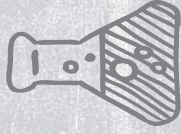
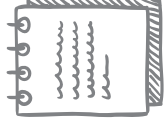


SOLUTIONS



Chapter 9



SOLUTIONS

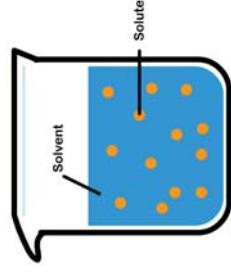
Chapter 9

- 9.1 – Solutions
- 9.2 – Electrolytes and Nonelectrolytes
- 9.3 – Solubility
- 9.4 – Solution Concentrations and Reactions
- 9.5 – Dilution of Solutions
- 9.6 – Properties of Solutions

9.1 - Solutions

Identify the solute and solvent in a solution.
Describe the formation of a solution.

Solution Definitions



A **solution** is a homogeneous mixture in which one substance, called the **solute**, is uniformly dispersed (dissolved) in another substance, called the **solvent**.

Review: a **homogenous mixture** is a combination of 2+ substances that are uniform throughout. Visually, it appears as composed of one substance. Example: milk



Solutions

The solute and solvent do not react with each other and so they can be mixed in varying proportions.

Salt water is a solution

- Solvent = water (present in larger amount)
- Solute = salt (smaller amount)

Different amount of salt can be dissolved in water.

A solution has at least one solute dispersed in a solvent, and can have more.



Types of Solutes and Solvents

- Solutes and solvents may be solids, liquids, or gases.
- The solution that forms has the same physical state as the solvent.
- Example: when sugar crystals dissolve in water, the resulting solution is a liquid.
 - Sugar – solute (s)
 - Water – solvent (l)
- Example: Soft drinks are carbonated by dissolving CO₂ in water and is a liquid.
 - CO₂ – solute (g)
 - Water – solvent (l)



TABLE 9.1 Some Examples of Solutions

Type	Example	Primary Solute	Solvent
Gas Solutions			
Gas in a gas	Air	O ₂ (g)	N ₂ (g)
Liquid Solutions			
Gas in a liquid	Soda water	CO ₂ (g)	H ₂ O(l)
	Household ammonia	NH ₃ (g)	H ₂ O(l)
Liquid in a liquid	Vinegar	HC ₂ H ₃ O ₂ (l)	H ₂ O(l)
Solid in a liquid	Seawater	NaCl(s)	H ₂ O(l)
	Tincture of iodine	I ₂ (s)	C ₂ H ₅ OH(l)
Solid Solutions			
Solid in a solid	Brass	Zn(s)	Cu(s)
	Steel	C(s)	Fe(s)



Water as a Solvent

Water is one of the most common solvents in nature, thanks to its structure and properties:

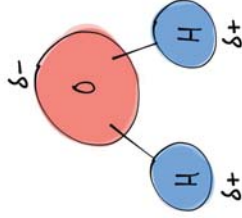
- Water is polar.
- Water can hydrogen bond.



Water is Polar (Review Section 6.8)

- The shared electrons in each of the O-H bonds in water are held tighter to Oxygen than to Hydrogen.
- This is because Oxygen is more electronegative than Hydrogen.
- Thus, the O-H bonds are *polar*.

- Because the shape of a water molecule is bent, its dipole don't cancel out.
- Thus the entire molecule is *polar* and water is said to be a **polar solvent**.



Water Molecules can Hydrogen Bond (Review Section 6.9)

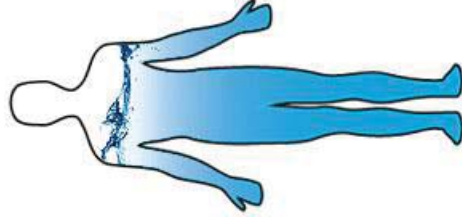
- **Hydrogen bonds:** the attraction between a partially positive H atom and a strongly electronegative atom of F, O, or N.
- Water molecules are linked together via hydrogen bonds.



- Hydrogen bonds are important in the properties of biological compounds such as proteins, carbohydrates, and DNA.

FYI: Water in the Body

- The average adult is about 60% water by mass and the average infant is about 75% water.
- About 60% of the body's water is contained within the cells. The other 40% makes up extracellular fluids, which include the interstitial fluid in tissue and the plasma in the blood.



FYI: Water in the Body

- Everyday you lose between 1500-3000 mL of water.
- Serious dehydration can occur in an adult if there is a 10% net loss in total body fluid; a 20% loss of fluid can be fatal.
- An infant suffers severe dehydration with only 5-10% loss in body fluid.
- Water loss is continually replaced by the liquids and foods in the diet and from metabolic processes that produce water in the cells in the body.

Water Loss	
Urine	1500 mL
Perspiration	300 mL
Breath	600 mL
Feces	100 mL
Total	2500 mL

Water Gain	
Liquid	1000 mL
Food	1200 mL
Metabolism	300 mL
Total	2500 mL

Forming Solutions

⚠️ Not all combinations of solute and solvent result in a solution.

- Steps to forming a solution:
 - 1. Break solute-solute and solvent-solvent attractions (*energy required*)
 - 2. Form solute-solvent attractions (*energy released*)



Oil and water have different polarities, so they will not form a solution.

Forming Solutions

Steps to forming a solution:

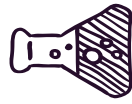
- 1. Break solute-solute and solvent-solvent attractions (*energy required*)
- 2. Form solute-solvent attractions (*energy released*)

In order to provide the energy for Step 1 to occur, the solute and solvent particles must be attracted to each other.

This occurs when the solute and solvent have similar *polarities*.

Forming Solutions

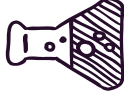
LIKE dissolves LIKE



Oil
Grease
Benzene
Hexane



Water
DMSO
Ethanol
Acetone



Polar solutes dissolve in Polar solvents.
Nonpolar solutes dissolve in Nonpolar solvents.



LIKE dissolves LIKE

Top layer: water (polar)
Bottom layer: methylene chloride, CH_2Cl_2 (nonpolar)

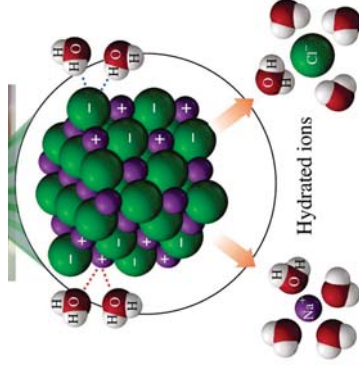
Solutions with Ionic and Polar Solutes

Ionic compounds will dissolve in polar solvents.

When NaCl crystals are placed in water, partially negative oxygen atoms in the water molecules attract the positive Na^+ ions. The partially positive hydrogen atoms in other water molecules attract negative Cl^- ions. Luring individual ions away from the crystal.

As soon as a water molecule lures a Na^+ or Cl^- off on its own, other water molecules will surround it, **hydrating** it.

Hydration of the ions diminishes their attraction to other ions and keeps them in solution.



Solutions with Nonpolar Solutes

- Compounds containing nonpolar molecules, such as iodine (I_2), oil, or grease, do not dissolve in water because there are essentially no attractions between the particles of a nonpolar solute and the polar solvent.

- Nonpolar solutes require nonpolar solvents for a solution to form.

Like dissolves Like

9.1 – Solutions

9.2 – Electrolytes and Nonelectrolytes

9.3 – Solubility

9.4 – Solution Concentrations and Reactions

9.5 – Dilution of Solutions

9.6 – Properties of Solutions



SOLUTIONS

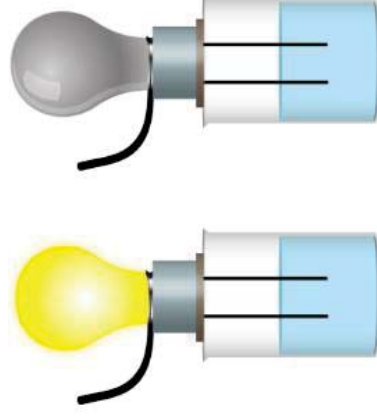
Chapter 9

9.2 – Electrolytes and Nonelectrolytes

Identify solutes as electrolytes or nonelectrolytes.

Electrolytes

- Solutes can be classified by their ability to conduct an electrical current.
 - Some compounds *dissociate* into ions when put into solution:
- $$AB(s, l, \text{ or } g) \xrightarrow{\text{solvent}} A^+(aq) + B^-(aq)$$
- Ions conduct electricity.
 - **Electrolytes:** compounds that dissociate into ions and conduct electricity.
 - **Nonelectrolytes:** compounds that *do not* dissociate into ions and therefore *do not* conduct electricity.

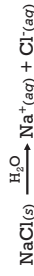


A conductivity apparatus is an incomplete electrical circuit that contains a source of electricity and a light bulb or meter that will show when current is flowing through the circuit. The ends of the incomplete circuit are prongs that can be lowered into a solution. If the prongs are lowered into a solution containing a sufficient number of ions, the circuit will be completed by the solution, current will flow, and the light bulb will light up. If the prongs are lowered into a solution with no ions or an insufficient number of ions, not enough current will flow to light the bulb.

Types of Electrolytes

Electrolytes can be further classified as **strong** electrolytes or **weak** electrolytes.

Strong electrolytes form from compounds dissociating completely into positive and negative ions when added to solution:



Strong electrolytes conduct electricity very well. (The lightbulb would have a bright glow.)

Compounds that do this include: all soluble ionic compounds, certain molecular compounds (strong acids and bases. More to come in chapter 11.)

Types of Electrolytes

Electrolytes can be further classified as **strong** electrolytes or **weak** electrolytes.

Weak electrolytes form from compounds that only partially dissociate into positive and negative ions when added to solution:



Weak electrolytes conduct electricity, but not as strong as strong electrolytes. This is due to few ions (as most of the HF is in the form of the molecule and not the ions.)

The lightbulb will be dim.

Compounds that will do this are certain molecular compounds (weak acids and bases – to come in chapter 11).

Types of Electrolytes

Most molecular compounds *do not* dissociate into ions but remain 100% molecules. These do not conduct electricity due to lack of ions and are therefore **nonelectrolytes**.



The lightbulb won't light up.

Most molecular compounds are nonelectrolytes.



TABLE 9.4 Classification of Solutes in Aqueous Solutions

Type of Solute	In Solution	Type(s) of Particles in Solution	Conducts Electricity?	Examples
Strong electrolyte	Dissociates completely	Ions only	Yes	Ionic compounds such as NaCl, KBr, MgCl ₂ , NaNO ₃ ; bases such as NaOH, KOH; acids such as HCl, HBr, HI, HNO ₃ , HClO ₄ , H ₂ SO ₄
Weak electrolyte	Ionizes partially	Mostly molecules and a few ions	Weakly	HF, H ₂ O, NH ₃ , HC ₂ H ₃ O ₂ (acetic acid)
Nonelectrolyte	No ionization	Molecules only	No	Carbon compounds such as CH ₃ OH (methanol), C ₂ H ₅ OH (ethanol), C ₁₂ H ₂₂ O ₁₁ (sucrose), CH ₃ N ₂ O (urea)

Practice

Identify the following as strong electrolyte, weak electrolyte, or nonelectrolyte:



Practice

Indicate whether the aqueous solutions of each of the following solutes contain only ions, only molecules, or mostly molecule with few ions:

Acetic acid, HC₂H₃O₂, weak electrolyte

NaBr, a strong electrolyte

Fructose, C₆H₁₂O₆, a nonelectrolyte

Practice

Write a balanced equation for the dissociation of the strong electrolytes in water:



$$1000 \text{ mEq} = 1 \text{ Eq}$$

The concentration of electrolytes in IV's are expressed in *equivalents per liter (Eq/L)* or more commonly *milliequivalents per liter (mEq/L)*.

In any solution, the charges of the positive ions is always balanced by the charge of the negative ions.

For example, a solution containing 25 mEq/L of Na^+ and 4 mEq/L of K^+ has a total positive charge of 29 mEq/L.

If Cl^- is the only negative ion, its concentration must be 29 mEq/L.

Equivalents (medical field)

Body fluids and intravenous (IV) solutions contain a mixture of electrolytes, such as Na^+ , Cl^- , K^+ , and Ca^{2+}

Each ion is measured in terms of an *equivalent (Eq)*

Equivalent (Eq): the amount of ion equal to 1 mole of positive or negative electrical charge.

TABLE 9.5 Equivalents of Electrolytes in Clinical Intravenous (IV) Solutions

Ion	Ionic Charge	Number of Equivalents in 1 Mole
Na^+ , K^+ , Li^+ , NH_4^+	1+	1 Eq
Ca^{2+} , Mg^{2+}	2+	2 Eq
Fe^{3+}	3+	3 Eq
Cl^- , $\text{C}_2\text{H}_3\text{O}_2^-$ (acetate), H_2PO_4^- , $\text{C}_3\text{H}_5\text{O}_3^-$ (lactate)	1-	1 Eq
CO_3^{2-} , HPO_4^{2-}	2-	2 Eq
PO_4^{3-} , $\text{C}_6\text{H}_5\text{O}_7^{3-}$ (citrate)	3-	3 Eq

Equivalents

The concentration of electrolytes in IV's are expressed in *equivalents per liter (Eq/L)* or more commonly *milliequivalents per liter (mEq/L)*.

$$1000 \text{ mEq} = 1 \text{ Eq}$$

In any solution, the charges of the positive ions is always balanced by the charge of the negative ions.

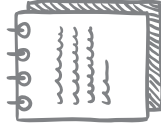
For example, a solution containing 25 mEq/L of Na^+ and 4 mEq/L of K^+ has a total positive charge of 29 mEq/L.

If Cl^- is the only negative ion, its concentration must be 29 mEq/L.

Practice

The lab tests for a patient indicate a blood calcium level of 8.8 mEq/L.

- How many moles of calcium ions are in 0.80 L of blood?
- If chloride ion is the only other ion present, what is its concentration in mEq/L?



- ✓ 9.1 – Solutions
- ✓ 9.2 – Electrolytes and Nonelectrolytes
- 9.3 – Solubility
- 9.4 – Solution Concentrations and Reactions
- 9.5 – Dilution of Solutions
- 9.6 – Properties of Solutions

SOLUTIONS

Chapter 9

9.3 - Solubility

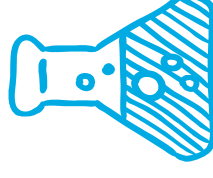
Define solubility; distinguish between an unsaturated and a saturated solution. Identify an ionic compound as soluble or insoluble.

Solubility

- The term *solubility* is used to describe the amount of solute that can dissolve in a given amount of solvent.
- Many factors can affect the solubility of a solute, such as:
 - the type of solute
 - the type of solvent
 - the temperature
- **Solubility:** the maximum amount of solute that can be dissolved at a certain temperature.
 - Usually expressed as: $\frac{\text{grams solute}}{100 \text{ grams solvent}}$

solubility

- If a solute easily dissolved when added to a solvent, the solution does not contain the maximum amount of solute yet.
 - This is an **unsaturated solution**.
- A solution that contains all the solute it can dissolve is called **saturated**.
- In special cases, a solution can be manipulated to dissolve more solute than it should be able to. This is a **supersaturated solution**.



Saturated Solutions

- In a saturated solution, solute will continue to dissolve. But for every particle that dissolves, another that is already dissolved will recrystallize (turn back into its original state, the opposite of dissolve).
- So there is no overall change in the amount of solute dissolved in solution.



For example: adding sugar to water. If you keep adding sugar, at some point, no more will dissolve. It remains a solid. You've saturated the solution.

Practice

At 20°C, the solubility of KCl is 34g KCl / 100g H₂O. In the lab, a student mixes 75g of KCl with 200g H₂O at a temperature at 20°C.

- How much of the KCl will dissolve?
- Is the solution saturated or unsaturated?
- How many grams of solid KCl will be left undissolved at the bottom of the container?

FYI: Gout and Kidney Stones

The conditions of gout and kidney stones involve compounds in the body that exceed their solubility levels and form solid products.

- **Gout** (men, 40+) – uric acid in the blood exceeds its solubility
 - 7mg uric acid per 100 mL of plasma at body temperature
 - Solid uric acid in the form of needle-like crystals for in cartilage, tendons, soft tissues, kidneys...
- **Common causes:** Kidney failure, diet high in uric acid precursors (meat, sardines, mushrooms, asparagus, and beans), alcohol
- **Treatment:** diet changes, medications that break down uric acid and/or prevent uric acid production



FYI: Gout and Kidney Stones

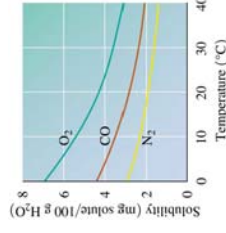
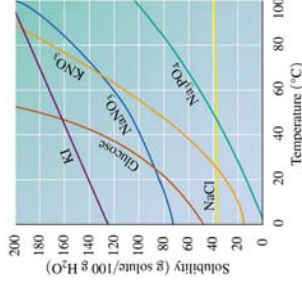
Kidney stones – solid crystals form in the urinary tract of calcium phosphate, calcium oxalate, and/or sometimes uric acid that have exceeded their solubilities.

- causes severe pain
- **Treatment:** wait it out, ultrasound, in extreme cases surgery



Effects of Temperature on Solubility

- For most solids, solubility **increases as temperature increases**. Which means more solute will dissolve at higher temperatures.
 - Example: making candy
- A few substances show little change in solubility as temperature rises.
- For a few gases, solubility **decreases as temperature increases**.

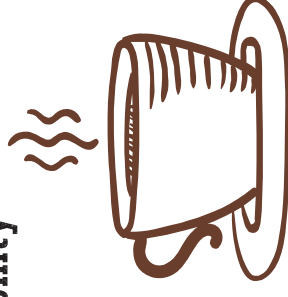


Effects of Temperature on Solubility

If you've ever added sugar to iced tea, then you know that, in general, not all the sugar dissolves. Some of it will accumulate at the bottom of the glass rather than dissolve into the tea.

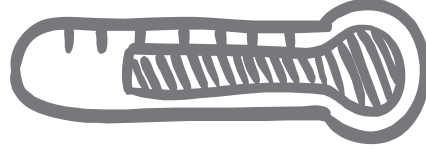
But if you add sugar to hot tea, many teaspoons will dissolve before solid sugar will start to remain.

Hot tea dissolves more sugar than cold tea because the solubility of sugar is much greater at higher temperatures.



Supersaturated Solutions

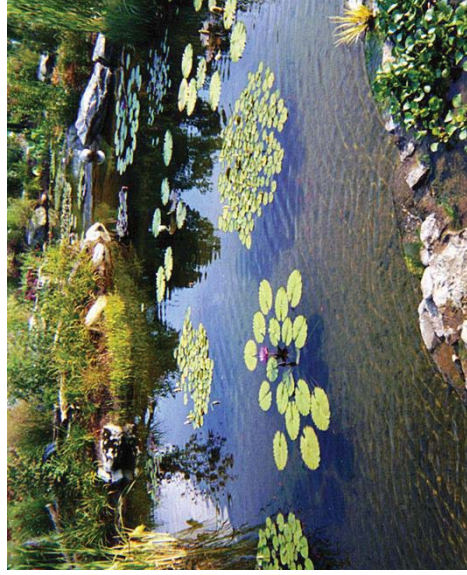
- After adding sugar to the hot tea, if you carefully let it cool, the sugar will remain dissolved.
- Now the iced tea has more sugar than the solubility at colder temperatures allows. It is **supersaturated**.
- Such a solution is unstable, and if the solution is disturbed, the excess solute may recrystallize to the saturated solution level.



Solubility of Gases

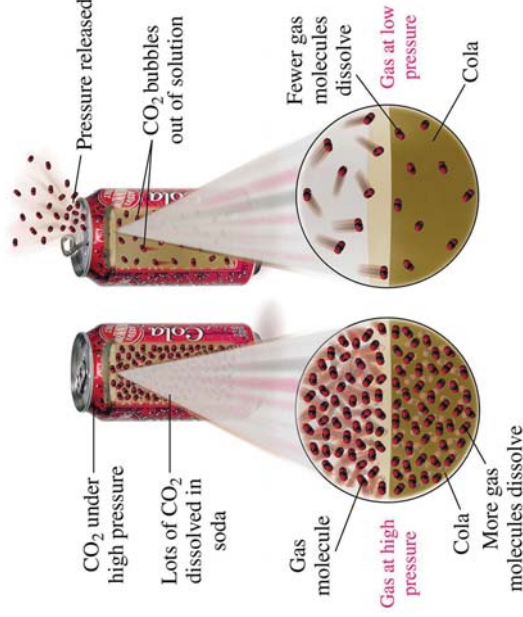
- The solubility of gases **decreases** as temperatures rise. This is because as temperatures rise, gas particles gain more kinetic energy and may escape the solution.
- Perhaps you've noticed more carbonation leaving a soda as it warms.
- At high temperatures, closed soda cans will burst as more gas molecules leave the solution and increase the gas pressure inside.





Solubility of Gases

- Biologists have found that increased temperature in lakes and rivers causes dissolved oxygen to escape.
- If temperatures rise sufficiently, there isn't enough oxygen left to support the biological communities.
- For this reason, power plants must make their own ponds to use with cooling towers so they don't warm the surrounding waterways and kill wildlife.



Henry's Law

Henry's Law: the solubility of a gas in a liquid is directly related to the pressure of the gas above or surrounding the liquid.

As the pressure outside of the liquid increases, the solubility of the gas increases

Soda is carbonated by using CO_2 under high pressure to increase the solubility of the CO_2 in the drink.

When you open the can at atmospheric pressure, the pressure in the can drops. This lowers the solubility of the CO_2 in the drink and dissolved CO_2 turns back into a gas, creating carbonation.

Soluble and Insoluble Ionic Compounds

Up until now, we've assumed all ionic compounds dissolve in water. But that's not entirely true. The **solubility rules** give some guidelines about ionic compounds in water.

TABLE 9.7 Solubility Rules for Ionic Compounds in Water

An ionic compound is soluble in water if it contains one of the following:

Positive ions: Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ , NH_4^+

Negative ions: NO_3^- , $\text{C}_2\text{H}_3\text{O}_2^-$

Cl^- , Br^- , I^- except when combined with Ag^+ , Pb^{2+} , or Hg_2^{2+}

SO_4^{2-} except when combined with Ba^{2+} , Pb^{2+} , Ca^{2+} , Sr^{2+} , or Hg_2^{2+}

Ionic compounds that do not contain at least one of these ions are usually insoluble.

Ionic compounds that are soluble in water typically contain at least one of these ions. Only an ionic compound containing a soluble cation or anion will dissolve in water.

Soluble and Insoluble Ionic Compounds

- Most compounds containing Cl^- are soluble in water.
 - Exceptions: AgCl , PbCl_2 , and Hg_2Cl_2 are insoluble
- Most ionic compounds containing SO_4^{2-} are soluble in water.
 - Exceptions: BaSO_4 , PbSO_4 , CaSO_4 , SrSO_4 , or Hg_2SO_4 are insoluble
- Most other ionic compounds are insoluble.

TABLE 9.7 Solubility Rules for Ionic Compounds in Water

An ionic compound is soluble in water if it contains one of the following:

Positive ions: Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ , NH_4^+

Negative ions: NO_3^- , $\text{C}_2\text{H}_3\text{O}_2^-$

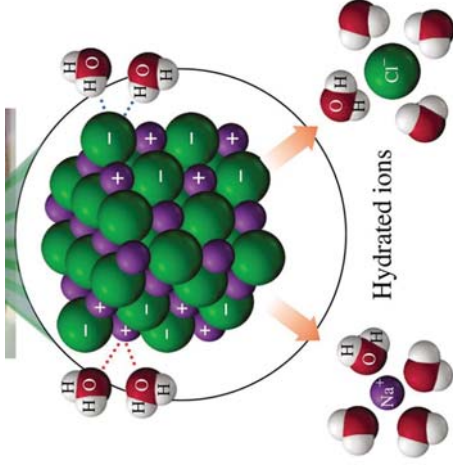
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SO_4^{2-} except when combined with Ba^{2+} , Pb^{2+} , Ca^{2+} , Sr^{2+} , or Hg_2^{2+}

Ionic compounds that do not contain at least one of these ions are usually insoluble.

Soluble and Insoluble Ionic Compounds

In an insoluble ionic compounds, the strength of the ionic bonds are too strong for the polar water molecules to break.



Practice

Use the solubility rules to predict whether the following ionic compounds are soluble or insoluble in water.



TABLE 9.7 Solubility Rules for Ionic Compounds in Water

An ionic compound is soluble in water if it contains one of the following:

Positive ions: Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ , NH_4^+

Negative ions: NO_3^- , $C_2H_3O_2^-$

Cl^- , Br^- , I^- except when combined with Ag^+ , Pb^{2+} , or Hg_2^{2+}

SO_4^{2-} except when combined with Ba^{2+} , Pb^{2+} , Ca^{2+} , Sr^{2+} , or Hg_2^{2+}

Ionic compounds that do not contain at least one of these ions are usually insoluble.



Formation of a Solid



We can use solubility rules to predict whether a solid, (called a precipitate), forms when two solutions containing soluble reactants are mixed...

Practice

When solutions of $NaCl$ and $AgNO_3$ are mixed, a white solid forms. Write the ionic and net ionic equation for the reaction.

Guide to Writing an Equation for the Formation of an Insoluble Ionic Compound

STEP 1 Write the ions of the reactants.

STEP 2 Write the combinations of ions and determine if any are insoluble.

STEP 3 Write the ionic equation including any solid.

STEP 4 Write the net ionic equation.



An ionic compound is soluble in water if it contains one of the following:

Positive ions: Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ , NH_4^+

Negative ions: NO_3^- , $C_2H_3O_2^-$

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SO_4^{2-} except when combined with Ba^{2+} , Pb^{2+} , Ca^{2+} , Sr^{2+} , or Hg_2^{2+}

Ionic compounds that do not contain at least one of these ions are usually insoluble.

Practice

Predict whether a solid would form. If so, write the net ionic equation.



Guide to Writing an Equation for the Formation of an Insoluble Ionic Compound

STEP 1 Write the ions of the reactants.

STEP 2 Write the combinations of ions and determine if any are insoluble.

STEP 3

Write the ionic equation including any solid.

STEP 4

Write the net ionic equation.

An ionic compound is soluble in water if it contains one of the following:

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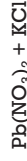
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SO_4^{2-} except when combined with Ba^{2+} , Pb^{2+} , Ca^{2+} , Sr^{2+} , or Hg_2^{2+}

Ionic compounds that do not contain at least one of these ions are usually insoluble.

Practice

Predict whether a solid would form. If so, write the net ionic equation.



Guide to Writing an Equation for the Formation of an Insoluble Ionic Compound

STEP 1 Write the ions of the reactants.

STEP 2 Write the combinations of ions and determine if any are insoluble.

STEP 3

Write the ionic equation including any solid.

STEP 4

Write the net ionic equation.

An ionic compound is soluble in water if it contains one of the following:

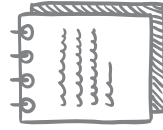
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Ionic compounds that do not contain at least one of these ions are usually insoluble.



9.1 Solutions

9.2 Electrolytes and Nonelectrolytes

9.3 Solubility

9.4 Solution Concentrations and Reactions

9.5 Dilution of Solutions

9.6 Properties of Solutions

9.4 – Solution Concentrations and Reactions



Calculate the concentration of a solute in a solution; use concentration units to calculate the amount of solute or solution. Given the volume and concentration of a solution, calculate the amount of another reactant or product in a reaction.

Concentration

- The amount of solute dissolved in a certain amount of solution is called the **concentration** of the solution.
- This section looks at different ways to express a concentration as a ration of solute and solution.

$$\text{concentration} = \frac{\text{amount of solute}}{\text{amount of solution}}$$

Amount of solute can be g, mL, or moles
Amount of solution can be g, mL, or L

Mass Percent (m/m) Concentration

Mass percent: describes the mass of the solute in grams for *exactly* 100g of solution.

$$\text{mass percent} = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 100$$

- The units of mass for solute and solution *must* be the same. (kg would be fine, as long as they are BOTH in kg.)
- The mass of solution = mass of solute + mass of solvent

Practice

What is the mass percent if you prepare a solution by mixing 8.00g KCl (solute) with 42.00g of water (solvent)?

Practice

What is the mass percent of NaOH in a solution prepared by dissolving 30.0g of NaOH in 120.0g of water?

Volume Percent (v/v) Concentration

Because the volume of liquids are easily measured, the concentrations of solutions are often expressed as volume percent (v/v).

Volume percent: the volume of solute in *exactly* 100mL of solution.

$$\text{volume percent} = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 100$$

- The units *must* be the same for both. Either mL or L usually
- Volume of solution = volume of solute + volume of solution



Volume Percent

The label indicates that vanilla extract contains 35% (v/v) alcohol.

35mL of ethanol in every 100 mL vanilla solution.

Practice

A bottle contains 59 mL of lemon extract solution. If the extract contains 49 mL of alcohol, what is the volume percent (v/v) of the alcohol in the solution?



Mass/Volume Percent (m/v) Concentration

Mass/volume percent (m/v): describes the mass of the solute in grams for exactly 100mL of solution.

solute units: g

solution units: mL

$$\text{mass/volume percent} = \frac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100$$

Mass/Volume Percent (m/v)

- The mass/volume percent is widely used in hospitals and pharmacies to prepare intravenous (IV) solutions and medicines.
- For example, a 5% (m/v) glucose solution contains 5g of glucose in 100mL of solution. The volume of solution represents the combined volumes of the glucose and water.



Practice

A potassium iodide solution may be used in a diet that is low in iodine. A KI solution is prepared by dissolving 5.0g of KI in enough water to give a final volume of 250 mL. What is the mass/volume percent (m/v) of the KI solution.

Molarity (M) Concentration

The most common type of concentration that chemists use is **molarity**: the number of moles of solute in exactly 1L of solution.

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

Practice

What is the molarity of 1.0 moles of NaCl dissolved in enough water to make 2.0L of solution?

Practice

What is the molarity (M) of 60.0g NaOH in 0.250L of solution?

11 **Na**
Sodium
22.990

8 **O**
Oxygen
15.999

1 **H**
Hydrogen
1.008

TABLE 9.9 Summary of Types of Concentration Expressions and Their Units

Concentration Units	Mass Percent (m/m)	Volume Percent (v/v)	Mass/Volume Percent (m/v)	Molarity (M)
Solute	g	mL	g	mole
Solution	g	mL	mL	L

Using Concentration as a Conversion Factor

In the preparation of solution, we often need to calculate the amount of solute or solution. If the concentration is known, it is useful to use as a conversion factor.

TABLE 9.10 Conversion Factors from Concentrations

Percent Concentration	Meaning	Conversion Factors
10% (m/m) KCl solution	10 g of KCl in 100 g of KCl solution	$\frac{10 \text{ g KCl}}{100 \text{ g solution}}$ and $\frac{100 \text{ g solution}}{10 \text{ g KCl}}$
12% (v/v) ethanol solution	12 mL of ethanol in 100 mL of ethanol solution	$\frac{12 \text{ mL ethanol}}{100 \text{ mL solution}}$ and $\frac{100 \text{ mL solution}}{12 \text{ mL ethanol}}$
5% (m/v) glucose solution	5 g of glucose in 100 mL of glucose solution	$\frac{5 \text{ g glucose}}{100 \text{ mL solution}}$ and $\frac{100 \text{ mL solution}}{5 \text{ g glucose}}$
Molarity		
6.0 M HCl solution	6.0 moles of HCl in 1 liter of HCl solution	$\frac{6.0 \text{ moles HCl}}{1 \text{ L solution}}$ and $\frac{1 \text{ L solution}}{6.0 \text{ moles HCl}}$

Practice

A topical antibiotic is 1.0% (m/v) clindamycin. How many grams of clindamycin are in 60 mL of the solution?

Practice

How many liters of a 2.00M NaCl solution are needed to provide 67.3g of NaCl?

11	Na Sodium 22.990
----	-------------------------------

17	Cl Chlorine 35.453
----	---------------------------------

Chemical Reactions in Solutions

When chemical reactions involve aqueous solutions, we use

- the balanced chemical equation
- the molarity
- the volume

go determine the moles or grams of the reactants or products.

Practice

Zinc reacts with HCl to produce hydrogen gas, H₂, and ZnCl₂.



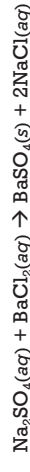
How many liters of a 1.50M HCl solution completely react with 5.32g of zinc?

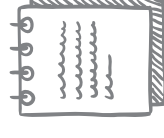
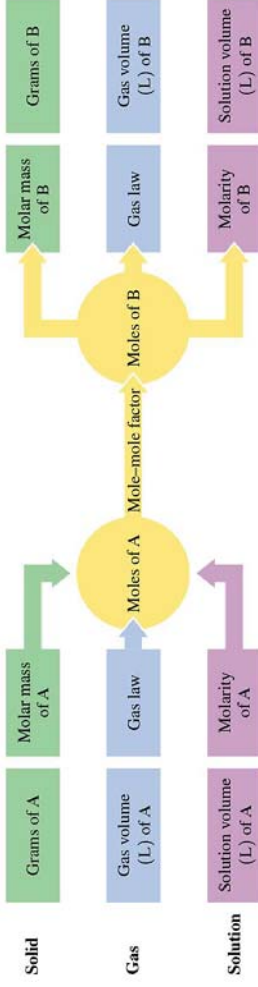
30

Zn
Zinc
65.39

Practice

How many mL of a 0.250M BaCl₂ solution are needed to react with 0.0325L of a 0.160M Na₂SO₄ solution?





- ✓ 9.1 – Solutions
- ✓ 9.2 – Electrolytes and Nonelectrolytes
- ✓ 9.3 – Solubility
- ✓ 9.4 – Solution Concentrations and Reactions
- 9.5 – Dilution of Solutions
- 9.6 – Properties of Solutions



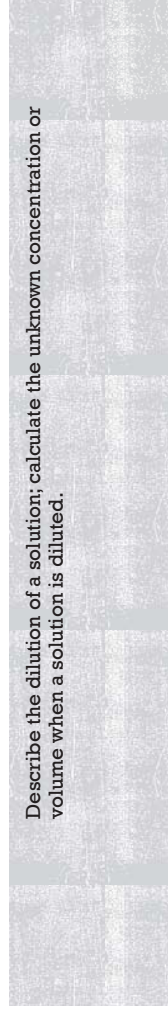
9.5 – Dilution of Solutions

Dilution

In chemistry and biology, we often prepare diluted solutions from more concentrated ones. In a process called **dilution**, a solvent (usually water) is added to a solution, which increases the volume.

As a result, the concentration of the solution **decreases**.

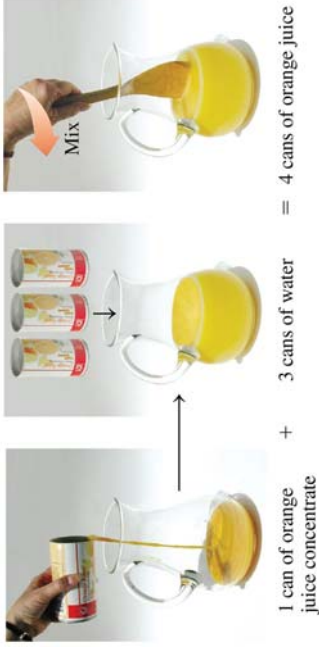
Describe the dilution of a solution; calculate the unknown concentration or volume when a solution is diluted.



Dilution

For example, when you make juice from frozen concentrate, you add 3 cans of water to dilute the orange juice (solute).

The amount of solute doesn't change. It is just spread out over a larger volume.



Dilution

$$C_1 V_1 = C_2 V_2$$

C – concentration (percent concentration or molarity)
V – volume

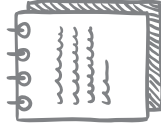
Practice

A doctor orders 1000mL of a 35.0% (m/v) dextrose solution.

If you have a 50.0% (m/v) dextrose solution, how many mL would you use to prepare 1000mL of 35.0% (m/v) dextrose solution?

Practice

What is the molarity of a solution when 75.0 mL of a 4.00M KCl solution is diluted to a volume of 500mL?



- ✔ 9.1 Solutions
- ✔ 9.2 Electrolytes and Nonelectrolytes
- ✔ 9.3 Solubility
- ✔ 9.4 Solution Concentrations and Reactions
- ✔ 9.5 Dilution of Solutions

9.6 – Properties of Solutions

The size and number of solute particles in different types of mixtures play an important role in determining the properties of those mixtures.

SOLUTIONS

Chapter 9

9.6 – PROPERTIES OF SOLUTIONS

Identify a mixture as a solution, a colloid, or a suspension. Describe how the number of particles in a solution affects the freezing point, the boiling point, and the osmotic pressure of a solution.

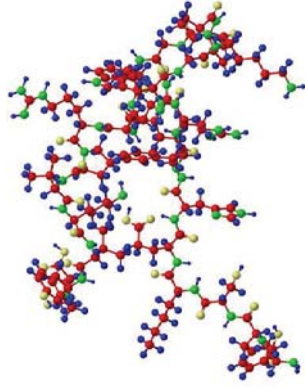
SOLUTIONS

- In **solutions**, the solute is dissolved as small particles that are uniformly dispersed throughout the solvent to give a homogeneous solution.
- When you look at a solution, such as salt water, you cannot visually distinguish the solute from the solvent.
- The solution appears transparent, although it may have a color.
- The particles are so small that they will pass through filters and semipermeable membranes.
 - Semipermeable membranes allow solvent molecules such as water and very small particles to pass through, but not large particles.



COLLOIDS

- The particles in a **colloid** are much larger than solute particles in a solution.
- Colloidal particles are large molecules, such as proteins, or groups of molecules or ions.
- Colloidal particles, similar to solutions, are homogeneous mixtures that do not separate or settle. Out.
- Colloidal particles are small enough to pass through filters, but too large for semipermeable membranes.



COLLOIDS

In colloids the solute isn't necessarily the same state as the solvent.

TABLE 9.11 Examples of Colloids

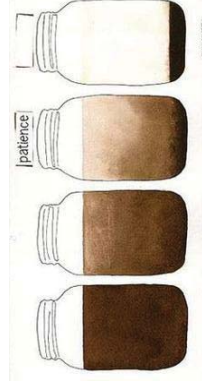
Colloid	Substance Dispersed	Dispersing Medium
Fog, clouds, hair sprays	Liquid	Gas
Dust, smoke	Solid	Gas
Shaving cream, whipped cream, soapsuds	Gas	Liquid
Styrofoam, marshmallows	Gas	Solid
Mayonnaise, homogenized milk	Liquid	Liquid
Cheese, butter	Liquid	Solid
Blood plasma, paints (latex), gelatin	Solid	Liquid

SUSPENSIONS

- **Suspensions:** heterogeneous, nonuniform mixtures that are very different from solutions to colloids.
- The particles of a suspension are so large that they can often be seen with the naked eye.
- They cannot pass through filters nor semipermeable membranes.
- The weight of the suspended solute particles causes them to settle out soon after mixing

SUSPENSION EXAMPLES

- If you stir muddy water, it mixes then quickly separates as the suspended dirt particles settle to the bottom and leave clear water on top.



- Water-treatment plants make use of the properties of suspensions to purify water. When chemicals such as aluminum sulfate or iron (III) sulfate are added to untreated water, they react with impurities to form large suspended particles called *floc*. Filters then trap the particles and allow the clean water to pass through.

TABLE 9.12 Comparison of Solutions, Colloids, and Suspensions

Type of Mixture	Type of Particle	Settling	Separation
Solution	Small particles such as atoms, ions, or small molecules	Particles do not settle	Particles cannot be separated by filters or semipermeable membranes
Colloid	Larger molecules or groups of molecules or ions	Particles do not settle	Particles can be separated by semipermeable membranes but not by filters
Suspension	Very large particles that may be visible	Particles settle rapidly	Particles can be separated by filters

When we add a solute to water, it changes several properties of the water:

- Vapor pressure
- Boiling point
- Freezing point

These changes depend only on how many solute particles are added and depend nothing on what specific compound it is.

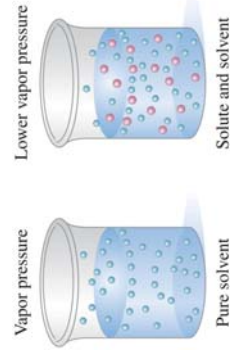
VAPOR PRESSURE LOWERING

Vapor pressure: The pressure exerted by the particles of vapor above a liquid.

In order for evaporation (and thus vapor pressure) to occur, molecules must be very near the liquid's surface.

If a nonvolatile (low vapor pressure) solute is added, solute molecules block as many solvent molecules from being near the surface. Causing a *lower* vapor pressure.

As more solute is added, more particles compete for the surface, and the lower pressure becomes lower.



BOILING POINT ELEVATION

The boiling point of a solvent is raised when a nonvolatile solute is added.

Recall: to boil, the vapor pressure must reach atmospheric pressure. If the vapor pressure is lower to begin with, it will take more heat to increase the pressure to atmospheric pressure.

FREEZING POINT DEPRESSION

- The freezing point of a solvent is lowered when a nonvolatile solvent is added.
- Recall: when a substance freezes, the molecules structure themselves very organized.
- When solute is added, it prevents the solvent molecules from becoming organized. So a lower temperature is required to force the organization.



OSMOSIS

- **Osmosis:** water molecules move through a semipermeable membrane from the solution with the **lower concentration** of solute into a solution with the **higher concentration** of solute.
 - This dilutes the more concentrated side with the eventual result of both solutions having the same concentration.
 - Only **solvent molecules** move through the membrane. Not solute.
- **Reverse osmosis:** applying pressure to force water to move from **higher to lower** concentration.
 - Removes impurities.



- ✓ 0.1—Solutions
- ✓ 0.2—Electrolytes and Nonelectrolytes
- ✓ 0.3—Solubility
- ✓ 0.4—Solution Concentrations and Reactions
- ✓ 0.5—Dilution of Solutions
- ✓ 0.6—Properties of Solutions

SOLUTIONS

Chapter 9

