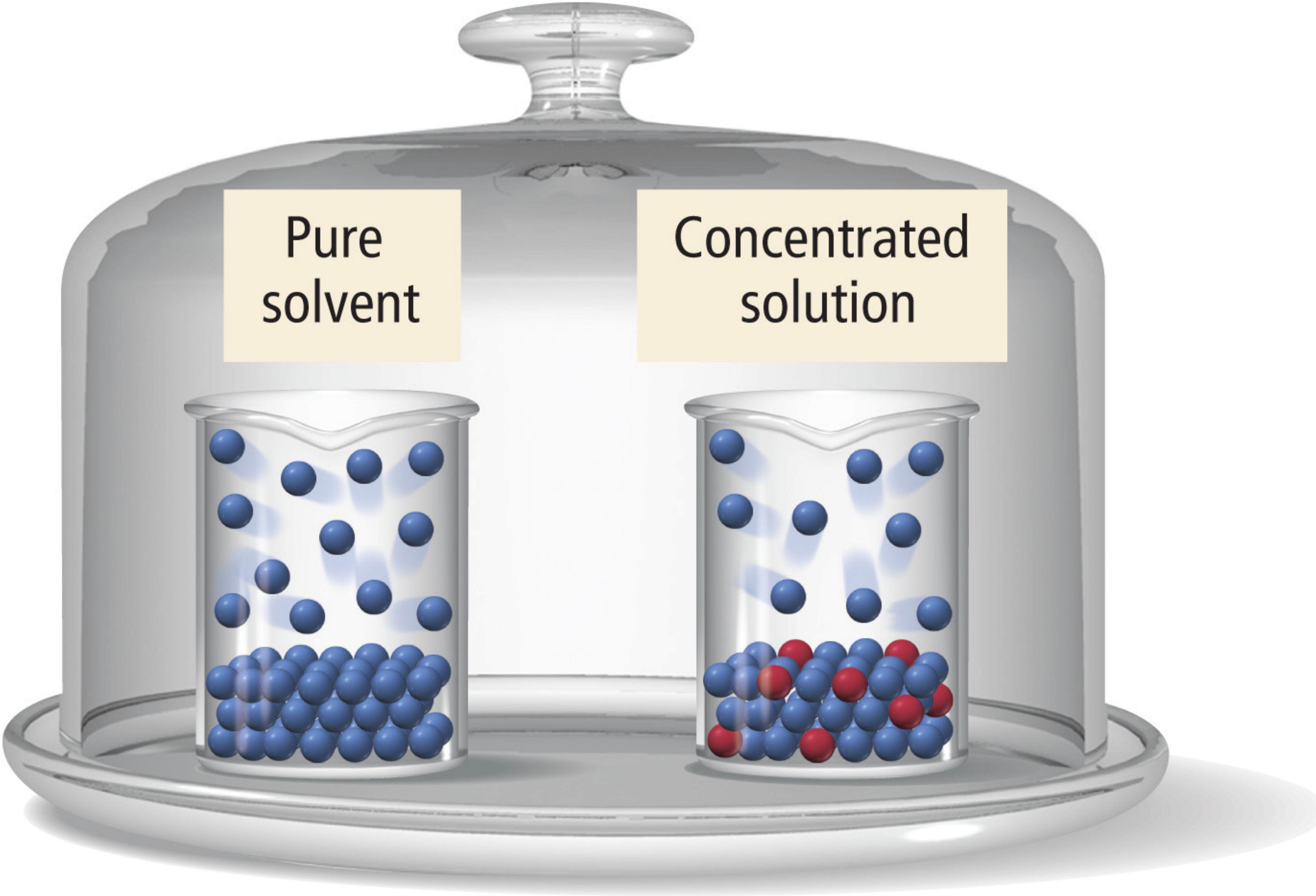
  - Tyndall Effect

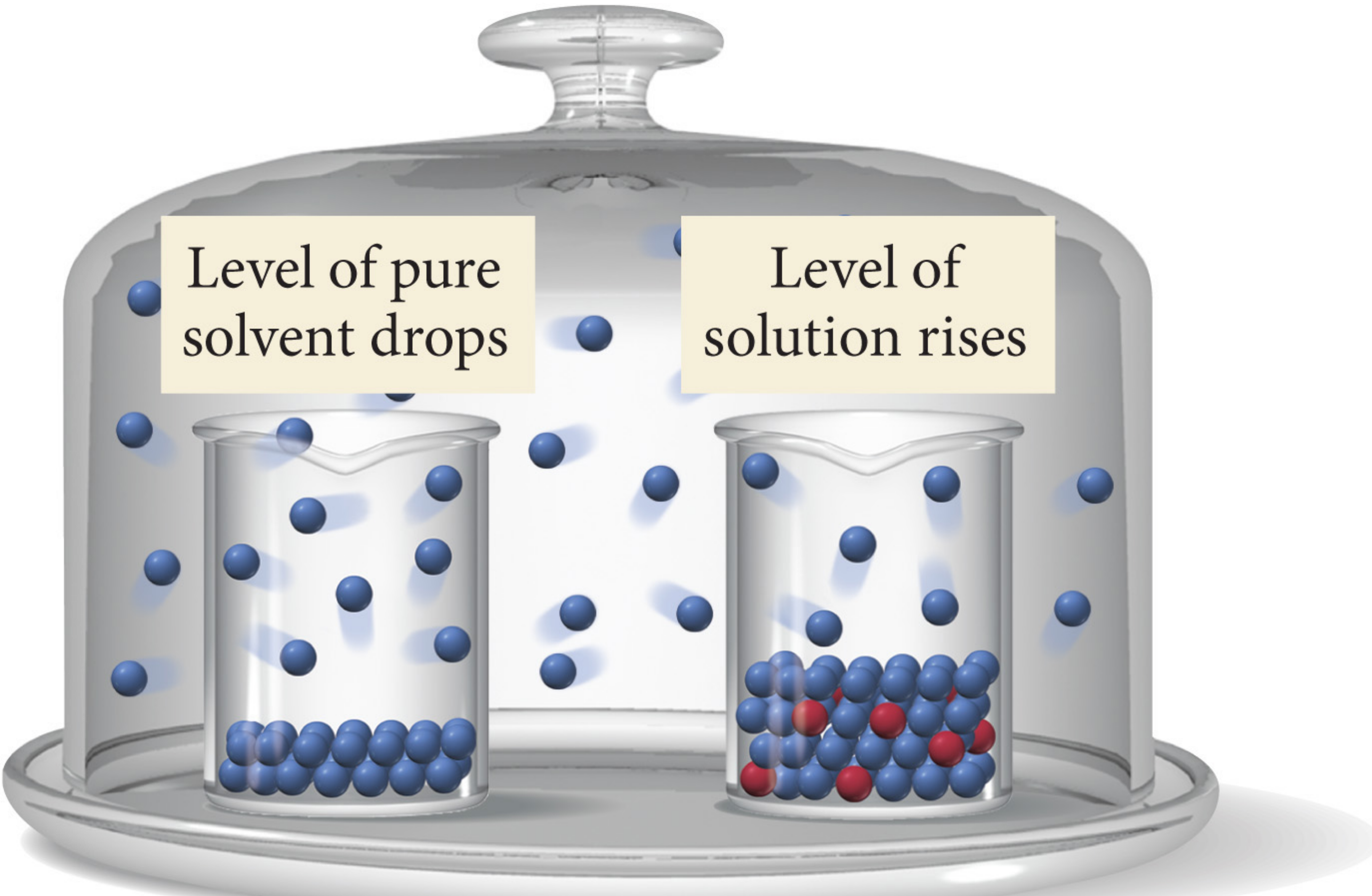


Beer's Law
$A = BC$
$A = \varepsilon \, l \, C$

Raoult's Law
$P_A = \chi_A \cdot P^\circ$




Pure solvent
Concentrated solution


Level of pure
solvent drops
Level of
solution rises

Raoult's Law

A solution of benzene (C6H6) and toluene (C7H8) is 25.0% benzene by mass. The vapor pressures of pure benzene and pure toluene at 25 °C are 94.2 torr and 28.4 torr, respectively. Assuming ideal behavior, calculate the following:

a. The vapor pressure of each of the solution components in the mixture.

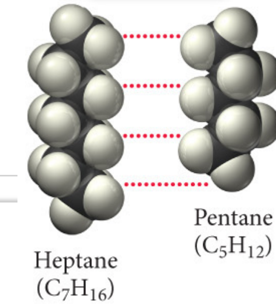
b. The total pressure above the solution.

Raoult's Law

$$P_A = \chi_A \bullet P^\circ$$

Ch13

Solution



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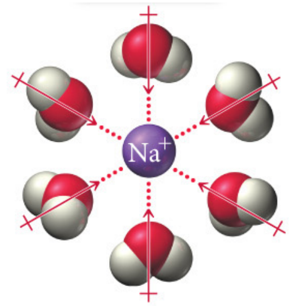

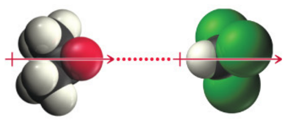

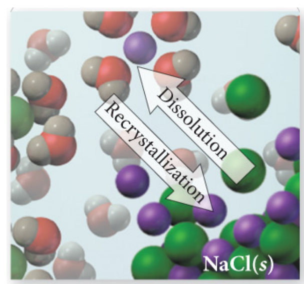

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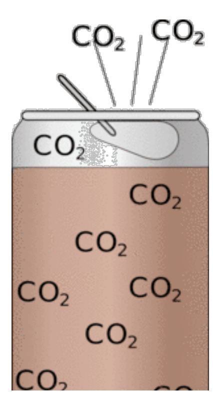
$[A] = k_A P_A$



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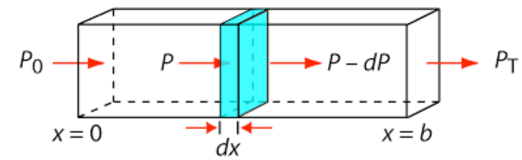

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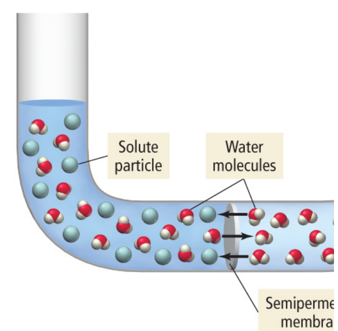
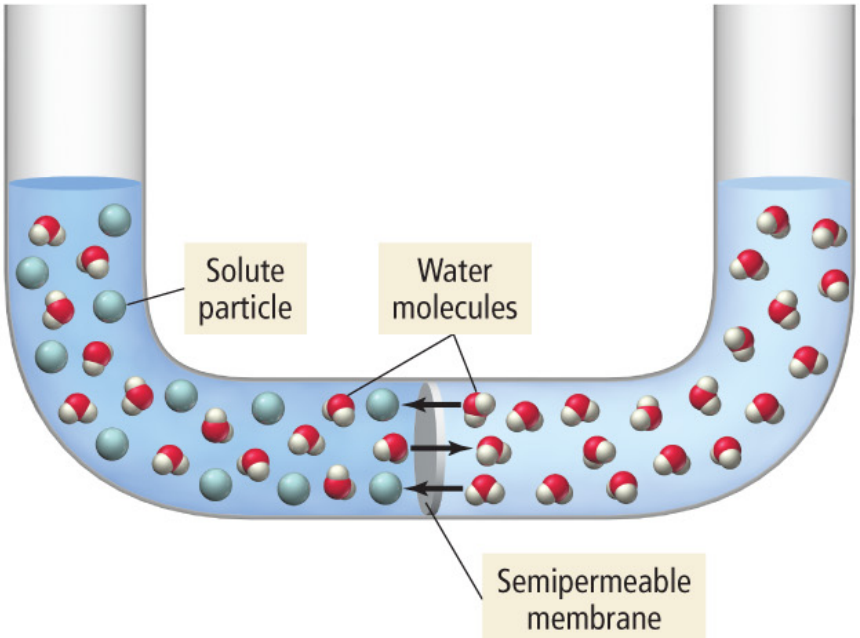
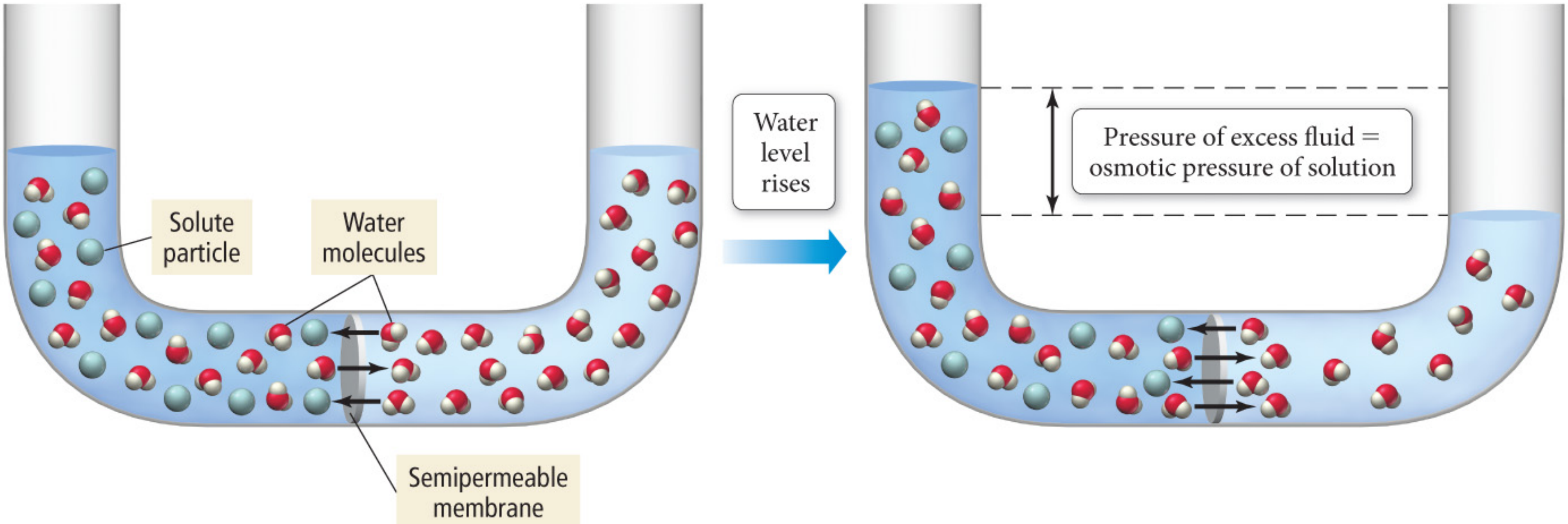


Osmosis



- A semi-permeable membrane can be used to allow solvent molecules to flow but block larger solute molecules.
- Osmosis is the tendency of solvent to favor flowing to the more concentrated side of that membrane.

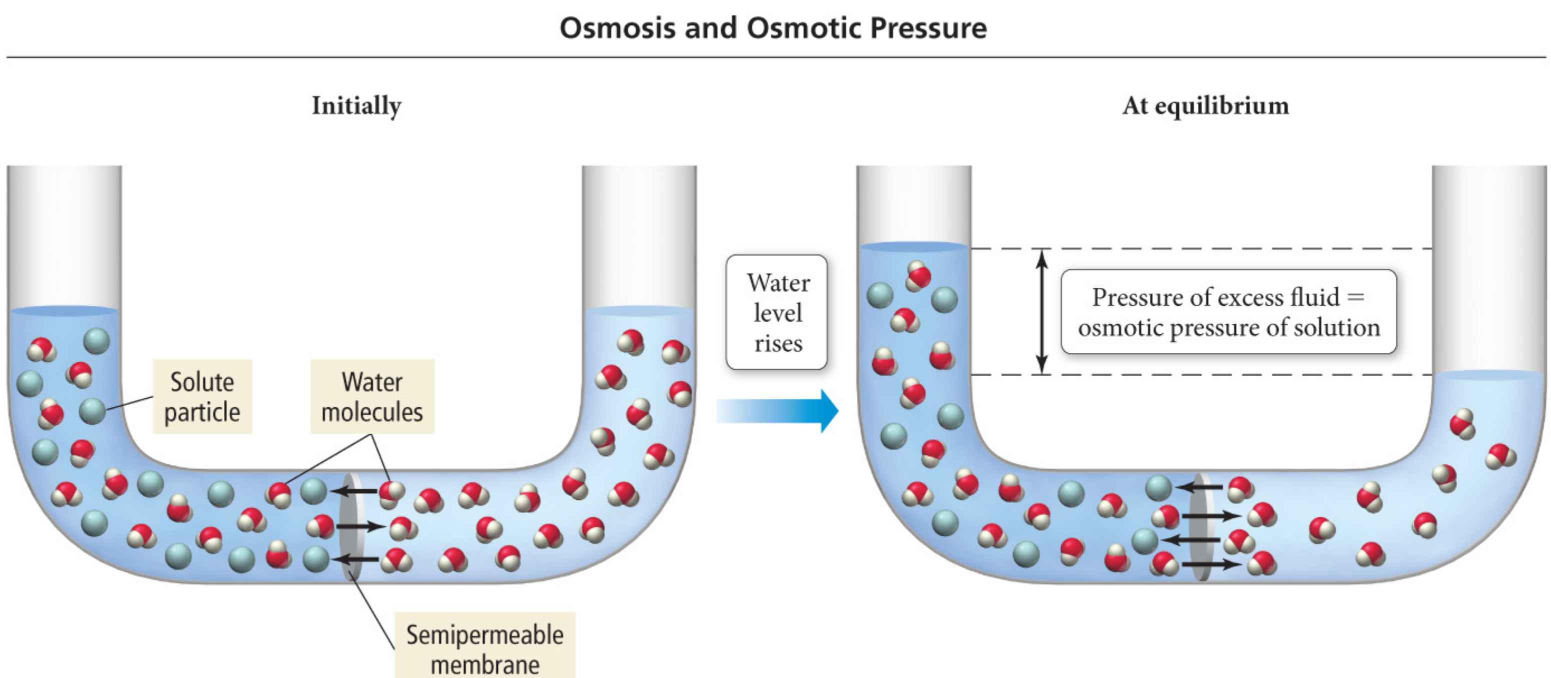
Osmotic Pressure



- A semi-permeable membrane can be used to allow solvent molecules to flow but block larger solute molecules.
- Osmosis is the tendency of solvent to favor flowing to the more concentrated side of that membrane.
- That tendency can be measured by the pressure required to offset the difference in flow.
- This is osmotic pressure.

Which of the following solutions will have the largest osmotic pressure?

a) 0.50 m $C_6H_{12}O_6$
b) 0.50 m NaCl
b) 0.50 m K_2SO_4
d) 0.50 m $FeCl_3$



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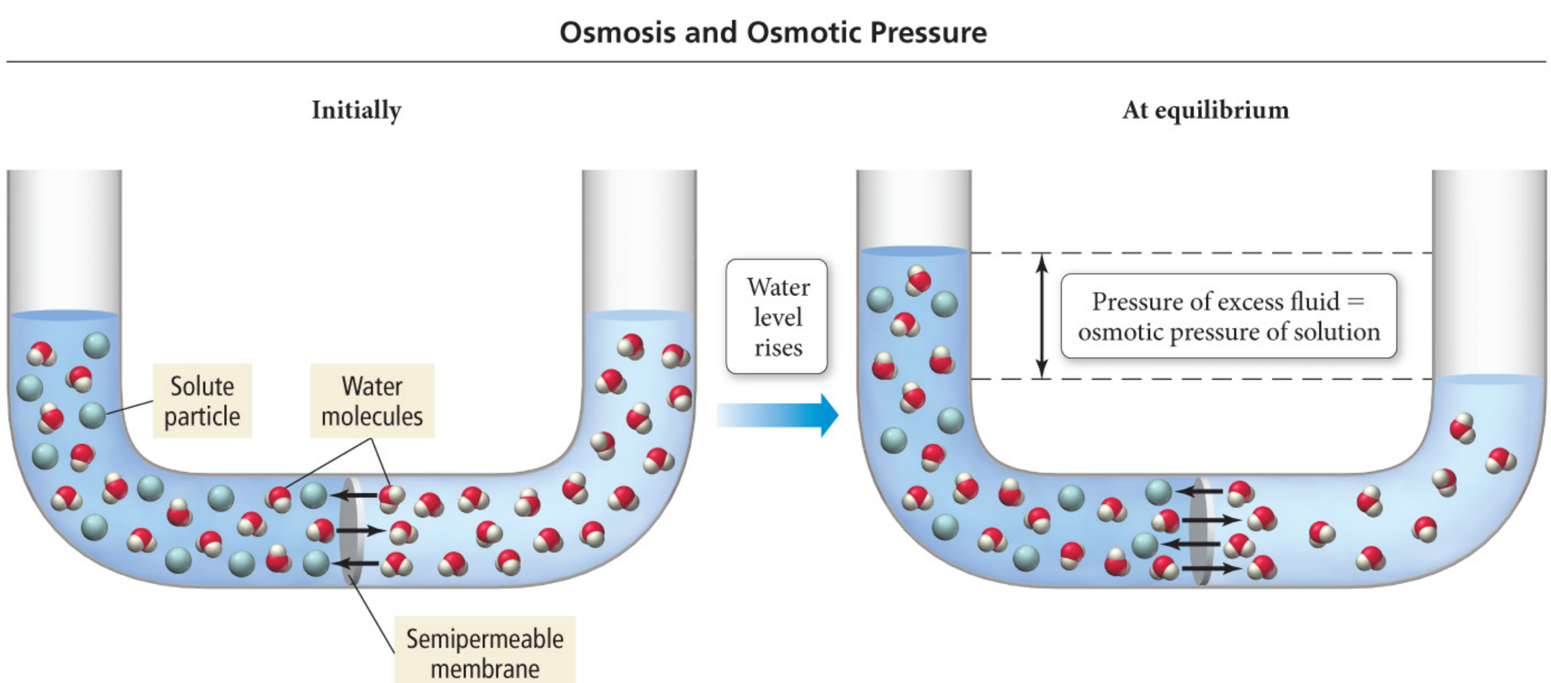
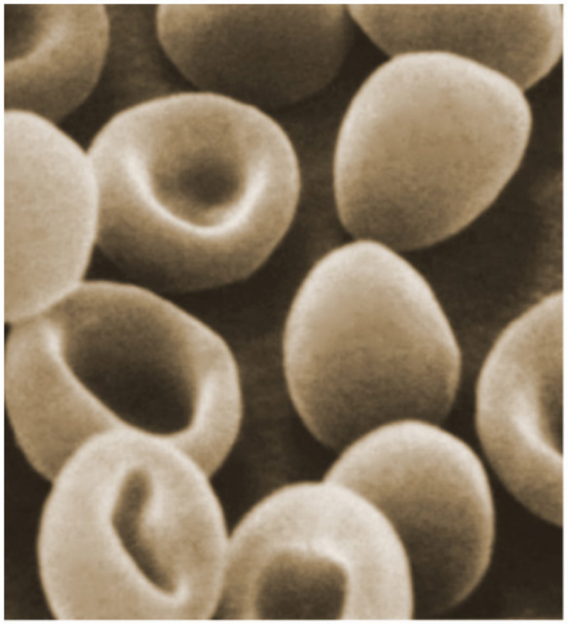
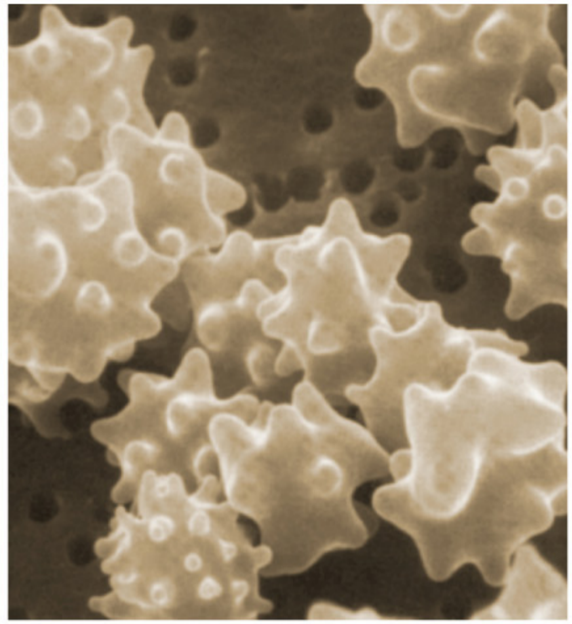


Isotonic



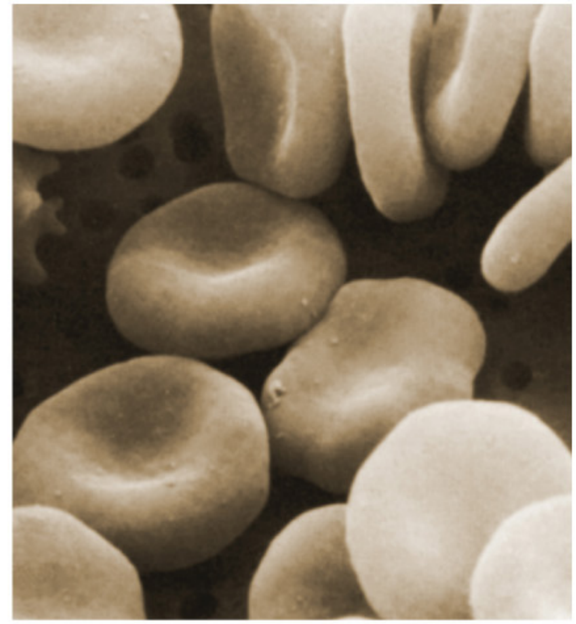
(a) Normal red blood cells

Hypertonic



(b) Shriveled red blood cells

Hypotonic



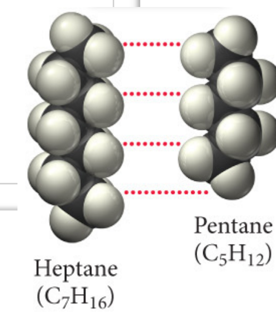
(c) Swollen red blood cells

Iso-, hyper-, or hypo- in which system is the osmotic pressure inside the cell greater than the pressure outside it?

In which system is the concentration of electrolytes greater outside the cell than inside it?

Ch13

Solution



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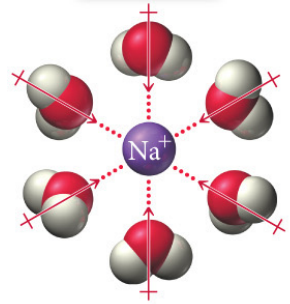

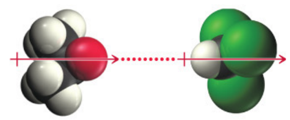

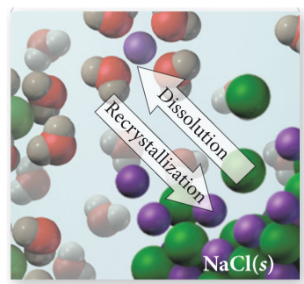

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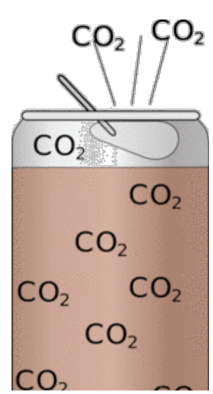

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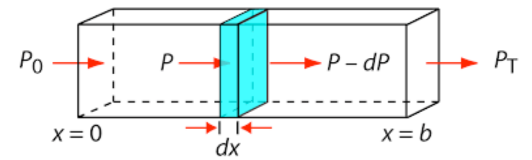

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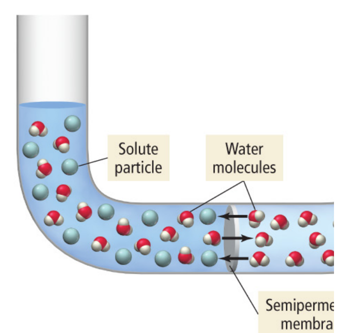
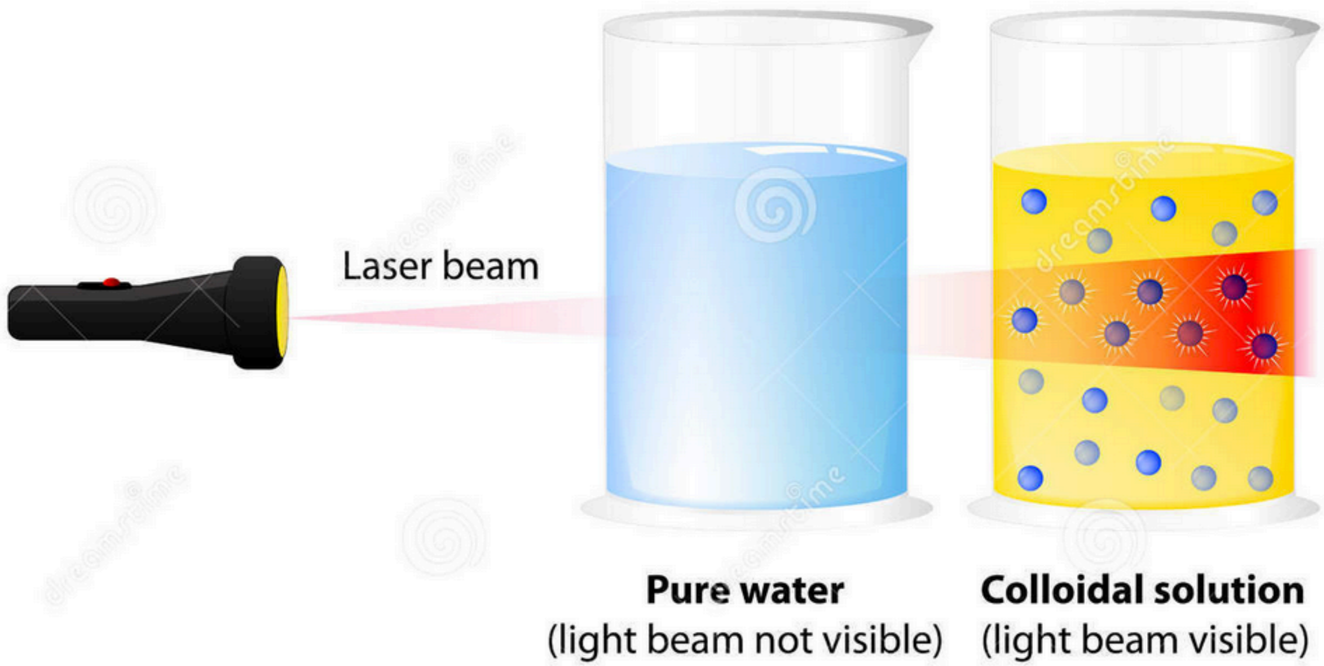


Tyndall Effect

- Colloids scatter light, so a light beam becomes visible in a colloid.
- You can identify a colloid or lightly dispersed mixture with the Tyndall Effect (does not occur for solutions or pure substances)



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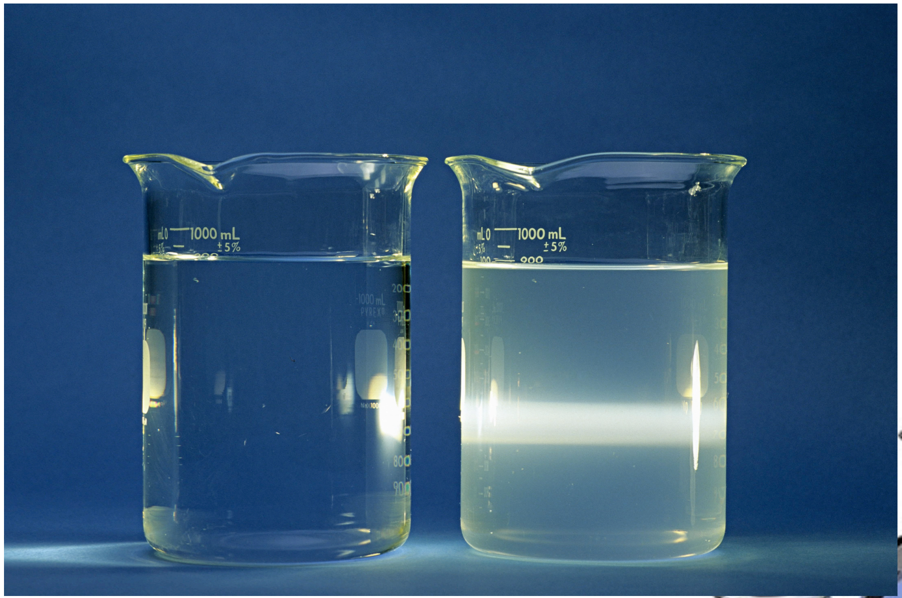

Ch13

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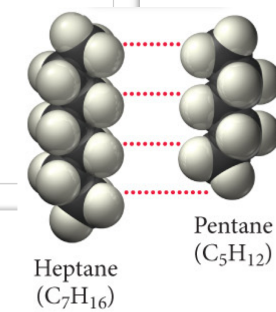

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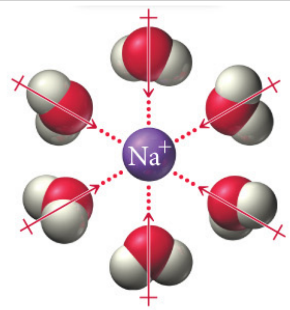

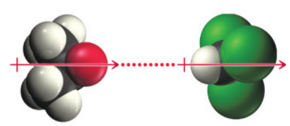

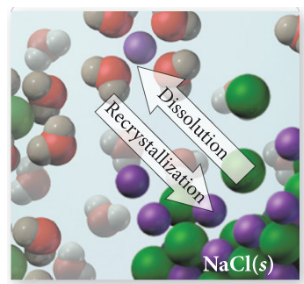

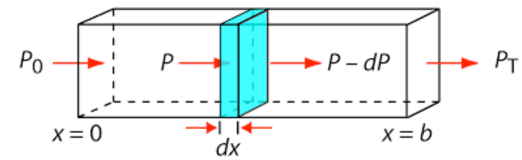

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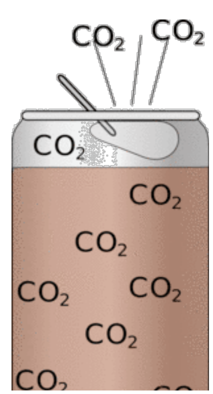

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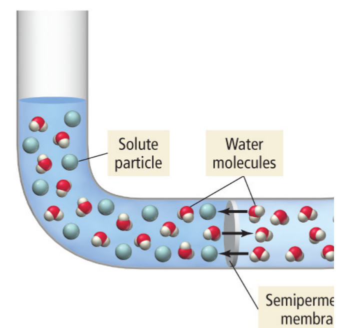

Questions?