SOLUTIONS

Definitions

A **solution** is a system in which one or more substances are *homogeneously* mixed or dissolved in another substance

- **homogeneous mixture**
  -- uniform appearance
  -- similar properties throughout mixture

- The **solvent** is the dissolving agent
  -- *i.e.,* the most abundant component of the solution

- The **solute** is the component that is dissolved
  -- *i.e.,* the least abundant component of the solution

Concentration of solutions

concentration -- the amount of solute dissolved in a given quantity of solvent or solution

There are many different types of concentration units:

molarity ⟵ *we will focus on this*

mass %
volume %
mass/volume %
parts per million (ppm)
parts per billion (ppb)
mole fraction
molality (not to be confused with *molarity*)

Concentration units based on the number of moles of solute

molarity -- number of moles of *solute* per liter of *solution*

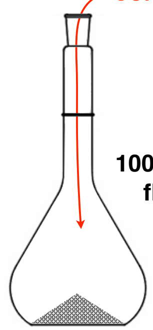
$$\textbf{Molarity} = \frac{\textbf{moles of solute}}{\textbf{liters solution}}$$

molarity has units of *moles per liter* $\left(\frac{\text{mol}}{\text{L}}\right)$

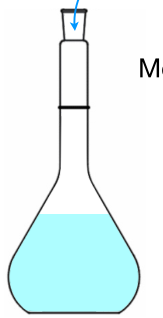
which can be abbreviated as *M*

Example: Preparation of a 1 molar solution of NaCl

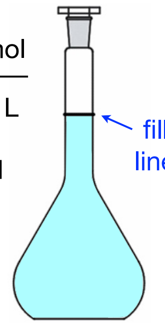
1 mole NaCl = 58.44 g NaCl



water



$$\text{Molarity} = \frac{1.000 \text{ mol}}{1.0000 \text{ L}} = \mathbf{1.000\ M}$$



Add 1 mole of NaCl to an empty 1-liter volumetric flask

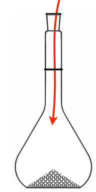
Add water to completely dissolve NaCl

Add water until the fill mark is reached and mix thoroughly

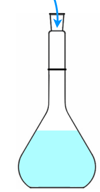
Example: Preparation of a 1 molar solution of NaCl

Note: *Dissolving 1 mole of solute to make 1 liter of solution is* ***not*** *the only way to prepare a solution with a concentration of 1 M (i.e., 1 mol / L)*

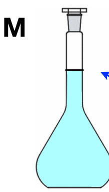
0.2500 mole NaCl = 14.61 g NaCl



water



$$\text{Molarity} = \frac{0.2500 \text{ mol}}{0.2500 \text{ L}} = \mathbf{1.000\ M}$$



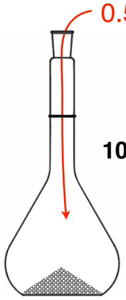
Add 0.250 mole of NaCl to an empty 250-ml volumetric flask

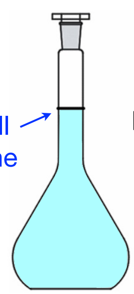
Add water to completely dissolve NaCl

Add water until the fill mark is reached and mix thoroughly

Example: Preparation of a 0.5 molar solution of NaCl

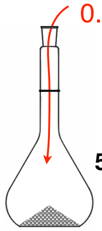
0.5000 mole NaCl = 29.22 g NaCl

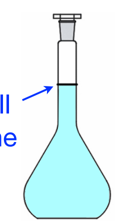




$$\text{Molarity} = \frac{0.5000 \text{ mol}}{1.0000 \text{ L}} = \mathbf{0.5000\ M}$$

0.2500 mole NaCl = 14.61 g NaCl





$$\text{Molarity} = \frac{0.2500 \text{ mol}}{0.5000 \text{ L}} = \mathbf{0.5000\ M}$$

What is the molarity of a solution made by dissolving 2.00 g of potassium chlorate in enough water to make 150. ml of solution?

Step 1: Start with the definition of molarity:

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

Step 2: Determine the number of moles of solute

Molar mass of $KClO_3$ = 39.10 + 35.45 + 3(16.00) = **122.6 g / mol**

$$2.00\ \cancel{g\ KClO_3}\left(\frac{1 \text{ mole } KClO_3}{122.6\ \cancel{g\ KClO_3}}\right) = \mathbf{0.0163 \text{ moles } KClO_3}$$

What is the molarity of a solution made by dissolving 2.00 g of potassium chloride in enough water to make 150. ml of solution?

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

Step 3: Determine the number of *liters* of solution

$$150.\ \cancel{\text{ml}} \left(\frac{1 \text{ liter}}{1000\ \cancel{\text{ml}}} \right) = \mathbf{0.150\ L}$$

What is the molarity of a solution made by dissolving 2.00 g of potassium chlorate in enough water to make 150. ml of solution?

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

Step 4: Plug values into molarity equation

$$\text{Molarity} = \frac{\mathbf{0.0163\ moles\ KClO_3}}{\mathbf{0.150\ L}}$$

$$\text{Molarity} = 0.109 \text{ moles } KClO_3 / L = \mathbf{0.109\ \mathit{M}\ \ KClO_3}$$

How many grams of potassium hydroxide are required to prepare 500. ml of 0.450 *M* KOH solution?

Step 1: Start with the definition of molarity:

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

Step 2: Determine the number of *liters* of solution

$$500.\ \cancel{\text{ml}} \left(\frac{1 \text{ liter}}{1000\ \cancel{\text{ml}}} \right) = \mathbf{0.500\ L}$$

How many grams of potassium hydroxide are required to prepare 500. ml of 0.450 *M* KOH solution?

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

Step 3: Plug known values into molarity equation and solve for unknown (moles of solute)

$$0.450\ M \text{ KOH} = \frac{x \text{ moles of KOH}}{\mathbf{0.500\ L}}$$

$$(0.500\ \cancel{L})\ 0.450 \frac{\text{moles KOH}}{\cancel{L}} = \frac{x \text{ moles of KOH}}{\cancel{\mathbf{0.500\ L}}} \cancel{(0.500\ L)}$$

$$\mathbf{0.225\ moles\ KOH} = \mathbf{\mathit{x}}$$

How many grams of potassium hydroxide are required to prepare 500. ml of 0.450 M KOH solution?

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

Step 4: Convert moles KOH to grams KOH

Molar mass of KOH = 39.10 + 16.00 + 1.008 = **56.11 g / mol**

$$0.225\ \cancel{\text{moles KOH}} \left(\frac{56.11\text{ g KOH}}{1\ \cancel{\text{mole KOH}}} \right) = \mathbf{12.6\text{ g KOH}}$$

Calculate the number of moles of nitric acid in 325 ml of 16 M HNO_3

Step 1: Start with the definition of molarity:

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

Step 2: Plug known values into molarity equation and solve for unknown (moles of solute)

$$16\ M\ HNO_3 = \frac{x\ \text{moles of } HNO_3}{\mathbf{0.325\ L}}$$

$$(0.325\ \cancel{L})\ 16\ \frac{\text{moles } HNO_3}{\cancel{L}} = \frac{x\ \text{moles of } HNO_3}{\cancel{\mathbf{0.325\ L}}}\ \cancel{(0.325\ L)}$$

$$\mathbf{5.2\ \text{moles } HNO_3 = x}$$

Concentrations of ions in aqueous solutions

When an *ionic compound* dissolves, the *concentrations* of the *individual ions* depend on the chemical formula of the compound

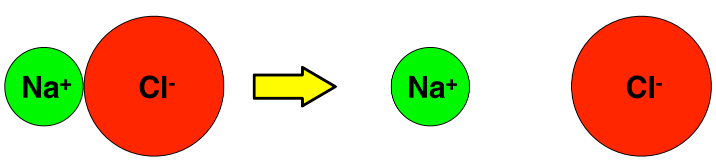
Example: *What are the concentrations of Na^+ and Cl^- ions in a 0.125 M aqueous solution of NaCl?*

$$0.125M\ NaCl = \frac{0.125\text{ mol NaCl}}{\text{L solution}}$$

$$\frac{0.125\text{ mol } Na^+}{\text{L solution}} = \mathbf{0.125}M\ \mathbf{Na^+}$$

$$\frac{0.125\text{ mol } Cl^-}{\text{L solution}} = \mathbf{0.125}\ M\ \mathbf{Cl^-}$$

Each NaCl formula unit produces 1 Na^+ ion and 1 Cl^- ion in solution



Concentrations of ions in aqueous solutions

When an *ionic compound* dissolves, the *concentrations* of the *individual ions* depend on the chemical formula of the compound

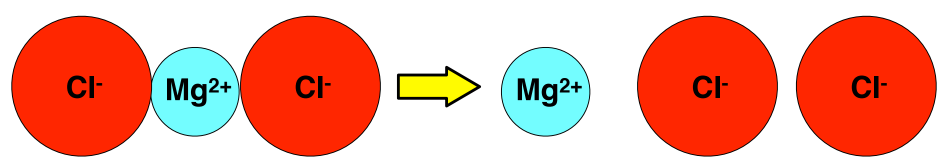
Example: *What are the concentrations of Mg^{2+} and Cl^- ions in a 0.125 M aqueous solution of $MgCl_2$?*

$$0.125M\ MgCl_2 = \frac{0.125\text{ mol } MgCl_2}{\text{L solution}}$$

$$\frac{0.125\text{ mol } Mg^{2+}}{\text{L solution}} = \mathbf{0.125}M\ \mathbf{Mg^{2+}}$$

$$\frac{0.250\text{ mol } Cl^-}{\text{L solution}} = \mathbf{0.250}M\ \mathbf{Cl^-}$$

Each $MgCl_2$ formula unit produces 1 Mg^{2+} ion and 2 Cl^- ions in solution



A solution is made by dissolving 9.82 g of copper (II) chloride ($CuCl_2$) in enough water to make 600. mL of solution. What is the molarity of Cl^- ions in solution?

Step 1: Determine the number of moles of solute

Molar mass of $CuCl_2$ = 63.55 + 2(35.45) = **134.45 g / mol**

$$9.82\ \cancel{g\ CuCl_2} \left(\frac{1\ \text{mole } CuCl_2}{134.45\ \cancel{g\ CuCl_2}} \right) = \mathbf{0.0730\ moles\ CuCl_2}$$

Step 2: Determine molarity of solute

$$\text{Molarity} = \frac{0.0730\ \text{moles } CuCl_2}{0.600\ \text{L}} = \mathbf{0.122\ \mathit{M}\ CuCl_2}$$

A solution is made by dissolving 9.82 g of copper (II) chloride ($CuCl_2$) in enough water to make 600. mL of solution. What is the molarity of Cl^- ions in solution?

Step 3: Determine the ion to solute ratio

$CuCl_2$ dissociates to give one Cu^{2+} ion and two Cl^- ions

$$CuCl_2 \longrightarrow Cu^{2+} + 2\,Cl^- \quad \Rightarrow \quad \frac{2\ \text{moles } Cl^-\ \text{ions}}{1\ \text{mole } CuCl_2}$$

Step 4: Determine molarity of the ion

$$0.122\ M\ CuCl_2 \left(\frac{2\ \text{moles } Cl^-\ \text{ions}}{1\ \text{mole } CuCl_2} \right) = \mathbf{0.244\ \mathit{M}\ Cl^-\ ions}$$

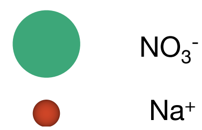
Dilutions

Dilution: Reducing the concentration of a solution by adding more solvent to the solution

- More solvent is added:
  -- *volume of the solution* increases
- No additional solute is added
  -- *number of moles of solute* stays the same

Net result: The ***molarity*** of the solution decreases

$$\downarrow \text{Molarity} = \frac{\text{moles of solute}\ \textit{(unchanged)}}{\text{liters solution} \uparrow}$$

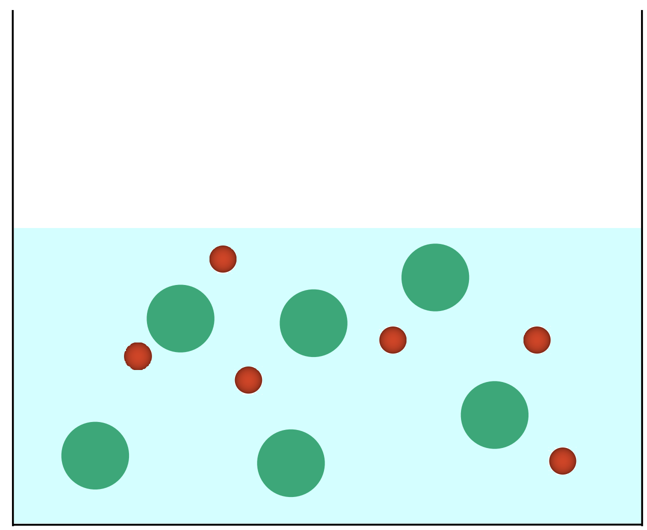


Moles = 1.0 mol

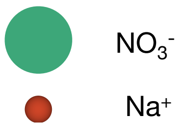
Volume = 1.0 L

$$\textbf{Molarity} = \frac{\textbf{1.0 mol}}{\textbf{1.0 L}}$$

= 1.0 M



$NaNO_3$ solution

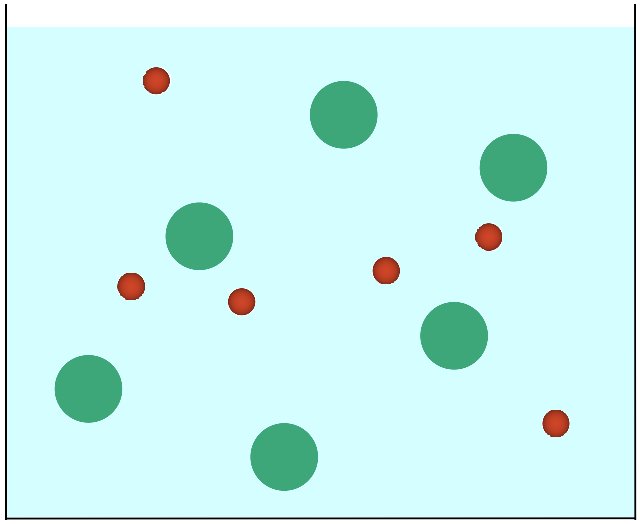


Moles = 1.0 mol

Volume = 2.0 L

$$\textbf{Molarity} = \frac{\textbf{1.0 mol}}{\textbf{2.0 L}}$$

= 0.5 M



- Solution volume is *doubled*
- Moles of solute remain the *same*
- Solution concentration is *halved*

Calculate the molarity of a solution prepared by diluting 125 ml of 0.400 *M* HCl to a final solution volume of 1.00 L.

For any dilution problem, remember that the number of moles of solute *remains the same*:

moles of solute (before) = moles of solute (after)

Based on the definition of molarity, this can be expressed as:

$$M_1 V_1 = M_2 V_2$$

Where M_1 is the molarity of the original solution

V_1 is the volume of the original solution

M_2 is the molarity of the diluted solution

V_2 is the volume of the diluted solution

Calculate the molarity of a solution prepared by diluting 125 ml of 0.400 *M* HCl to a final solution volume of 1.00 L.

$$M_1 V_1 = M_2 V_2$$

$M_1 = 0.400\ M$

$V_1 = 125\text{ ml} = 0.125\text{ L}$

$M_2 = ?$

$V_2 = 1.00\text{ L}$

$$\frac{(0.400\ M)\ (0.125\ \cancel{L})}{1.00\ \cancel{L}} = \frac{(M_2)\ \cancel{(1.00\ L)}}{\cancel{1.00\ L}}$$

$$\mathbf{0.0500\ M = M_2}$$

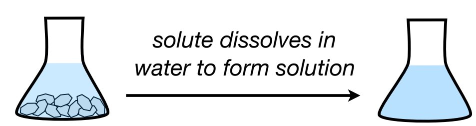
Solubility

Solubility refers to the ability of a compound to dissolve in a solvent

-- different compounds will dissolve to different extents in a given solvent

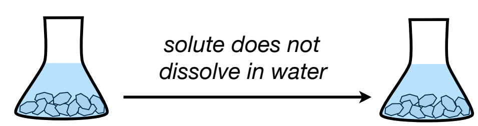
When water is the solvent:

If a solute dissolves readily in water, it is said to be ***soluble*** in water



Examples:
$NaCl$
$AgNO_3$
$(NH_4)_2CO_3$

If a solute will not dissolve in water, it is said to be ***insoluble*** in water



Examples:
$Fe(OH)_3$
$PbCl_2$
$CaCO_3$

Solubility rules

AN IONIC COMPOUND IS **SOLUBLE** IN WATER IF IT CONTAINS THE FOLLOWING IONS:	EXCEPTIONS
Ammonium ion (NH_4^+)	none
Alkali metal (Group IA) ions (Li^+, Na^+, K^+)	none
Nitrate (NO_3^-) Acetate ($C_2H_3O_2^-$)	none
Halides (Cl^-, Br^-, I^-)	Compounds containing Ag^+, Pb^{2+}, Hg_2^{2+}
Sulfate (SO_4^{2-})	Compounds containing Ag^+, Pb^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}

Solubility rules

AN IONIC COMPOUND IS **NOT SOLUBLE** IN WATER IF IT CONTAINS THE FOLLOWING IONS:	EXCEPTIONS
Carbonate (CO_3^{2-}) Phosphate (PO_4^{3-})	Compounds containing Li^+, Na^+, K^+, NH_4^+ (soluble)
Hydroxide (OH^-)	Compounds containing Li^+, Na^+, K^+, NH_4^+ (soluble) Compounds containing Ca^{2+}, Ba^{2+}, Sr^{2+} (slightly soluble)
Sulfide (S^{2-})	Compounds containing Li^+, Na^+, K^+, NH_4^+ (soluble) Compounds containing Ca^{2+}, Ba^{2+}, Sr^{2+} (soluble)

Sample problems

Are the following compounds soluble or insoluble in water?

$NaCl$

$(NH_4)_3PO_4$

$CaCO_3$

$MgSO_4$

$BaSO_4$

Double-displacement reactions

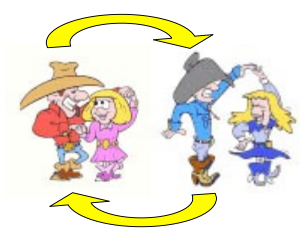
double-displacement reaction -- two ionic compounds exchange partners (*i.e.,* cations and anions) to form two different compounds

General form:

$$AB + CD \longrightarrow AD + BC$$

cations switch places

two new compounds are formed



Precipitation reactions are a type of double-displacement reaction

Reactions of aqueous solutions: *Precipitation reactions*

Precipitation reactions are a type of double-displacement reaction

In a **precipitation reaction**, an insoluble solid (called a precipitate) is formed when reactants in *aqueous solution* (*i.e.,* dissolved in water) are combined

$$BaCl_2\,(aq) \;+\; 2\,AgNO_3\,(aq) \longrightarrow 2\,AgCl\,(s) \;+\; Ba(NO_3)_2\,(aq)$$

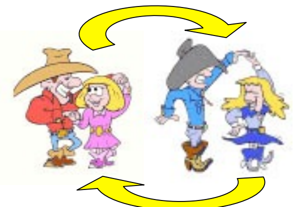
insoluble precipitate
indicated by *(s)* after its formula

Precipitation reactions

Most **precipitation reactions** occur when the anions and cations of two *aqueous ionic compounds* switch partners

General form:

AB + CD ⟶ AD + BC



To predict whether a precipitation reaction will occur:

- look at the *potential* products of the reaction (*i.e.,* make the anions and cations switch partners)
- determine whether either product is an insoluble solid

Example: Will a precipitation reaction occur when aqueous zinc chloride and potassium hydroxide are mixed?

$$ZnCl_2\,(aq) \;+\; 2\,KOH\,(aq) \longrightarrow Zn(OH)_2\,(s) \;+\; 2\,KCl\,(aq)$$

Use solubility rules to determine if either of these is an insoluble solid

Will the following reactions take place?

$$Pb(NO_3)_2(aq) \;+\; 2\,KI(aq) \longrightarrow PbI_2(s) \;+\; 2\,KNO_3(aq)$$

insoluble precipitate

$$Na_2SO_4(aq) \;+\; CuBr_2(aq) \longrightarrow 2\,NaBr(aq) + CuSO_4(aq)$$

no reaction

$$2\,CoCl_3(aq) \;+\; 3\,H_2S(aq) \longrightarrow Co_2S_3(s) + 6\,HCl(aq)$$

insoluble precipitate

For chemical reactions involving aqueous solutions, three types of equations can be written

Molecular equation

Formulas written for all reactants & products do *not* show their ionic character

-- *i.e., aqueous substances are shown as neutral compounds*

$$Pb(NO_3)_2(aq) \;+\; 2\,KI(aq) \longrightarrow PbI_2(s) \;+\; 2\,KNO_3(aq)$$

Complete ionic equation

All soluble strong electrolytes are shown as ions

-- *aqueous substances are shown as separate cations and anions*

$$Pb^{2+}(aq) + 2\,NO_3^-(aq) + 2\,K^+(aq) + 2\,I^-(aq) \longrightarrow PbI_2(s) + 2\,K^+(aq) + 2\,NO_3^-(aq)$$

Net ionic equation

Includes *only* the substances that undergo change

-- *ions that are present but do not react (spectator ions) are not shown*

$$Pb^{2+}(aq) \;+\; 2\,I^-(aq) \longrightarrow PbI_2(s)$$

Procedure for writing net ionic equations

1. Write a balanced molecular equation for the reaction

$$Pb(NO_3)_2(aq) \ + \ 2\ KI(aq) \longrightarrow PbI_2(s) \ + \ 2\ KNO_3(aq)$$

2. Rewrite equation to show aqueous substances as separate cations and anions (i.e., complete ionic equation)

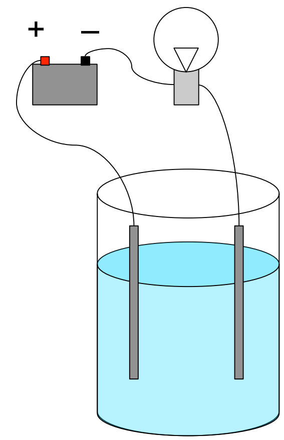
$$Pb^{2+}(aq) + \cancel{2\ NO_3^-(aq)} + \cancel{2\ K^+(aq)} + 2\ I^-(aq) \longrightarrow PbI_2(s) + \cancel{2\ K^+(aq)} + \cancel{2\ NO_3^-(aq)}$$

3. Rewrite equation after identifying and canceling spectator ions

$$Pb^{2+}(aq) \ + \ 2\ I^-(aq) \longrightarrow PbI_2(s)$$

Electrolytes

electrolyte -- a substance that forms ions when dissolved in water, resulting in a solution that conducts electricity



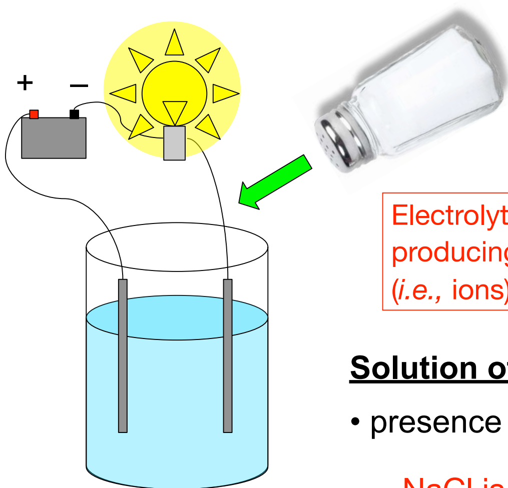
Electrolytes are capable of producing charge carriers (*i.e.,* ions) in solution

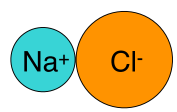
Pure water

Non-electrolyte

Electrolytes

electrolyte -- a substance that forms ions when dissolved in water, resulting in a solution that conducts electricity





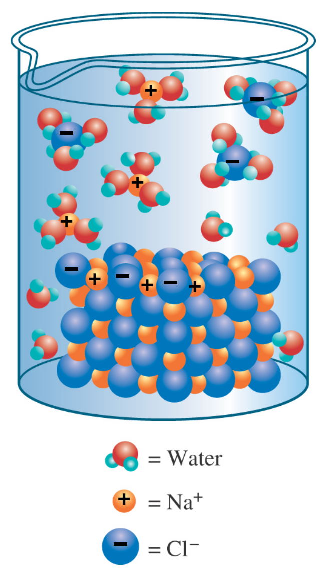
Electrolytes are capable of producing charge carriers (*i.e.,* ions) in solution

Solution of NaCl in water

- presence of Na^+ and Cl^- ions

NaCl is an ***electrolyte***

Formation of ions in solution



Electrolytes -- substances that form *aqueous solutions* containing ions

Non-electrolytes -- substances that do not form ions in solution

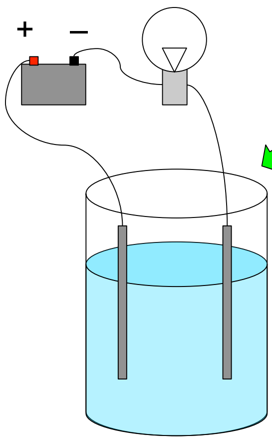
dissociation -- the separation of an ionic compound into its cations and anions as the compound dissolves

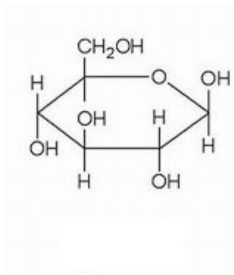
Example: Sodium chloride

$$NaCl\ (s) \longrightarrow Na^+\ (aq) \ + \ Cl^-\ (aq)$$

Electrolytes

electrolyte -- a substance that forms ions when dissolved in water, resulting in a solution that conducts electricity





Solution of sugar in water

- neutral molecules (no charge carriers)

sugar in a ***non-electrolyte***

Strong and weak electrolytes

Strong electrolytes are solutes that exist in solution *completely* or *nearly completely* as *ions*

Weak electrolytes are solutes that dissociate only *partially* to form ions in solution

-- exist primarily as *non-dissociated molecules* in solution, with only a *small fraction* in the form of *ions*

- Nearly all *soluble ionic compounds* are strong electrolytes
- *Strong acids and bases* are strong electrolytes
- *Weak acids and bases* are weak electrolytes

We will talk about strong/weak acids and bases shortly