What evidence is there that ocean water is a solution of a solid in a liquid?

What evidence, shown by plants and animals, proves that ocean water is a solution of gas in a liquid?
The Unique Properties of Water
Chapter 15 Solutions

I. What are Solutions (15.1)

A. **Solution** - homogeneous mixture made up of individual molecules, atoms or ions.

B. **Solute** - the substance being **dissolved**

C. **Solvent** - the substance **doing the dissolving**

D. **Soluble** - substance that **dissolves** in a solvent

E. **Insoluble** - substance that **does not dissolve** in a solvent

F. **Immiscible** – liquids that are **NOT soluble** in each other

G. **Miscible** – two liquids that **are soluble** in each other
H. Types of solutions

1. Aqueous solution - water is the solvent

2. Tincture - alcohol is the solvent
H. Types of solutions

3. Alloys - solid solution of two or more metals
H. **Types of solutions**

- **Amalgam** - an alloy in which one metal is Hg
<table>
<thead>
<tr>
<th>Type of solution</th>
<th>Example</th>
<th>Solvent</th>
<th>Solute</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gas</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Gas in gas</td>
<td>Nitrogen (gas)</td>
<td>Oxygen (gas)</td>
<td></td>
</tr>
<tr>
<td><strong>Liquid</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Gas in liquid</td>
<td>Water (liquid)</td>
<td>Carbon dioxide (gas)</td>
<td></td>
</tr>
<tr>
<td>Gas in liquid</td>
<td>Ocean water</td>
<td>Water (liquid)</td>
<td>Oxygen gas (gas)</td>
</tr>
<tr>
<td>Liquid in liquid</td>
<td>Water (liquid)</td>
<td>Ethylene glycol (liquid)</td>
<td></td>
</tr>
<tr>
<td>Liquid in liquid</td>
<td>Water (liquid)</td>
<td>Acetic acid (liquid)</td>
<td></td>
</tr>
<tr>
<td>Solid in liquid</td>
<td>Ocean water</td>
<td>Water (liquid)</td>
<td>Sodium chloride (solid)</td>
</tr>
<tr>
<td><strong>Solid</strong></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Liquid in solid</td>
<td>Silver (solid)</td>
<td>Mercury (liquid)</td>
<td></td>
</tr>
<tr>
<td>Solid in solid</td>
<td>Iron (solid)</td>
<td>Carbon (solid)</td>
<td></td>
</tr>
</tbody>
</table>
What happens in an Aqueous Solution?

A. Electrolytes - any substance that, when it is dissolved in a solution, will conduct an electric current by means of movement of ions
<table>
<thead>
<tr>
<th>Substance</th>
<th>dissociation in water</th>
<th>1. NaCl(_{(s)})</th>
<th>2. HCl(_{(g)}) + H(_2)O</th>
<th>3. NaOH(_{(s)})</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCl(_{(aq)})</td>
<td>Yes</td>
<td>Yes</td>
<td>No</td>
<td>No</td>
</tr>
<tr>
<td>HCl(_{(aq)})</td>
<td>Yes</td>
<td>Yes</td>
<td>No</td>
<td>No</td>
</tr>
<tr>
<td>HC(_2)H(_3)O(_2)(aq)</td>
<td>No</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaOH(_{(aq)})</td>
<td>Yes</td>
<td>tap H(_2)O</td>
<td>Yes</td>
<td>dH(_2)O</td>
</tr>
</tbody>
</table>

**Strong Electrolytes**

- Substance that completely or largely dissociates in water to produce ions and conduct electricity.

**Hydronium ion** (H\(_3\)O\(^{+1}\)), H\(^{+1}\) covalently bonded to water molecule.

1. NaCl\(_{(s)}\) + H\(_2\)O → Na\(^{+1}\)(aq) + Cl\(^{-1}\)(aq)  
   **Dissociation**

2. HCl\(_{(g)}\) + H\(_2\)O → H\(_3\)O\(^{+1}\)(aq) + Cl\(^{-1}\)(aq)  
   **Ionization**

3. NaOH\(_{(s)}\) + H\(_2\)O → Na\(^{+1}\)(aq) + OH\(^{-1}\)(aq)  
   **Dissociation**
Weak (or non) electrolyte - a substance that experiences only a small degree of dissociation (or none) in an aqueous solution, poor (or not a) conductor of electricity

1. \( \text{HC}_2\text{H}_3\text{O}_2(aq) + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^{+1}(aq) + \text{C}_2\text{H}_3\text{O}_2^{-1}(aq) \)

2. \( \text{H}_2\text{O}(l) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_3\text{O}^{+1}(aq) + \text{OH}^{-1}(aq) \)
II. Solvation in Aqueous Solutions

A Model of the Dissolving of NaCl

- Remember that ionic solids are composed of a three-dimensional network of positive and negative ions, which form strong ionic bonds.
II. Solvation in Aqueous Solutions

- The process by which the charged particles in an ionic solid separate from one another is called **dissociation**.
Hydrated Ion
Dissociation Equation:

\[ \text{NaCl} \ (s) \rightarrow \text{Na}^{+1} \ (aq) + \text{Cl}^{-1} \ (aq) \]

- **Solvation** – the process of surrounding solute particles with solvent particles to form a solution.
Factors that affect rate of solvation — must increase collisions between solute and solvent

1. Agitating the mixture
Factors that affect rate of solvation – must increase collisions between solute and solvent

1. Increasing the surface area
3. Increasing the temperature which increasing the kinetic energy
dissolving sugar for syrup
III. Solubility (15.1)

- measure of how much solute dissolves in a given amount of solvent.
- usually expressed as a quantity of solute per definite amount of solvent at a particular temperature (ex. 36g NaCl/100g H₂O at 20°C)
Unsaturated ➤ the amount of solute dissolved is less than the maximum that could be dissolved

\[ \text{NaCl}_\text{(s)} \rightarrow \text{Na}^{+1}_\text{(aq)} + \text{Cl}^{-1}_\text{(aq)} \]
Saturated solution holds the maximum amount of solute per amount of the solution under the given conditions.

\[ \text{Saturated } \text{solution} \Rightarrow \text{solute per amount of solution} \]

\[ \text{NaCl}_\text{(s)} \rightleftharpoons \text{Na}^{+1}_\text{(aq)} + \text{Cl}^{-1}_\text{(aq)} \]
III. Solubility (15.1)

Supersaturated solutions contain more solute than the usual maximum amount and are unstable.

They cannot permanently hold the excess solute in solution and may release it suddenly.

\[
\text{NaCl}_\text{(s)} \leftrightarrow \text{Na}^{+1}\text{(aq)} + \text{Cl}^{-1}\text{(aq)}
\]
Super Saturated Solutions

Show Quick Time
Super Saturated Video Here
How do you make great fudge?

Start with ingredients
How do you make great fudge?

Heat and stir to increase the kinetic energy and dissolve as much sugar into solution as possible.
How do you make great fudge?

Using copper pots helps
How do you make great fudge?

Pour out to cool SLOWLY!
What kind of table is he pouring onto?
How do you make great fudge?

A marble table allows the fudge to cool slowly. It helps to continue to stir as it cools.
How do you make great fudge?

You may also add nuts and goodies as it cools. Then slice and eat.

Bad fudge tastes like granulated sugar because it cools too fast.
Graphs and solubility tables are useful planning tools for making solutions.

Can you draw the curve of a graph as a salt increases in solubility as temperature increases?
Factors affecting solubility

The tendency of a substance to dissolve in another depends on:
Factors affecting solubility

1. Nature of the solute and solvent (or types of solutes and solvents)
   ➢ example: at room temp, 1g of PbCl$_2$/100g H$_2$O and 200g ZnCl$_2$/100g H$_2$O

Which solute is more soluble and why, can you explain?

ZnCl$_2$
Factors affecting solubility

- Intermolecular forces are an important factor in determining solubility of a solute in a solvent.
- The stronger the attraction between solute and solvent molecules, the greater the solubility.
Factors affecting solubility

- For example, polar liquids tend to dissolve in polar solvents.
- Favorable dipole-dipole interactions exist (solute-solute, solvent-solvent, and solute-solvent).
- Water and ethanol are miscible because the broken hydrogen bonds in both pure liquids are re-established in the mixture.
- However, not all alcohols are miscible with water.
• Why? The number of carbon atoms in a chain affects solubility.
• The greater the number of carbons in the chain, the more the molecule behaves like a hydrocarbon.
• Thus, the more C atoms in the alcohol, the lower its solubility in water.

<table>
<thead>
<tr>
<th>Alcohol</th>
<th>Solubility in $H_2O^a$</th>
<th>Solubility in $C_6H_{14}$</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH$_3$OH (methanol)</td>
<td>$\infty$</td>
<td>0.12</td>
</tr>
<tr>
<td>CH$_3$CH$_2$OH (ethanol)</td>
<td>$\infty$</td>
<td>$\infty$</td>
</tr>
<tr>
<td>CH$_3$CH$_2$CH$_2$OH (propanol)</td>
<td>$\infty$</td>
<td>$\infty$</td>
</tr>
<tr>
<td>CH$_3$CH$_2$CH$_2$CH$_2$OH (butanol)</td>
<td>0.11</td>
<td>$\infty$</td>
</tr>
<tr>
<td>CH$_3$CH$_2$CH$_2$CH$_2$CH$_2$OH (pentanol)</td>
<td>0.030</td>
<td>$\infty$</td>
</tr>
<tr>
<td>CH$_3$CH$_2$CH$_2$CH$_2$CH$_2$CH$_2$OH (hexanol)</td>
<td>0.0058</td>
<td>$\infty$</td>
</tr>
<tr>
<td>CH$_3$CH$_2$CH$_2$CH$_2$CH$_2$CH$_2$CH$_2$OH (heptanol)</td>
<td>0.0008</td>
<td>$\infty$</td>
</tr>
</tbody>
</table>
• Increasing the number of –OH groups within a molecule increases its solubility in water.
• The greater the number of –OH groups along the chain, the more solute-water H-bonding is possible.

Vitamin A
Vitamin C
• Generalization: “Like dissolves like”. Substances with similar intermolecular attractive forces tend to be soluble in one another.
• The more polar bonds in the molecule, the better it dissolves in a polar solvent.
• The less polar the molecule the less likely it is to dissolve in a polar solvent and the more likely it is to dissolve in a nonpolar solvent.
• Network solids do not dissolve because the strong intermolecular forces in the solid are not reestablished in any solution.
2. Pressure

The solubility of a gas in a liquid is a function of the pressure of the gas over the solution.

- a. solid in liquid?
  - No Effect

- b. gas in liquid?

- What happen to when the pressure goes up?
  - Increase Pressure, Increase solubility of the gas
2. Pressure

- c. Henry’s Law

At a given temperature the solubility (S) of a gas in a liquid is directly proportional to the pressure (P) of the gas above the liquid.
2. Pressure

• c. Henry’s Law

- At a given temperature the solubility ($S$) of a gas in a liquid is directly proportional to the pressure ($P$) of the gas above the liquid.
2. Pressure

- Henry’s Law

\[
\frac{S_1}{P_1} = \frac{S_2}{P_2}
\]

\[S = \text{Solubility}\]

\[P = \text{Pressure}\]
3. Temperature ___ temp, ___ kinetic energy, ___ randomness
   a. solid in liquid?

➢ what if cools back down?

Dissolved solid falls back out of solution
b. gas in liquid?

What happen to a gas as the liquid warms up?

Increase temp, decrease solubility
Heat of Solution

• Equation:

\[ \Delta H_{\text{soln}} = \Delta H_1 + \Delta H_2 + \Delta H_3 \]

H = enthalpy = heat content of a solution
H units expressed in kJ/mole (kJ=kilojoules)
If \( \Delta H_{\text{soln}} \) is: negative=exothermic
positive=endothermic
Heat of Solution

• Three steps involving heat of solution
  – Separation of Solute Molecules $= \Delta H_1$ (endo)
  – Separation of Solvent Molecules $= \Delta H_2$ (endo)
  – Formation of solute/solvent interactions $= \Delta H_3$ (exo)
Heat of Solution:

Energy Changes and Solution Formation
• There are three steps involving energy in the formation of a solution:
• Separation of solute molecules ($\Delta H_1$),

$\Delta H_1$: Separation of solute molecules
Heat of Solution:

Energy Changes and Solution Formation

• Separation of solvent molecules ($\Delta H_2$), and

$\Delta H_2$: Separation of solvent molecules
Heat of Solution:

Energy Changes and Solution Formation
• Formation of solute-solvent interactions ($\Delta H_3$).
Heat of Solution:

\( \Delta H \) soln can either be positive or negative depending on the intermolecular forces.

• To determine whether \( \Delta H \) soln is positive or negative, we consider the strengths of all solute-solute, solvent-solvent and solute-solvent interactions:
Heat of Solution:

• Breaking attractive intermolecular forces is always endothermic.
• \( \Delta H_1 \) and \( \Delta H_2 \) are both positive.
• Forming attractive intermolecular forces is always exothermic.
Heat of Solution:

- Forming attractive intermolecular forces is always exothermic.
- $\Delta H_3$ is always negative.
Heat of Solution:

It is possible to have either \( \Delta H_3 > (\Delta H_1 + \Delta H_2) \) or \( \Delta H_3 < (\Delta H_1 + \Delta H_2) \).

Examples:

- \( \text{MgSO}_4 \) added to water has \( \Delta H_{\text{soln}} = -91.2 \text{ kJ/mol} \).
- \( \text{NH}_4\text{NO}_3 \) added to water has \( \Delta H_{\text{soln}} = +26.4 \text{ kJ/mol} \).

\( \text{MgSO}_4 \) is often used in instant heat packs and \( \text{NH}_4\text{NO}_3 \) is often used in instant cold packs.
Heat of Solution:

How can we predict if a solution will form?

• In general, solutions form if the $\Delta H_{\text{soln}}$ is negative.
• If $\Delta H_{\text{soln}}$ is too endothermic a solution will not form.
• "Rule of thumb": Polar solvents dissolve polar solutes.
• Nonpolar solvents dissolve nonpolar solutes.
• Consider the process of mixing NaCl in gasoline.
• Only weak interactions are possible because gasoline is nonpolar.
• These interactions do not compensate for the separation of ions from one another.
Heat of Solution:

• Result: NaCl doesn't dissolve to any great extent in gasoline.

• Consider the process of mixing water in octane ($C_8H_{18}$).
  • Water has strong H-bonds.
  • The energy required to break these H-bonds is not compensated for by interactions between water and octane.

• Result: Water and octane do not mix.
Heat of Solution:
- **Heat of Solution**: overall energy change that occurs during the solution process

- Describe the solutions you made:
  - Ammonium chloride was endothermic
  - Calcium chloride was exothermic
B. Factors affecting rate of solution

• 1. size of particle
  a. dissolving only occurs at surface
  b. smaller pieces, ___ surface area, ___ rate of solution
B. Factors affecting rate of solution

2. stirring
   a. brings fresh parts of solvent in contact with solute, so ___ rate of solution
B. Factors affecting rate of solution

3. amount of solute already dissolved
   a. the more solute dissolved, ______
      rate of solution
B. Factors affecting rate of solution

4. temperature
a. solids: ____ temp, ____ rate of solution
b. gases: ____ temp, ____ rate of solution
IV. Solution Concentrations

What does concentration mean? (15.2)

<table>
<thead>
<tr>
<th>Concentration description</th>
<th>Ratio</th>
</tr>
</thead>
<tbody>
<tr>
<td>Percent by mass</td>
<td>$\frac{\text{mass of solute}}{\text{mass of solution}} \times 100$</td>
</tr>
<tr>
<td>Percent by volume</td>
<td>$\frac{\text{volume of solute}}{\text{volume of solution}} \times 100$</td>
</tr>
<tr>
<td>Molarity</td>
<td>$\frac{\text{moles of solute}}{\text{liter of solution}}$</td>
</tr>
<tr>
<td>Molality</td>
<td>$\frac{\text{moles of solute}}{\text{kilogram of solvent}}$</td>
</tr>
<tr>
<td>Mole fraction</td>
<td>$\frac{\text{moles of solute}}{\text{moles of solute} + \text{moles of solvent}}$</td>
</tr>
</tbody>
</table>
IV. Solution Concentrations

What does concentration mean? (15.2)

What is concentrated or dilute?

**Concentrated** - how much of certain substance the solution contains

Frozen orange juice? 3 parts water to 1 part orange juice

Frozen grape juice?

4 parts water to 1 part grape juice

Same as 25 parts per 100 or 25%
IV. Solution Concentrations

What does concentration mean? (15.2)

What is concentrated or dilute?

Dilute? – phosphoric acid in soft drinks

Concentrated phosphoric acid is toxic

How dilute is dilute? Serial dilution
Serial Dilution

1 stock/ml

Concentration

1/10

Actual Stock Dilution

.1 stock/ml

1/10 (1/100)

.01 stock/ml

.01/10 (1/1000)

.001 stock/ml

.001/10 (1/10000)

ppt

Part per thousand
Serial Dilution

ppt
Part per thousand

ppm
Part per million

.00001
0.0001/10
1/100,000

.000001
0.000001/10
1/1,000,000
If ppm is so small, who cares?

0.7 – 1.0 ppm NaF in water prevents tooth decay

> 1.0 ppm cause mottling of tooth enamel
If ppm is so small, who cares?

5-20 ppm of Iodine prevents formation of goiter

Goiter is enlargement of thyroid gland
If ppm is so small, who cares?

0.5 ppm chlorine in swimming pools kills bacteria

2-3 ppm chlorinates swimming pools
A. Molarity - the number of moles of solute in 1 liter of solution

\[ M \text{ (molarity)} = \frac{\# \text{ of moles of solute}}{1 \text{ Liter of solution}} \]
example: What is the molarity of a solution which contains 4.9 g of H$_2$SO$_4$ in 250 ml of solution?

Find moles first?

\[
4.9 \text{ g } \text{H}_2\text{SO}_4 \times \frac{1 \text{ mole}}{98.1 \text{ g}} = 0.050 \text{ moles}
\]

Find liters next?

\[
250 \text{ ml} \times \frac{1 \text{ liter}}{1000 \text{ ml}} = 0.250 \text{ liters}
\]

Now find molarity?

\[
\frac{0.050 \text{ moles}}{0.250 \text{ liters}} = 0.20 \text{ M}
\]
example #2: How many moles of KNO$_3$ are contained in 300 ml of a 0.50 M solution?

Write molarity as moles over liters?

\[
\frac{0.50 \text{ moles}}{L} \times 0.300 \text{ liters} = 0.15 \text{ moles}
\]

Multiply by liters given

Units cancel and your left with moles
B. Molality - the number of moles of solute per kilogram (1000 grams) of solvent

\[ m \text{ (molality)} = \frac{\text{# of moles of solute}}{1 \text{ kg of solvent}} \]
Molality

Calculate the molality (m) of a solution made by adding 4.5 g NaCl to 100.0 g water

Convert g of NaCl to moles:

\[
4.5 \text{ g NaCl} \times \frac{1 \text{ mole NaCl}}{58.44 \text{ g NaCl}} = 0.077 \text{ moles NaCl}
\]
Molality

Convert g of water to kg:

$100 \text{ g } H_2O \times \frac{1 \text{ kg } H_2O}{1000 \text{ g } H_2O} = 0.100 \text{ kg water}$

Just move decimal point 3 places to the left!
Molality

\[ m = \frac{\text{moles solute}}{\text{kg solvent}} \]

.077 moles NaCl = .77 m or .77 molal

.100 kg

We use molality (m) because its value does NOT change with temperature
Now finish your (50 question) solution walkaround and (22 question) molarity walkaround.
Fewer solvent molecules leave surface, lowering vapor pressure
VI. Freezing point depression and boiling point elevation

1. VP: the pressure exerted in a closed container by liquid particles that have escaped the liquid’s surface and have entered the gaseous state.
VII. Freezing point depression and boiling point elevation

4. Remember Boiling Point is defined as:

When vapor pressure = atm pressure
VI. Freezing point depression and boiling point elevation

B. Boiling point elevation and freezing point depressions depend on:
1. molal concentration - number of moles of solute in a kilogram of solvent
2. nature of the solvent
Example: #2 What would lower the freezing point better on the roads, sugar (C$_{12}$H$_{22}$O$_{11}$), table salt (NaCl), another salt CaCl$_2$, and why?

\[
\begin{align*}
\text{C}_{12}\text{H}_{22}\text{O}_{11} & \rightarrow \text{C}_{12}\text{H}_{22}\text{O}_{11} \quad \text{(1 mole)} \\
\text{NaCl} & \rightarrow \text{Na}^{+1} + \text{Cl}^{-1} \quad \text{(2 moles)} \\
\text{CaCl}_2 & \rightarrow \text{Ca}^{+2} + 2\text{Cl}^{-1} \quad \text{(3 moles)}
\end{align*}
\]
C. Molal freezing and boiling point constants

1. A 1 molal solution will lower the freezing point of water 1.86 °C. Water would freeze at -1.86 °C instead of 0 °C. This value is called the molal freezing point constant for water. (1.86 C/molal)

relationship: \[ \Delta T_f = k_f m \]
relationship: \[ \Delta T_f = k_f m \]

\( \Delta T_f \) = change in freezing point

\( k_f \) = constant for water

\( m \) = molal concentration of solution
2. A 1 molal solution will raise the boiling point of water 0.52 °C.

\[ \Delta T_b = K_b m \]
Examples: If 85.0 grams of sugar are dissolve in 392 grams of water, what will be the boiling point and freezing points of the resulting solution? The molecular formula of sugar is $C_{12}H_{22}O_{11}$. 

(Answer = 100.33 °C, -1.18 °C)
Fewer solvent molecules leave surface, lowering vapor pressure.

- Pure solvent
- Solvent with nonvolatile solute
Lowering the Vapor Pressure

• Nonvolatile solutes (with no measurable vapor pressure) reduce the ability of the surface solvent molecules to escape the liquid.
• Therefore, vapor pressure is lowered.
• The amount of vapor pressure lowering depends on the amount of solute.
<table>
<thead>
<tr>
<th>Solvent</th>
<th>Normal Boiling Point (°C)</th>
<th>$K_b$ (°C/m)</th>
<th>Normal Freezing Point (°C)</th>
<th>$K_f$ (°C/m)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water, $\text{H}_2\text{O}$</td>
<td>100.0</td>
<td>0.51</td>
<td>0.0</td>
<td>1.86</td>
</tr>
<tr>
<td>Benzene, $\text{C}_6\text{H}_6$</td>
<td>80.1</td>
<td>2.53</td>
<td>5.5</td>
<td>5.12</td>
</tr>
<tr>
<td>Ethanol, $\text{C}_2\text{H}_5\text{OH}$</td>
<td>78.4</td>
<td>1.22</td>
<td>$-114.6$</td>
<td>1.99</td>
</tr>
<tr>
<td>Carbon tetrachloride, $\text{CCl}_4$</td>
<td>76.8</td>
<td>5.02</td>
<td>$-22.3$</td>
<td>29.8</td>
</tr>
<tr>
<td>Chloroform, $\text{CHCl}_3$</td>
<td>61.2</td>
<td>3.63</td>
<td>$-63.5$</td>
<td>4.68</td>
</tr>
</tbody>
</table>
VII. Heterogeneous Mixtures

1. **Suspension** (NOT A SOLUTION) – a mixture containing particles that remain while thoroughly mixed, but then settles out when left undisturbed.

Can you think of examples of a suspension?
Can you think of examples of a suspension?
Chalk + Water = Suspension
2. **Colloids** (NOT A SOLUTION) – **Colloids** or colloidal dispersions are suspensions in which the suspended particles are larger than molecules but too small to separate out of the suspension due to gravity. Particle size: 10 to 2000 Å.

Can you think of examples colloids?
<table>
<thead>
<tr>
<th>Phase of Colloid</th>
<th>Dispersing (solventlike) Substance</th>
<th>Dispersed (solutelike) Substance</th>
<th>Colloid Type</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Gas</td>
<td>Gas</td>
<td>Gas</td>
<td>—</td>
<td>None (all are solutions)</td>
</tr>
<tr>
<td>Gas</td>
<td>Gas</td>
<td>Liquid</td>
<td>Aerosol</td>
<td>Fog</td>
</tr>
<tr>
<td>Gas</td>
<td>Gas</td>
<td>Solid</td>
<td>Aerosol</td>
<td>Smoke</td>
</tr>
<tr>
<td>Liquid</td>
<td>Liquid</td>
<td>Gas</td>
<td>Foam</td>
<td>Whipped cream</td>
</tr>
<tr>
<td>Liquid</td>
<td>Liquid</td>
<td>Liquid</td>
<td>Emulsion</td>
<td>Milk</td>
</tr>
<tr>
<td>Liquid</td>
<td>Liquid</td>
<td>Solid</td>
<td>Sol</td>
<td>Paint</td>
</tr>
<tr>
<td>Solid</td>
<td>Solid</td>
<td>Gas</td>
<td>Solid foam</td>
<td>Marshmallow</td>
</tr>
<tr>
<td>Solid</td>
<td>Solid</td>
<td>Liquid</td>
<td>Solid emulsion</td>
<td>Butter</td>
</tr>
<tr>
<td>Solid</td>
<td>Solid</td>
<td>Solid</td>
<td>Solid sol</td>
<td>Ruby glass</td>
</tr>
</tbody>
</table>
Aerosol: gas + liquid or solid (e.g., fog and smoke),
• Foam: liquid + gas (e.g., whipped cream),
• Emulsion: liquid + liquid (e.g., milk),
• Sol: liquid + solid (e.g., paint),
• Solid foam: solid + gas (e.g., marshmallow),
• Solid emulsion: solid + liquid (e.g., butter),
• Solid sol: solid + solid (e.g., ruby glass).
Colloidal: smaller than suspension, larger than solution
3. **Tyndall Effect** Dispersed colloid particles that are large enough to scatter light. The path of a beam of light projected through a colloidal suspension can be seen through the suspension.

📖 Watch and observe the demo?
Hydrophilic and Hydrophobic Colloids

Water-loving colloids: hydrophilic.
Water-hating colloids: hydrophobic.
In the human body, large biological molecules such as proteins are kept in suspension by association with surrounding water molecules.

- These macromolecules fold up so that hydrophobic groups are away from the water (inside the folded molecule).
- Hydrophilic groups are on the surface of these molecules and interact with solvent (water) molecules.
Typical hydrophilic groups are polar (containing C–O, O–H, N–H bonds) or charged.
• Consider a small drop of oil in water. They will not mix unless soaps act on the oil and water.
• Soaps are molecules with long hydrophobic tails and hydrophilic heads that remove dirt by stabilizing the colloid in water.
• Most dirt stains on people and clothing are oil-based.
Triglyceride

A molecule of fat or oil. It consists of 3 free-swinging fatty acid molecules hooked to a glycerol backbone.
• Biological application of this principle:
• The gallbladder excretes a fluid called bile.
• Bile contains substances (bile salts) that form an emulsion with fats in our small intestine.
Oil and Vinegar Salad Dressing?
Do they mix?
You can form an emulsion with lecithin.
Lecithin acts as a emulsifying agent to keep oil and vinegar mixed