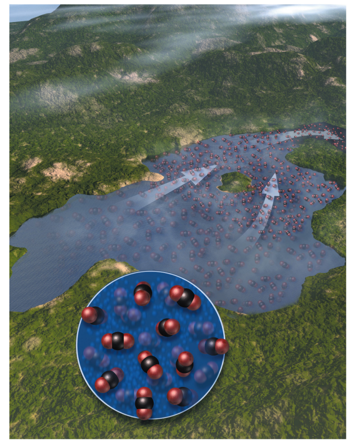
*Introductory Chemistry*, 3rd Edition
Nivaldo Tro

# Chapter 13
# Solutions





Roy Kennedy
Massachusetts Bay Community College
Wellesley Hills, MA

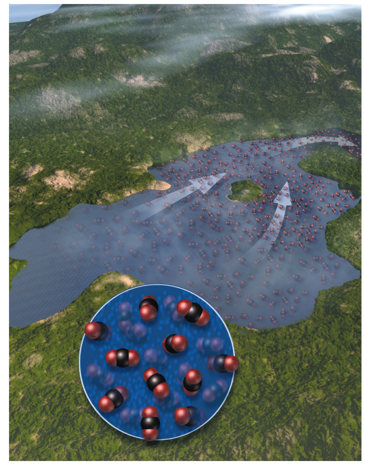
2009, Prentice Hall

# Tragedy in Cameroon

- Lake Nyos
  - ✓ Lake in Cameroon, West Africa.
  - ✓ On August 22, 1986, 1,700 people and 3,000 cattle died.
- Released carbon dioxide cloud.
  - ✓ $CO_2$ seeps in from underground and dissolves in lake water to levels above normal saturation.
  - ✓ Though not toxic, $CO_2$ is heavier than air—the people died from asphyxiation.

# Tragedy in Cameroon: A Possible Solution

- Scientists have studied Lake Nyos and similar lakes in the region to try and keep such a tragedy from reoccurring.
- Currently, they are trying to keep the $CO_2$ levels in the lake water from reaching the very high supersaturation levels by venting $CO_2$ from the lake bottom with pipes.







# Solutions

- Homogeneous mixtures.
  - ✓Composition may vary from one sample to another.
  - ✓Appears to be one substance, though really contains multiple materials.
- Most homogeneous materials we encounter are actually solutions.
  - ✓E.g., air and lake water.

# Solutions, Continued

- **Solute** is the dissolved substance.
  - ✓Seems to "disappear."
  - ✓"Takes on the state" of the solvent.
- **Solvent** is the substance solute dissolves in.
  - ✓Does not appear to change state.
- When both solute and solvent have the same state, the **solvent** is the component present in the **highest percentage**.
- Solutions in which the solvent is water are called **aqueous solutions**.



# Brass

| Type | Color | % Cu | % Zn | Density g/cm$^3$ | MP °C | Tensile strength psi | Uses |
|---|---|---|---|---|---|---|---|
| Gilding | **Reddish** | 95 | 5 | 8.86 | 1066 | 50K | Pre-83 pennies, munitions, plaques |
| Commercial | Bronze | 90 | 10 | 8.80 | 1043 | 61K | Door knobs, grillwork |
| Jewelry | Bronze | 87.5 | 12.5 | 8.78 | 1035 | 66K | Costume jewelry |
| Red | Golden | 85 | 15 | 8.75 | 1027 | 70K | Electrical sockets, fasteners, eyelets |
| Low | Deep yellow | 80 | 20 | 8.67 | 999 | 74K | Musical instruments, clock dials |
| Cartridge | Yellow | 70 | 30 | 8.47 | 954 | 76K | Car radiator cores |
| Common | Yellow | 67 | 33 | 8.42 | 940 | 70K | Lamp fixtures, bead chain |
| Muntz metal | Yellow | 60 | 40 | 8.39 | 904 | 70K | Nuts & bolts, brazing rods |

# Common Types of Solution

| Solution phase | Solute phase | Solvent phase | Example |
|---|---|---|---|
| Gaseous solutions | Gas | Gas | Air (mostly $N_2$ and $O_2$) |
| Liquid solutions | Gas | Liquid | Soda ($CO_2$ in $H_2O$) |
| | Liquid | Liquid | Vodka ($C_2H_5OH$ in $H_2O$) |
| | Solid | Liquid | Seawater (NaCl in $H_2O$) |
| Solid solutions | Solid | Solid | Brass (Zn in Cu) |



# Solubility

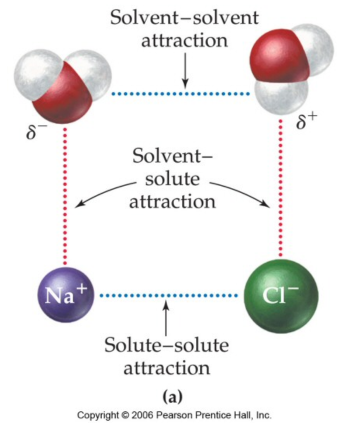
- When one substance (solute) dissolves in another (solvent) it is said to be **soluble**.
  - ✓Salt is soluble in water.
  - ✓Bromine is soluble in methylene chloride.
- When one substance does not dissolve in another it is said to be **insoluble**.
  - ✓Oil is insoluble in water.
- The solubility of one substance in another depends on two factors: nature's tendency towards mixing and the types of intermolecular attractive forces.

# Will It Dissolve?

- Chemist's rule of thumb:

  ***Like dissolves like***

- A chemical will dissolve in a solvent if it has a similar structure to the solvent.
- When the solvent and solute structures are similar, the solvent molecules will attract the solute particles at least as well as the solute particles to each other.






# Classifying Solvents

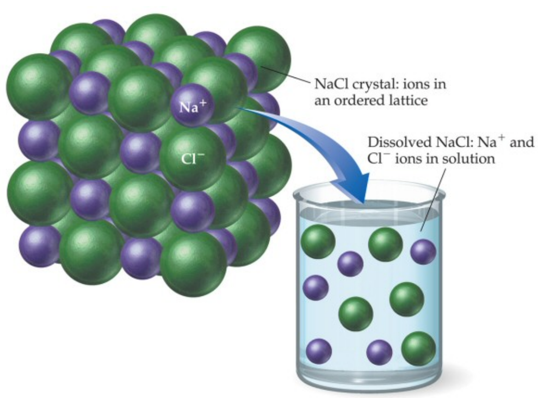
| Solvent | Class | Structural feature |
|---|---|---|
| Water, $H_2O$ | Polar | O-H |
| Ethyl alcohol, $C_2H_5OH$ | Polar | O-H |
| Acetone, $C_3H_6O$ | Polar | C=O |
| Toluene, $C_7H_8$ | Nonpolar | C-C and C-H |
| Hexane, $C_6H_{14}$ | Nonpolar | C-C and C-H |
| Diethyl ether, $C_4H_{10}O$ | Nonpolar | C-C, C-H, and C-O |

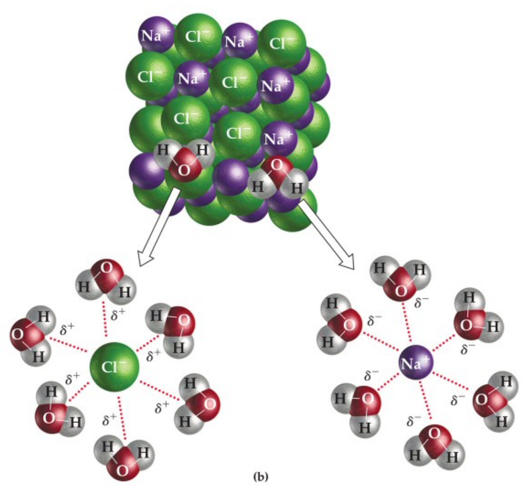
# Will It Dissolve in Water?

- Ions are attracted to polar solvents.
  - ✓ Many ionic compounds dissolve in water.
    - ➢ Generally, if the ions total charges < 4.
- Polar molecules are attracted to polar solvents.
  - ✓ Table sugar, ethyl alcohol, and glucose all dissolve well in water.
    - ➢ Have either multiple OH groups or little CH.
- Nonpolar molecules are attracted to nonpolar solvents.
  - ✓ β-carotene ($C_{40}H_{56}$) is not water soluble; it dissolves in fatty (nonpolar) tissues.
- Many molecules have both polar and nonpolar structures—whether they will dissolve in water depends on the kind, number, and location of polar and nonpolar structural features in the molecule.

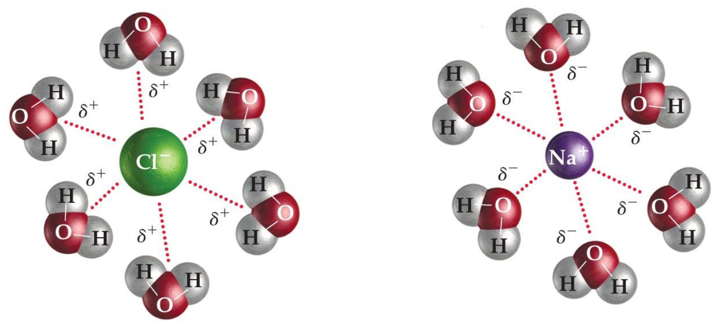


# Salt Dissolving in Water



# Solvated Ions





When materials dissolve, the solvent molecules surround the solvent particles due to the solvent's attractions for the solute. This process is called **solvation**. Solvated ions are effectively isolated from each other.



# Practice—Decide if Each of the Following Will Be Significantly Soluble in Water.

- potassium iodide, KI soluble.
- octane, $C_8H_{18}$ insoluble.
- methanol, $CH_3OH$ soluble.
- copper, **Cu** insoluble.
- cetyl alcohol, $CH_3(CH_2)_{14}CH_2OH$ insoluble.
- iron(III) sulfide, $Fe_2S_3$ insoluble.

# Solubility

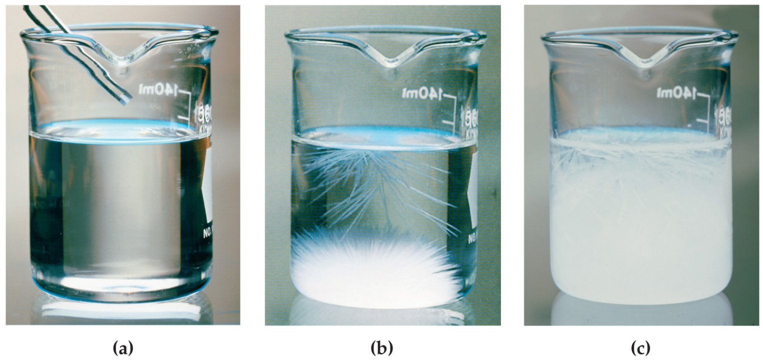
- There is usually a limit to the solubility of one substance in another.
  - ✓Gases are ***always*** soluble in each other.
  - ✓Two liquids that are mutually soluble are said to be **miscible**.
    - ➢Alcohol and water are miscible.
    - ➢Oil and water are immiscible.
- The maximum amount of solute that can be dissolved in a given amount of solvent is called **solubility**.



# Descriptions of Solubility

- **Saturated** solutions have the maximum amount of solute that will dissolve in that solvent at that temperature.
- **Unsaturated** solutions can dissolve more solute.
- **Supersaturated** solutions are holding more solute than they should be able to at that temperature.
  - ✓Unstable.

# Supersaturated Solution

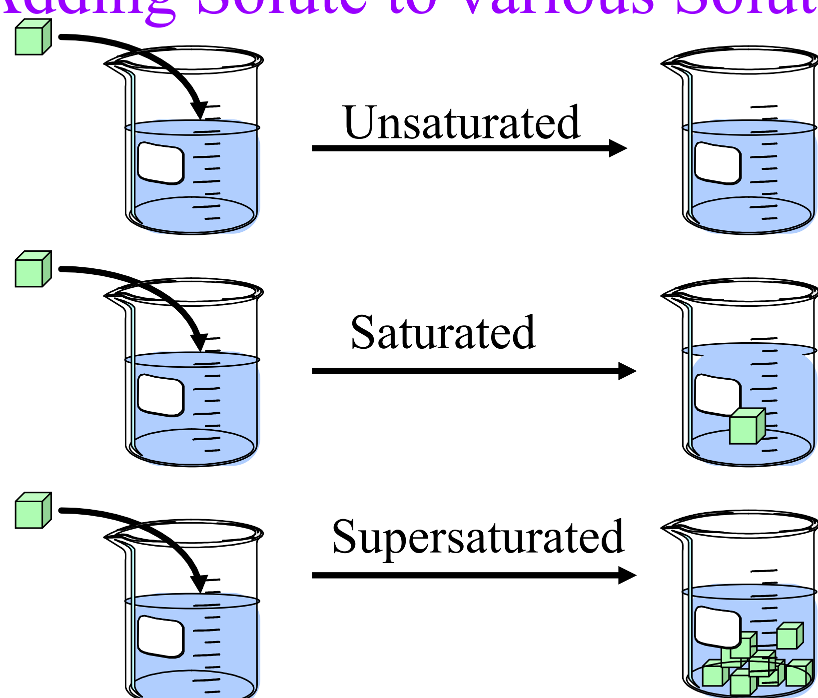





A supersaturated solution has more dissolved solute than the solvent can hold. When disturbed, all the solute above the saturation level comes out of solution.

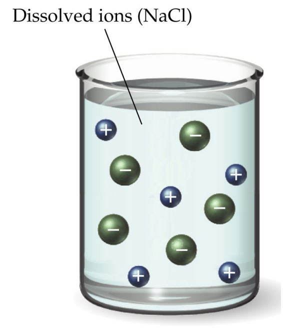


# Adding Solute to various Solutions

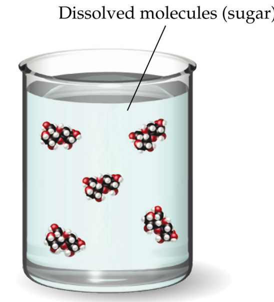


# Electrolytes

- Electrolytes are substances whose aqueous solution is a conductor of electricity.
- In **strong** electrolytes, **all** the electrolyte molecules are dissociated into ions.
- In **non**electrolytes, **none** of the molecules are dissociated into ions.
- In **weak** electrolytes, a **small percentage** of the molecules are dissociated into ions.



Electrolyte solution

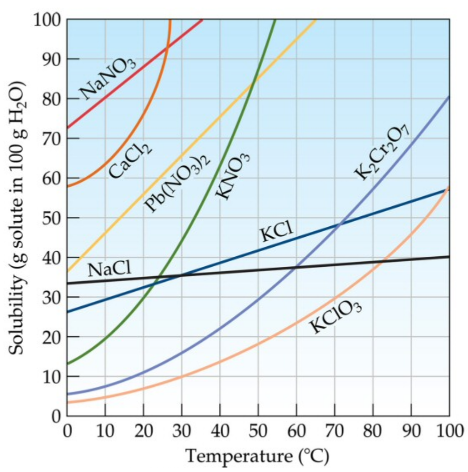


Nonelectrolyte solution



# Solubility and Temperature

- The solubility of the solute in the solvent depends on the temperature.
  - ✓ Higher temperature = Higher solubility of solid in liquid.
  - ✓ Lower temperature = Higher solubility of gas in liquid.



# Solubility and Temperature, Continued

Warm soda pop fizzes more than cold soda pop because the solubility of $CO_2$ in water decreases as temperature increases.



# Changing Temperature = Changing Solubility

- When a solution is saturated, it is holding the maximum amount of solute it can at that temperature.
- If the temperature is changed, the solubility of the solute changes.
  - ✓If a solution contains 71.3 g of $NH_4Cl$ in 100 g of water at 90 °C, it will be saturated.
  - ✓If the temperature drops to 20 °C, the saturation level of $NH_4Cl$ drops to 37.2 g.
  - ✓Therefore, 24.1 g of $NH_4Cl$ will precipitate.

# Purifying Solids: Recrystallization

- When a solid precipitates from a solution, crystals of the pure solid form by arranging the particles in a crystal lattice.
- Formation of the crystal lattice tends to reject impurities.
- To purify a solid, chemists often make a saturated solution of it at high temperature; when it cools, the precipitated solid will have much less impurity than before.



# Solubility of Gases: Effect of Temperature

- Many gases dissolve in water.
  - ✓However, most have very limited solubility.
- The solubility of a gas in a liquid decreases as the temperature increases.
  - ✓Bubbles seen when tap water is heated (before the water boils) are gases that are dissolved, coming out of the solution.
  - ✓Opposite of solids.

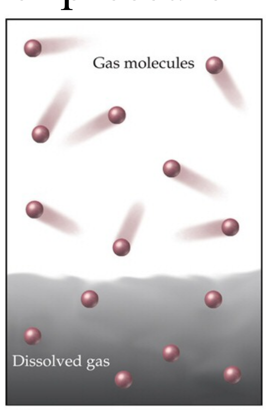
# Solubility of Gases: Effect of Pressure

- The solubility of a gas is directly proportional to its partial pressure.
  - ✓Henry's law.
  - ✓The solubility of solid is not effected by pressure.
- The solubility of a gas in a liquid increases as the pressure increases.

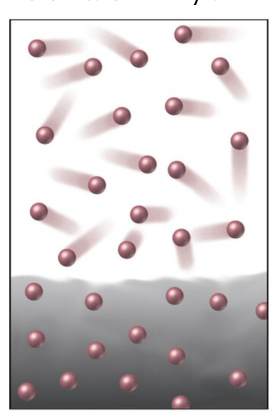


# Solubility and Pressure

- The solubility of gases in water depends on the pressure of the gas.
- Higher pressure = higher solubility.

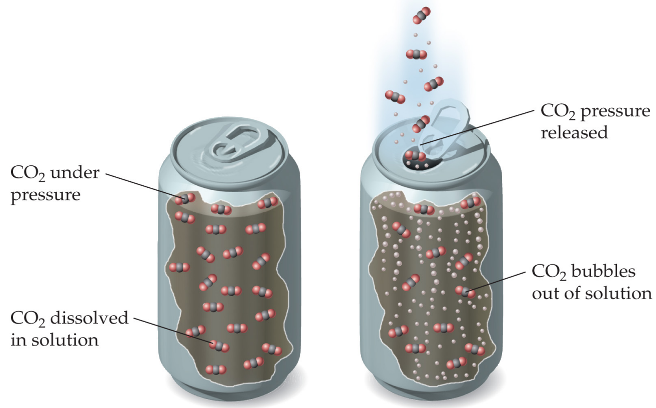



Gas at low pressure over a liquid



Gas at high pressure over a liquid

# Solubility and Pressure, Continued






When soda pop is sealed, the $CO_2$ is under pressure. Opening the container lowers the pressure, which decreases the solubility of $CO_2$ and causes bubbles to form.



# Solution Concentrations

# Describing Solutions

- Solutions have variable composition.
- To describe a solution, you need to describe both the components *and* their relative amounts.
- **Dilute** solutions have low amounts of solute per amount of solution.
- **Concentrated** solutions have high amounts of solute per amount of solution.



# Concentrations—Quantitative Descriptions of Solutions

- A more precise method for describing a solution is to quantify the amount of solute in a given amount of solution.
- **Concentration** = Amount of solute in a given amount of solution.
  - ✓Occasionally amount of solvent.

# Mass Percent

- Parts of solute in every 100 parts solution.
  - ✓ If a solution is 0.9% by mass, then there are 0.9 grams of solute in every 100 grams of solution.
    - ➢ Or 10 kg solute in every 100 kg solution.
- Since masses are additive, the mass of the solution is the sum of the masses of solute and solvent.

$$\text{Mass Percent} = \frac{\text{Mass of Solute, g}}{\text{Mass of Solution, g}} \times 100\%$$

$$\text{Mass of Solute} + \text{Mass of Solvent} = \text{Mass of Solution}$$



# Example 13.1—Calculate the Mass Percent of a Solution Containing 27.5 g of Ethanol in 175 mL $H_2O$.

| | |
|---|---|
| **Given:** | 27.5 g ethanol, 175 mL $H_2O$, 202.5 g solution |
| **Find:** | % by mass |
| **Solution Map:** | g EtOH, g $H_2O$ → g $H_2O$ → g sol'n → % <br> $\frac{1.00\text{ g }H_2O}{1\text{ mL }H_2O}$   g solute + g solvent = g solution   $\%\text{ by Mass} = \frac{\text{g solute}}{\text{g solution}} \times 100\%$ |
| **Relationships:** | 1 mL $H_2O$ = 1.00 g |
| **Solve:** | $175\text{ mL }H_2O \times \frac{1.00\text{ g }H_2O}{1\text{ mL }H_2O} = 175\text{ g }H_2O$ <br> $27.5\text{ g ethanol} + 175\text{ g }H_2O = 202.5\text{ g solution}$ <br> $\%\text{ by Mass} = \frac{27.5\text{ g ethanol}}{202.5\text{ g solution}} \times 100\% = 13.6\%$ |
| **Check:** | The answer seems reasonable as it is less than 100%. |

Example 13.1:

- Calculate the mass percent of a solution containing 27.5 g of ethanol ($C_2H_6O$) and 175 mL of $H_2O$. (Assume the density of $H_2O$ is 1.00 g/mL.)



Example:
Calculate the mass percent of a solution containing 27.5 g of ethanol ($C_2H_6O$) and 175 mL of $H_2O$. (Assume the density of $H_2O$ is 1.00 g/mL.)

- Write down the given quantity and its units.

Given: 27.5 g $C_2H_6O$

175 mL $H_2O$

| Example: Calculate the mass percent of a solution containing 27.5 g of ethanol ($C_2H_6O$) and 175 mL of $H_2O$. (Assume the density of $H_2O$ is 1.00 g/mL.) | Information: Given: 27.5 g $C_2H_6O$; 175 mL $H_2O$ |
|---|---|

- Write down the quantity to find and/or its units.

  Find: mass percent



| Example: Calculate the mass percent of a solution containing 27.5 g of ethanol ($C_2H_6O$) and 175 mL of $H_2O$. (Assume the density of $H_2O$ is 1.00 g/mL.) | Information:<br>Given: 27.5 g $C_2H_6O$; 175 mL $H_2O$<br>Find: % by mass |
|---|---|

- Collect needed equations:

$$\text{Mass Percent} = \frac{\text{Mass of Solute, g}}{\text{Mass of Solution, g}} \times 100\%$$

Example:
Calculate the mass percent of a solution containing 27.5 g of ethanol ($C_2H_6O$) and 175 mL of $H_2O$. (Assume the density of $H_2O$ is 1.00 g/mL.)

Information:
Given: 27.5 g $C_2H_6O$; 175 mL $H_2O$
Find: % by mass
Equation: $\text{Mass \%} = \frac{\text{g Solute}}{\text{g Sol'n}} \times 100\%$

- Collect needed conversion factors:

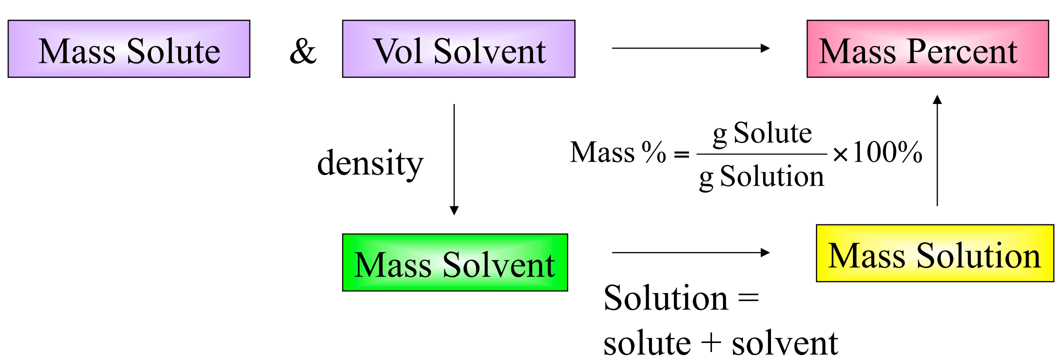
$$d(H_2O) = 1.00 \text{ g/mL} \;\therefore\; 1.00 \text{ g } H_2O = 1 \text{ mL } H_2O$$



Example:
Calculate the mass percent of a solution containing 27.5 g of ethanol ($C_2H_6O$) and 175 mL of $H_2O$. (Assume the density of $H_2O$ is 1.00 g/mL.)

Information:
Given: 27.5 g $C_2H_6O$; 175 mL $H_2O$
Find: % by mass
Equation: $\text{Mass \%} = \frac{\text{g Solute}}{\text{g Solution}} \times 100\%$
Conversion Factor:
1.00 g $H_2O$ = 1 mL $H_2O$

- Design a solution map:



Example:
Calculate the mass percent of a solution containing 27.5 g of ethanol ($C_2H_6O$) and 175 mL of $H_2O$. (Assume the density of $H_2O$ is 1.00 g/mL.)

Information:
Given: 27.5 g $C_2H_6O$; 175 mL $H_2O$
Find: % by mass
Equation: $\text{Mass \%} = \frac{\text{g Solute}}{\text{g Solution}} \times 100\%$
Conversion Factor:
1.00 g $H_2O$ = 1 mL $H_2O$
Solution Map:
mass solution and volume solvent →
mass solvent → mass solution → mass percent

- Apply the solution maps:

$$175\ \text{mL}\ H_2O \times \frac{1.00\ \text{g}\ H_2O}{1\ \text{mL}\ H_2O} = 175\ \text{g}\ H_2O$$

Mass of solution = mass $C_2H_6O$ + mass $H_2O$
= 27.5 g $C_2H_6O$ + 175 g $H_2O$
= 202.5 g



Example:
Calculate the mass percent of a solution containing 27.5 g of ethanol ($C_2H_6O$) and 175 mL of $H_2O$. (Assume the density of $H_2O$ is 1.00 g/mL.)

Information:
Given: 27.5 g $C_2H_6O$; 175 mL $H_2O$
Find: % by mass
Equation: $\text{Mass \%} = \frac{\text{g Solute}}{\text{g Solution}} \times 100\%$
Conversion Factor:
1.00 g $H_2O$ = 1 mL $H_2O$
Solution Map:
mass solution and volume solvent →
mass solvent → mass solution → mass percent

- Apply the solution maps and equation:

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

$$\text{Mass Percent} = \frac{27.5\ \text{g}\ C_2H_6O}{202.5\ \text{g solution}} \times 100\%$$

= 13.5802%
= 13.6%

Example:
Calculate the mass percent of a solution containing 27.5 g of ethanol ($C_2H_6O$) and 175 mL of $H_2O$. (Assume the density of $H_2O$ is 1.00 g/mL.)

Information:

Given: 27.5 g $C_2H_6O$; 175 mL $H_2O$

Find: % by mass

Equation: $\text{Mass \%} = \frac{\text{g Solute}}{\text{g Solution}} \times 100\%$

Conversion Factor:

1.00 g $H_2O$ = 1 mL $H_2O$

Solution Map:

mass solution and volume solvent → mass solvent → mass solution → mass percent

- Check the solution:

mass percent = 13.6%

The units of the answer, %, are correct.
The magnitude of the answer makes sense since the mass of solute is less than the mass of solvent.



## Practice—Calculate the Mass Percent of a Solution that Has 10.0 g of $I_2$ Dissolved in 150.0 g of Ethanol.

## Practice—Calculate the Mass Percent of a Solution that Has 10.0 g of $I_2$ Dissolved in 150.0 g of Ethanol, Continued

| | |
|---|---|
| **Given:** | 10.0 g $I_2$, 150.0 g ethanol / 160.0 g solution $H_2O$ |
| **Find:** | % by mass |
| **Solution Map:** | g EtOH, g $H_2O$ → g sol'n → % |
| **Relationships:** | $\text{g solute} + \text{g solvent} = \text{g solution}$ $\quad \%\text{ by Mass} = \frac{\text{g solute}}{\text{g solution}} \times 100\%$ |
| **Solve:** | $10.0\text{ g } I_2 + 150.0\text{ g ethanol} = 160.0\text{ g solution}$ <br> $\%\text{ by Mass} = \frac{10.0\text{ g } I_2}{160.0\text{ g solution}} \times 100\%$ <br> $= 6.25\%$ |
| **Check:** | The answer seems reasonable as it is less than 100%. |

# Using Concentrations as Conversion Factors

- Concentrations show the relationship between the amount of solute and the amount of solvent.
  - ✓ 12% by mass sugar (*aq*) means 12 g sugar ≡ 100 g solution.
- The concentration can then be used to convert the amount of solute into the amount of solution or visa versa.

## Example 13.2—What Volume of 11.5% by Mass Soda Contains 85.2 g of Sucrose?

| | |
|---|---|
| **Given:** | 85.2 g sugar |
| **Find:** | volume, mL |
| **Solution Map:** | g solute → $\frac{100\text{ g sol'n}}{11.5\text{ g sucrose}}$ → g sol'n → $\frac{1\text{ mL sol'n}}{1.00\text{ g sol'n}}$ → mL sol'n |
| **Relationships:** | 100 g sol'n = 11.5 g sugar, 1 mL solution = 1.00 g |
| **Solve:** | $85.2\text{ g sugar} \times \frac{100\text{ g}}{11.5\text{ g sugar}} \times \frac{1\text{ mL}}{1.00\text{ g}} = 741\text{ mL}$ |
| **Check:** | The unit is correct. The magnitude seems reasonable as the mass of sugar ≈ 10% the volume of solution. |



Example 13.2:

- A soft drink contains 11.5% sucrose ($C_{12}H_{22}O_{11}$) by mass. What volume of soft drink in milliliters contains 85.2 g of sucrose? (Assume the density is 1.00 g/mL.)

Example:
A soft drink contains 11.5% sucrose ($C_{12}H_{22}O_{11}$) by mass. What volume of soft drink in milliliters contains 85.2 g of sucrose? (Assume the density is 1.00 g/mL.)

- Write down the given quantity and its units.

Given: 85.2 g $C_{12}H_{22}O_{11}$



Example:
A soft drink contains 11.5% sucrose ($C_{12}H_{22}O_{11}$) by mass. What volume of soft drink in milliliters contains 85.2 g of sucrose? (Assume the density is 1.00 g/mL.)

Information:
Given: 85.2 g $C_{12}H_{22}O_{11}$

- Write down the quantity to find and/or its units.

Find: volume of solution, mL

| Example: A soft drink contains 11.5% sucrose ($C_{12}H_{22}O_{11}$) by mass. What volume of soft drink in milliliters contains 85.2 g of sucrose? (Assume the density is 1.00 g/mL.) | Information:<br>Given: 85.2 g $C_{12}H_{22}O_{11}$<br>Find: mL solution |
|---|---|

- Collect needed conversion factors:

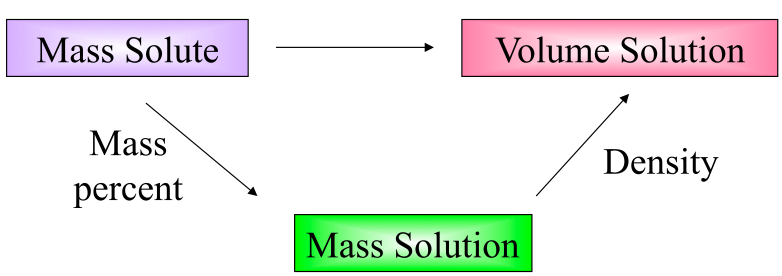
$$11.5\% \text{ by mass } \therefore\ 11.5 \text{ g } C_{12}H_{22}O_{11} \equiv 100 \text{ g solution}$$

$$d = 1.00 \text{ g/mL } \therefore\ 1.00 \text{ g solution} = 1 \text{ mL solution}$$



| Example: A soft drink contains 11.5% sucrose ($C_{12}H_{22}O_{11}$) by mass. What volume of soft drink in milliliters contains 85.2 g of sucrose? (Assume the density is 1.00 g/mL.) | Information:<br>Given: 85.2 g $C_{12}H_{22}O_{11}$<br>Find: mL solution<br>Conversion Factors:<br>11.5 g $C_{12}H_{22}O_{11}$ ≡ 100 g solution<br>1.00 g solution = 1 mL solution |
|---|---|

- Design a solution map:



Example:
A soft drink contains 11.5% sucrose ($C_{12}H_{22}O_{11}$) by mass. What volume of soft drink in milliliters contains 85.2 g of sucrose? (Assume the density is 1.00 g/mL.)

Information:
Given: 85.2 g $C_{12}H_{22}O_{11}$
Find: mL solution
Conversion Factors:
11.5 g $C_{12}H_{22}O_{11}$ ≡ 100 g solution
1.00 g solution = 1 mL solution
Solution Map:
g sucrose → g solution → mL solution

- Apply the solution map:

$$85.2\ \text{g}\ C_{12}H_{22}O_{11} \times \frac{100\ \text{g solution}}{11.5\ \text{g}\ C_{12}H_{22}O_{11}} \times \frac{1\ \text{mL solution}}{1.00\ \text{g solution}}$$

$= 740.87$ mL

$= 741$ mL



Example:
A soft drink contains 11.5% sucrose ($C_{12}H_{22}O_{11}$) by mass. What volume of soft drink in milliliters contains 85.2 g of sucrose? (Assume the density is 1.00 g/mL.)

Information:
Given: 85.2 g $C_{12}H_{22}O_{11}$
Find: mL solution
Conversion Factors:
11.5 g $C_{12}H_{22}O_{11}$ ≡ 100 g solution
1.00 g solution = 1 mL solution
Solution Map:
g sucrose → g solution → mL solution

- Check the solution:

volume of solution = 741 mL

The units of the answer, mL, are correct.
The magnitude of the answer makes sense since the mass of solute is less than the volume of solution.

## Practice—Milk Is 4.5% by Mass Lactose. Determine the Mass of Lactose in 175 g of Milk.



## Practice—Milk Is 4.5% by Mass Lactose. Determine the Mass of Lactose in 175 g of Milk, Continued.

**Given:** 175 g milk ≡ 175 g solution

**Find:** g lactose

**Equivalence:** 4.5 g lactose ≡ 100 g solution

**Solution Map:** g solution → g Lactose

$$\frac{4.5 \text{ g Lactose}}{100 \text{ g solution}}$$

**Apply Solution Map:**

$$175 \text{ g solution} \times \frac{4.5 \text{ g Lactose}}{100 \text{ g solution}} = 7.9 \text{ g Lactose}$$

**Check Answer:**

Units are correct. Number makes sense because lactose is a component of the mixture, therefore, its amount should be less.

# Preparing a Solution

- Need to know **amount** of solution and **concentration** of solution.
- Calculate the mass of solute needed.
  - ✓Start with amount of solution.
  - ✓Use concentration as a conversion factor.
    - ➢5% by mass ⇒ 5 g solute ≡ 100 g solution.
  - ✓"Dissolve the grams of solute in enough solvent to total the total amount of solution."



# Example—How Would You Prepare 250.0 g of 5.00% by Mass Glucose Solution (Normal Glucose)?

**Given:** 250.0 g solution

**Find:** g glucose

**Equivalence:** 5.00 g glucose ≡ 100 g solution

**Solution Map:** g solution → g glucose

$$\frac{5.00\ \text{g Glucose}}{100\ \text{g solution}}$$

**Apply Solution Map:**

$$250.0\ \cancel{\text{g solution}} \times \frac{5.00\ \text{g glucose}}{100\ \cancel{\text{g solution}}} = 12.5\ \text{g glucose}$$

**Answer:** Dissolve 12.5 g of glucose in enough water to total 250.0 g.

## Practice—How Would You Prepare 450.0 g of 15.0% by Mass Aqueous Ethanol?



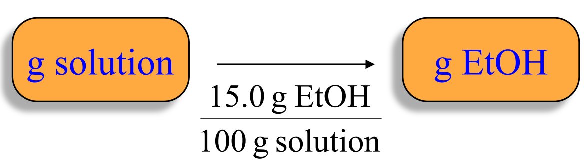
## Practice—How Would You Prepare 450.0 g of 15.0% by Mass Aqueous Ethanol?, Continued

**Given:** 450.0 g solution

**Find:** g ethanol (EtOH)

**Equivalence:** 15.0 g EtOH ≡ 100 g solution

**Solution map:**




**Apply solution map**:

$$450.0\ \cancel{\text{g solution}} \times \frac{15.0\ \text{g EtOH}}{100\ \cancel{\text{g solution}}} = 67.5\ \text{g EtOH}$$

**Answer:**

Dissolve 67.5 g of ethanol in enough water to total 450.0 g.

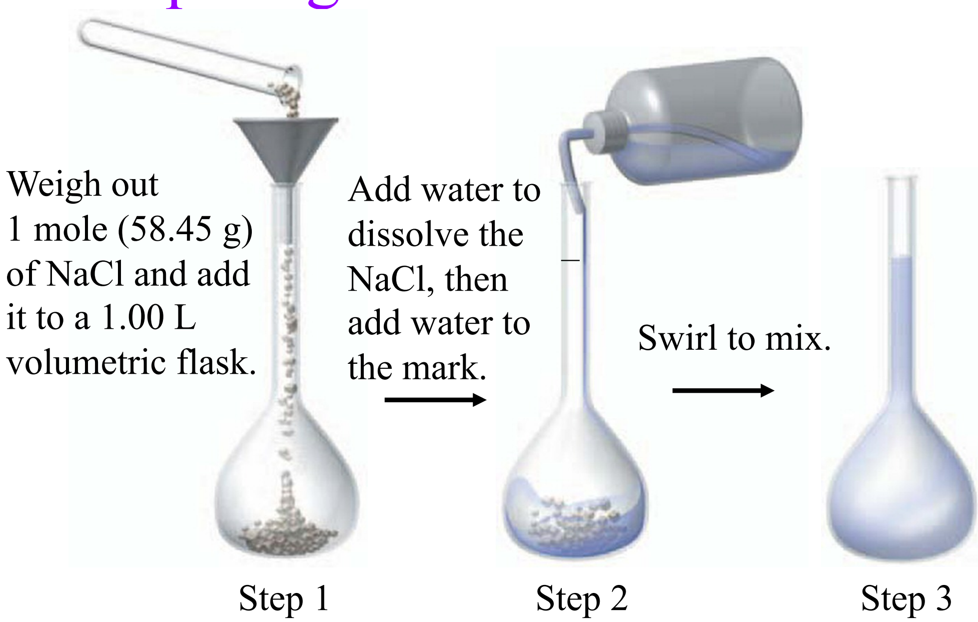
# Solution Concentration
# Molarity

- Moles of solute per 1 liter of solution.
- Used because it describes how many molecules of solute in each liter of solution.
- If a sugar solution concentration is 2.0 M , 1 liter of solution contains 2.0 moles of sugar, 2 liters = 4.0 moles sugar, 0.5 liters = 1.0 mole sugar:

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



# Preparing a 1.00 M NaCl Solution



## Example 13.3—Calculate the Molarity of a Solution Made by Dissolving 15.5 g of NaCl in 1.50 L of Solution

| | |
|---|---|
| **Given:** | 15.5 g NaCl, 1.50 L solution |
| **Find:** | M |
| **Solution Map:** | g NaCl → mol NaCl; mol NaCl + L solution → M, $M = \frac{mol}{L}$ |
| **Relationships:** | M = mol/L, 1 mol NaCl = 58.44 g |
| **Solve:** | $15.5 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 0.2652 \text{ mol NaCl}$ <br> $M = \frac{0.2652 \text{ mol NaCl}}{1.50 \text{ L}}$ <br> $M = 0.177 \text{ M}$ |
| **Check:** | The unit is correct, the magnitude is reasonable. |



Example 13.3:

- Calculate the molarity of a solution made by putting 15.5 g of NaCl into a beaker and adding water to make 1.50 L of NaCl solution.

Example:
Calculate the molarity of a solution made by putting 15.5 g of NaCl into a beaker and adding water to make 1.50 L of NaCl solution.

- Write down the given quantity and its units.

Given: 15.5 g NaCl

1.50 L solution



Example:
Calculate the molarity of a solution made by putting 15.5 g of NaCl into a beaker and adding water to make 1.50 L of NaCl solution.

Information:
Given: 15.5 g NaCl; 1.50 L solution

- Write down the quantity to find and/or its units.

Find: molarity (M)

| Example: Calculate the molarity of a solution made by putting 15.5 g of NaCl into a beaker and adding water to make 1.50 L of NaCl solution. | Information: Given: 15.5 g NaCl; 1.50 L solution Find: molarity, M |
|---|---|

- Collect needed equations and conversion factors:

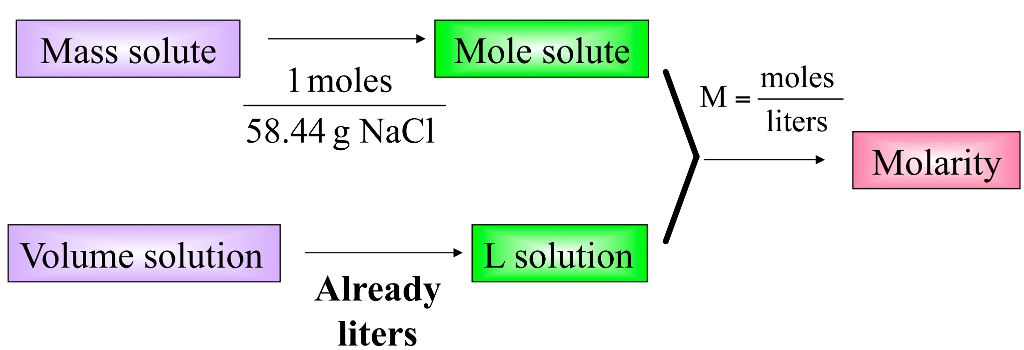
$$\text{Molarity} = \frac{\text{moles solute}}{\text{liters solution}}$$

$$\text{Molar mass NaCl} = 58.44 \text{ g/mol} \therefore 58.44 \text{ g NaCl} = 1 \text{ mol.}$$



| Example: Calculate the molarity of a solution made by putting 15.5 g of NaCl into a beaker and adding water to make 1.50 L of NaCl solution. | Information: Given: 15.5 g NaCl; 1.50 L solution Find: molarity, M Conversion Factors: 58.44 g = 1 mol NaCl; $\text{Molarity} = \frac{\text{moles solute}}{\text{liters solution}}$ |
|---|---|

- Design a solution map:



| Example: Calculate the molarity of a solution made by putting 15.5 g of NaCl into a beaker and adding water to make 1.50 L of NaCl solution. | Information: Given: 15.5 g NaCl; 1.50 L solution<br>Find: molarity, M<br>Conversion Factors: 58.44 g = 1 mol NaCl; $\text{Molarity} = \frac{\text{moles solute}}{\text{liters solution}}$ |
|---|---|

- Apply the solution map:

$$15.5\ \cancel{g}\ \text{NaCl} \times \frac{1\ \text{mol NaCl}}{58.44\ \cancel{g}} = 0.26\underline{5}2\ \text{mol NaCl}$$

$$\text{Molarity} = \frac{\text{moles solute}}{\text{liters solution}} = \frac{0.26\underline{5}2\ \text{mol NaCl}}{1.50\ \text{L solution}}$$

$$= 0.177\ \text{M NaCl}$$



| Example: Calculate the molarity of a solution made by putting 15.5 g of NaCl into a beaker and adding water to make 1.50 L of NaCl solution. | Information: Given: 15.5 g NaCl; 1.50 L solution<br>Find: molarity, M<br>Conversion Factors: 58.44 g = 1 mol NaCl; $\text{Molarity} = \frac{\text{moles solute}}{\text{liters solution}}$ |
|---|---|

- Check the solution:

Molarity of solution = 0.177 M

The units of the answer, M, are correct.
The magnitude of the answer makes sense since the mass of solute is less than the 1 mole and the volume is more than 1 L.

# Practice—What Is the Molarity of a Solution Containing 3.4 g of $NH_3$ (MM 17.03) in 200.0 mL of Solution?



# Practice—What Is the Molarity of a Solution Containing 3.4 g of $NH_3$ (MM 17.03) in 200.0 mL of Solution?, Continued

| | |
|---|---|
| **Given:** | 3.4 g $NH_3$, 200.0 mL solution |
| **Find:** | M |
| **Solution Map:** | g $NH_3$ → mol $NH_3$; mL sol'n → L sol'n; $M = \frac{mol}{L}$ → M |
| **Relationships:** | M = mol/L, 1 mol $NH_3$ = 17.03 g, 1 mL = 0.001 L |
| **Solve:** | $3.4\ \text{g NH}_3 \times \frac{1\ \text{mol NH}_3}{17.03\ \text{g NH}_3} = 0.20\ \text{mol}$<br>$200.0\ \text{mL} \times \frac{0.001\ \text{L}}{1\ \text{mL}} = 0.2000\ \text{L}$<br>$M = \frac{0.20\ \text{mol NH}_3}{0.2000\ \text{L}}$<br>$M = 1.0\ M$ |
| **Check:** | The unit is correct, the magnitude is reasonable. |

# Using Concentrations as Conversion Factors

- Concentrations show the relationship between the amount of solute and the amount of solvent.
  - ✓ 0.12 M sugar (*aq*) means 0.12 mol sugar ≡ 1.0 L solution.
- The concentration can then be used to convert the moles of solute into the liters of solution, or visa versa.
- Since we normally measure the amount of solute in grams, we will need to convert between grams and moles.



# Example 13.4—How Many Liters of a 0.114 M NaOH Solution Contains 1.24 mol of NaOH?

| | |
|---|---|
| **Given:**<br>**Find:** | 1.24 mol NaOH<br>volume, L |
| **Solution Map:**<br><br>**Relationships:** | **mol NaOH** → **L solution** $\frac{1.00\ \text{L sol'n}}{0.114\ \text{mol NaOH}}$<br><br>1.00 L solution = 0.114 mol NaOH |
| **Solve:** | $1.24\ \text{mol NaOH} \times \frac{1.00\ \text{L}}{0.114\ \text{mol NaOH}} = 10.9\ \text{L}$ |
| **Check:** | The unit is correct, the magnitude seems reasonable as the moles of NaOH > 10x the amount in 1 L. |

Example 13.4:

- How many liters of a 0.114 M NaOH solution contains 1.24 mol of NaOH?



Example:
How many liters of a 0.114 M NaOH solution contains 1.24 mol of NaOH?

- Write down the given quantity and its units.

Given: 1.24 mol NaOH

Example:
How many liters of a 0.114 M NaOH solution contains 1.24 mol of NaOH?

Information:
Given: 1.24 mol NaOH

- Write down the quantity to find and/or its units.

Find: volume of solution (L)



Example:
How many liters of a 0.114 M NaOH solution contains 1.24 mol of NaOH?

Information:
Given: 1.24 mol NaOH
Find: L solution

- Collect needed conversion factors:

Molarity = 0.114 mol/L ∴ 0.114 mol NaOH = 1 L solution.

| Example: How many liters of a 0.114 M NaOH solution contains 1.24 mol of NaOH? | Information:<br>Given: 1.24 mol NaOH<br>Find: L solution<br>Conversion Factor:<br>0.114 mol = 1 L |
|---|---|

- Design a solution map:

Mole solute → L solution

$$\frac{1\text{ L solution}}{0.114\text{ moles}}$$



| Example: How many liters of a 0.114 M NaOH solution contains 1.24 mol of NaOH? | Information:<br>Given: 1.24 mol NaOH<br>Find: L solution<br>Conversion Factor:<br>0.114 mol = 1 L<br>Solution Map: mol → L |
|---|---|

- Apply the solution map:

$$1.24\ \cancel{\text{mol}}\ \text{NaOH} \times \frac{1\text{ L solution}}{0.114\ \cancel{\text{moles}}} = 10.9\text{ L solution}$$

| Example: How many liters of a 0.114 M NaOH solution contains 1.24 mol of NaOH? | Information:<br>Given: 1.24 mol NaOH<br>Find: L solution<br>Conversion Factor:<br>0.114 mol = 1 L<br>Solution Map: mol → L |
|---|---|

- Check the solution:

Volume of solution = 10.9 L

The units of the answer, L, are correct.
The magnitude of the answer makes sense.
Since 1 L only contains 0.114 moles,
the volume must be more than 1 L.



# Practice—Determine the Mass of $CaCl_2$ (MM = 110.98) in 1.75 L of 1.50 M Solution.

## Practice—Determine the Mass of $CaCl_2$ (MM = 110.98) in 1.75 L of 1.50 M Solution, Continued.

**Given:** 1.75 L solution

**Find:** g $CaCl_2$

**Equivalence:** 1.50 mol $CaCl_2$ ≡ 1 L solution; 110.98 g = 1 mol $CaCl_2$

**Solution Map:**

L solution → $\frac{1.50 \text{ mol } CaCl_2}{1 \text{ L}}$ → mol $CaCl_2$ → $\frac{110.98 \text{ g}}{1 \text{ mol } CaCl_2}$ → g $CaCl_2$

**Apply Solution Map:**

$$1.75 \text{ L solution} \times \frac{1.50 \text{ mol } CaCl_2}{1 \text{ L}} \times \frac{110.98 \text{ g}}{1 \text{ mol } CaCl_2} = 291 \text{ g } CaCl_2$$

**Check Answer:**
Units are correct.



## Practice—How Many Grams of $CuSO_4 \cdot 5\ H_2O$ (MM 249.69) are in 250.0 mL of a 1.00 M Solution?

## Practice—How Many Grams of $CuSO_4 \cdot 5\ H_2O$ (MM 249.69) are in 250.0 mL of a 1.00 M Solution?, Continued

| | |
|---|---|
| **Given:** | 250.0 mL solution |
| **Find:** | mass $CuSO_4 \cdot 5\ H_2O$, g |
| **Solution Map:** | mL sol'n → L sol'n → mol $CuSO_4$ → g $CuSO_4$; $\frac{0.001\ \text{L}}{1\ \text{mL}}$, $\frac{1.00\ \text{mol}}{1\ \text{L sol'n}}$, $\frac{249.69\ \text{g}}{1\ \text{mol}}$ |
| **Relationships:** | 1.00 L solution = 1.00 mol; 1 mL = 0.001 L; 1 mol = 249.69 g |
| **Solve:** | $250.0\ \cancel{\text{mL}} \times \frac{0.001\ \cancel{\text{L}}}{1\ \cancel{\text{mL}}} \times \frac{1.00\ \cancel{\text{mol}}\ CuSO_4 \cdot 5H_2O}{1\ \cancel{\text{L}}} \times \frac{249.69\ \text{g}}{1\ \cancel{\text{mol}}\ CuSO_4 \cdot 5H_2O} = 62.4\ \text{g}\ CuSO_4 \cdot 5H_2O$ |
| **Check:** | The unit is correct, the magnitude seems reasonable as the volume is ¼ of a liter. |



## Example—How Would You Prepare 250 mL of 0.20 M NaCl?

**Given:** 250 mL solution

**Find:** g NaCl

**Equivalence:** 0.20 moles NaCl ≡ 1 L solution; 0.001 L = 1 mL; 58.44 g = 1 mol NaCl

**Solution Map:**

**mL solution** → ($\frac{0.001\ \text{L}}{1\ \text{mL}}$) → **L solution** → ($\frac{0.20\ \text{mol NaCl}}{1\ \text{L}}$) → **moles NaCl** → ($\frac{58.44\ \text{g}}{1\ \text{mol NaCl}}$) → **g NaCl**

**Apply Solution Map:**

$$250\ \cancel{\text{mL}} \times \frac{0.001\ \cancel{\text{L}}}{1\ \cancel{\text{mL}}} \times \frac{0.20\ \cancel{\text{mol}}\ \text{NaCl}}{1\ \cancel{\text{L}}} \times \frac{58.44\ \text{g}}{1\ \cancel{\text{mol}}\ \text{NaCl}} = 2.9\ \text{g NaCl}$$

**Answer:**

Dissolve 2.9 g of NaCl in enough water to total 250 mL.

## Practice—How Would You Prepare 100.0 mL of 0.100 M $K_2SO_4$ (MM = 174.26)?



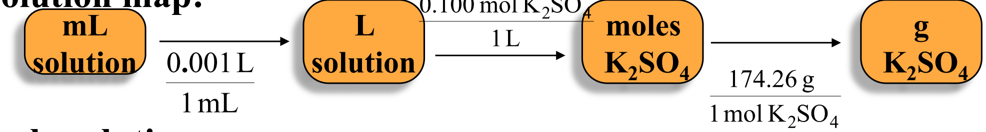
## Practice—How Would You Prepare 100.0 mL of 0.100 M $K_2SO_4$ (MM = 174.26)?, Continued

**Given:** 100.0 mL solution

**Find:** g $K_2SO_4$

**Equivalence:** 0.100 moles $K_2SO_4$ ≡ 1 L solution; 0.001 L = 1 mL; 174.26 g = 1 mol $K_2SO_4$

**Solution map:**




**Apply solution map:**

$$100.0\ \text{mL} \times \frac{0.001\ \text{L}}{1\ \text{mL}} \times \frac{0.100\ \text{mol}\ K_2SO_4}{1\ \text{L}} \times \frac{174.26\ \text{g}}{1\ \text{mol}} = 1.74\ \text{g}\ K_2SO_4$$

**Answer:**

Dissolve 1.74 g of $K_2SO_4$ in enough water to total 100.0 mL.

# Molarity and Dissociation

- When strong electrolytes dissolve, all the solute particles dissociate into ions.
- By knowing the formula of the compound and the molarity of the solution, it is easy to determine the molarity of the dissociated ions. Simply multiply the salt concentration by the number of ions.



# Molarity and Dissociation

$NaCl(aq) = Na^{+}(aq) + Cl^{-}(aq)$

1 "molecule" = 1 ion + 1 ion

100 "molecules" = 100 ions + 100 ions

1 mole "molecules" = 1 mole ions + 1 mole ions

1 M NaCl "molecules" = 1 M $Na^{+}$ ions + 1 M $Cl^{-}$ ions

0.25 M NaCl = 0.25 M $Na^{+}$ + 0.25 M $Cl^{-}$

# Molarity and Dissociation, Continued

$CaCl_2(aq) = Ca^{2+}(aq) + 2\ Cl^{-}(aq)$

1 "molecule" = 1 ion + 2 ion

100 "molecules" = 100 ions + 200 ions

1 mole "molecules" = 1 mole ions + 2 mole ions

1 M $CaCl_2$ = 1 M $Ca^{2+}$ ions + 2 M $Cl^{-}$ ions

0.25 M $CaCl_2$ = 0.25 M $Ca^{2+}$ + 0.50 M $Cl^{-}$



# Example 13.5—Determine the Molarity of the Ions in a 0.150 M $Na_3PO_4(aq)$ Solution.

| | |
|---|---|
| **Given:** | 0.150 M $Na_3PO_4(aq)$ |
| **Find:** | concentration of $Na^+$ and $PO_4^{3-}$, *M* |
| **Relationships:** | $Na_3PO_4(aq) \rightarrow 3\ Na^+(aq) + PO_4^{3-}(aq)$ |
| **Solve:** | $0.150\ \text{M}\ \cancel{Na_3PO_4} \times \frac{1\ \cancel{\text{mol}}\ PO_4^{3-}}{1\ \cancel{\text{mol}}\ \cancel{Na_3PO_4}} = 0.150\ \text{M}\ PO_4^{3-}$ <br> $0.150\ \text{M}\ \cancel{Na_3PO_4} \times \frac{3\ \cancel{\text{mol}}\ Na^+}{1\ \cancel{\text{mol}}\ \cancel{Na_3PO_4}} = 0.450\ \text{M}\ Na^+$ |
| **Check:** | The unit is correct, the magnitude seems reasonable as the ion molarities are at least as large as the $Na_3PO_4$. |

## Practice—Find the Molarity of All Ions in the Given Solutions of Strong Electrolytes.

- 0.25 M $MgBr_2(aq)$.
- 0.33 M $Na_2CO_3(aq)$.
- 0.0750 M $Fe_2(SO_4)_3(aq)$.



## Practice—Find the Molarity of All Ions in the Given Solutions of Strong Electrolytes, Continued.

- $MgBr_2(aq) \rightarrow Mg^{2+}(aq) + 2\ Br^{-}(aq)$
  0.25 M 0.25 M 0.50 M
- $Na_2CO_3(aq) \rightarrow 2\ Na^{+}(aq) + CO_3^{2-}(aq)$
  0.33 M 0.66 M 0.33 M
- $Fe_2(SO_4)_3(aq) \rightarrow 2\ Fe^{3+}(aq) + 3\ SO_4^{2-}(aq)$
  0.0750 M 0.150 M 0.225 M

# Dilution

- Dilution is adding extra solvent to decrease the concentration of a solution.
- The amount of solute stays the same, but the concentration decreases.
- **Dilution Formula:**

$Conc_{start\ soln} \times Vol_{start\ soln} = Conc_{final\ soln} \times Vol_{final\ sol}$

- Concentrations and volumes can be most units as long as they are consistent.



# Example—What Volume of 12.0 M KCl Is Needed to Make 5.00 L of 1.50 M KCl Solution?

**Given:**

| | *Initial solution* | *Final solution* |
|---|---|---|
| ***Concentration*** | 12.0 M | 1.50 M |
| ***Volume*** | ? L | 5.00 L |

**Find:** L of initial KCl

**Equation:** $(conc_1)\cdot(vol_1) = (conc_2)\cdot(vol_2)$

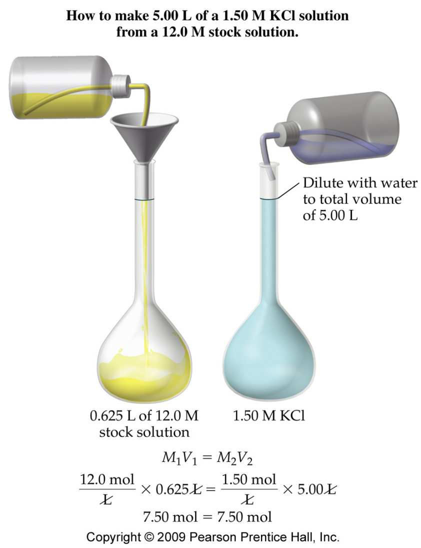
**Rearrange and apply equation:**

$$vol_1 = \frac{(conc_2)\bullet(vol_2)}{(conc_1)}$$

$$vol_1 = \frac{(1.50\ M)\bullet(5.00\ L)}{(12.0\ M)}$$

$$vol_1 = 0.625\ L$$

# Making a Solution by Dilution




$M_1 \text{ x } V_1 = M_2 \text{ x } V_2$

$M_1 = 12.0 \text{ M} \quad V_1 = ? \text{ L}$

$M_2 = 1.50 \text{ M} \quad V_2 = 5.00 \text{ L}$

$$M_1 \times V_1 = M_2 \times V_2$$

$$V_1 = \frac{M_2 \times V_2}{M_1}$$

$$V_1 = \frac{(1.50 \text{ M}) \times (5.00 \text{ L})}{(12.0 \text{ M})} = 0.625 \text{ L}$$

Dilute 0.625 L of 12.0 M solution to 5.00 L.



## Example—Dilution Problems

- What is the concentration of a solution made by diluting 15 mL of 5.0% sugar to 135 mL?

$M_1 = 5.0 \%$ $\quad M_2 = ? \%$

$V_1 = 15 \text{ mL}$ $\quad V_2 = 135 \text{ mL}$

$(5.0\%)(15 \text{ mL}) = M_2 \text{ x } (135 \text{ mL})$

$M_2 = 0.55\%$

- How would you prepare 200 mL of 0.25 M NaCl solution from a 2.0 M solution?

$M_1 = 2.0 \text{ M}$ $\quad M_2 = 0.25 \text{ M}$ $\quad (2.0 \text{ M}) \text{ x } V_1 = (0.25 \text{ M})(200 \text{ mL})$

$V_1 = ? \text{ mL}$ $\quad V_2 = 200 \text{ mL}$ $\quad V_1 = 25 \text{ mL}$

Dilute 25 mL of 2.0 M NaCl solution to 200 mL.

## Practice—Determine the Concentration of the Following Solutions.

- Made by diluting 125 mL of 0.80 M HCl to 500 mL.

- Made by adding 200 mL of water to 800 mL of 400 ppm.



## Practice—Determine the Concentration of the Following Solutions, Continued.

- Made by diluting 125 mL of 0.80 M HCl to 500 mL.

$M_1$ = 0.80 M $M_2$ = ? M
$V_1$ = 125 mL $V_2$ = 500 mL

(0.80 M)(125 mL) = $M_2$ x (500 mL)
$M_2$ = 0.20 M

- Made by adding 200 mL of water to 800 mL of 400 ppm.

$M_1$ = 400 ppm $M_2$ = ? ppm
$V_1$ = 800 mL $V_2$ = 200 + 800 mL

(400 PPM)(800 mL) = $M_2$ x (1000 mL)
$M_2$ = 320 PPM

# Example—To What Volume Should You Dilute 0.200 L of 15.0 M NaOH to Make 3.00 M NaOH?

| | | |
|---|---|---|
| • Sort information. | **Given:**<br>**Find:** | $V_1 = 0.200\text{L}, M_1 = 15.0\text{ M}, M_2 = 3.00\text{ M}$<br>$V_2$, L |
| • Strategize. | **Solution Map:**<br><br>**Relationships:** | $V_1, M_1, M_2$ → $V_2$<br>$\frac{M_1 \bullet V_1}{M_2} = V_2$<br>$M_1V_1 = M_2V_2$ |
| • Follow the solution map to **Solve** the problem. | **Solve:** | $\frac{\left(15.0\frac{\text{mol}}{\text{L}}\right) \bullet (0.200\text{ L})}{\left(3.00\frac{\text{mol}}{\text{L}}\right)} = 1.00\text{ L}$ |
| • Check. | **Check:** | Since the solution is diluted by a factor of 5, the volume should increase by a factor of 5, and it does. |



# Practice Question 1—How Would You Prepare 400 mL of a 4.0% Solution From a 12% Solution?

# Practice Question 2—How Would You Prepare 250 mL of a 3.0% Solution From a 7.5% Solution?

## Practice Question 1—How Would You Prepare 400 ML of a 4.0% Solution From a 12% Solution?, Continued

$M_1$ = 12 % $M_2$ = 4.0 % (12%) x $V_1$ = (4.0%)(400 mL)
$V_1$ = ? mL $V_2$ = 400 mL $V_1$ = 133 mL

Dilute 133 mL of 12% solution to 400 mL.

## Practice Question 2—How Would You Prepare 250 ML of a 3.0% Solution From a 7.5% Solution?, Continued

$M_1$ = 7.5 % $M_2$ = 3.0 % (7.5%) x $V_1$ = (3.0%)(250 mL)
$V_1$ = ? mL $V_2$ = 250 mL $V_1$ = 100 mL

Dilute 100 mL of 7.5% solution to 250 mL.



# Solution Stoichiometry

- We know that the balanced chemical equation tells us the relationship between moles of reactants and products in a reaction.
  - ✓ $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$ implies that for every 2 moles of $H_2$ you use, you need 1 mole of $O_2$ and will make 2 moles of $H_2O$.
- Since molarity is the relationship between moles of solute and liters of solution, we can now measure the moles of a material in a reaction in solution by knowing its molarity and volume.

## Example 13.7—How Many Liters of 0.115 M KI Is Needed to React with 0.104 L of a 0.225 M $Pb(NO_3)_2$?

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

**Given:** 0.104 L $Pb(NO_3)_2$

**Find:** L KI

**Solution Map:** L $Pb(NO_3)_2$ → mol $Pb(NO_3)_2$ → mol KI → L KI

$$\frac{0.225\text{ mol}}{1\text{ L}} \qquad \frac{2\text{ mol KI}}{1\text{ mol } Pb(NO_3)_2} \qquad \frac{1\text{ L}}{0.115\text{ mol}}$$

**Relationships:** 0.225 mol $Pb(NO_3)_2$ = 1 L; 2 mol KI = 1 mol $Pb(NO_3)_2$; 0.115 mol KI = 1 L

**Solve:**

$$0.104\,\cancel{L}\,Pb(NO_3)_2 \times \frac{0.225\,\cancel{\text{mol }Pb(NO_3)_2}}{1\,\cancel{L}} \times \frac{2\,\cancel{\text{mol KI}}}{1\,\cancel{\text{mol }Pb(NO_3)_2}} \times \frac{1\text{ L}}{0.115\,\cancel{\text{mol KI}}} = 0.407\text{ L KI}$$

**Check:** The unit is correct.



Example 13.7:

- How much 0.115 M KI solution, in liters, is required to completely precipitate all the $Pb^{2+}$ in 0.104 L of 0.225 M $Pb(NO_3)_2$?

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

Example:
How much 0.115 M KI solution, in liters, is required to completely precipitate all the $Pb^{2+}$ in 0.104 L of 0.225 M $Pb(NO_3)_2$?
2 KI(*aq*) + $Pb(NO_3)_2$(*aq*) → $PbI_2$(*s*) + 2 $KNO_3$(*aq*)

- Write down the given quantity and its units.

Given: 0.104 L $Pb(NO_3)_2$



Example:
How much 0.115 M KI solution, in liters, is required to completely precipitate all the $Pb^{2+}$ in 0.104 L of 0.225 M $Pb(NO_3)_2$?
2 KI(*aq*) + $Pb(NO_3)_2$(*aq*) → $PbI_2$(*s*) + 2 $KNO_3$(*aq*)

Information:
Given: 0.104 L $Pb(NO_3)_2$

- Write down the quantity to find and/or its units.

Find: volume of KI solution, L

Example:
How much 0.115 M KI solution, in liters, is required to completely precipitate all the $Pb^{2+}$ in 0.104 L of 0.225 M $Pb(NO_3)_2$?
$2\,KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2\,KNO_3(aq)$

Information:
Given: 0.104 L $Pb(NO_3)_2$
Find: L KI

- Collect needed conversion factors:

$$0.115\text{ M KI} \therefore 0.115\text{ mol KI} \equiv 1\text{ L solution.}$$

$$0.225\text{ M } Pb(NO_3)_2 \therefore 0.225\text{ mol } Pb(NO_3)_2 \equiv 1\text{ L solution.}$$

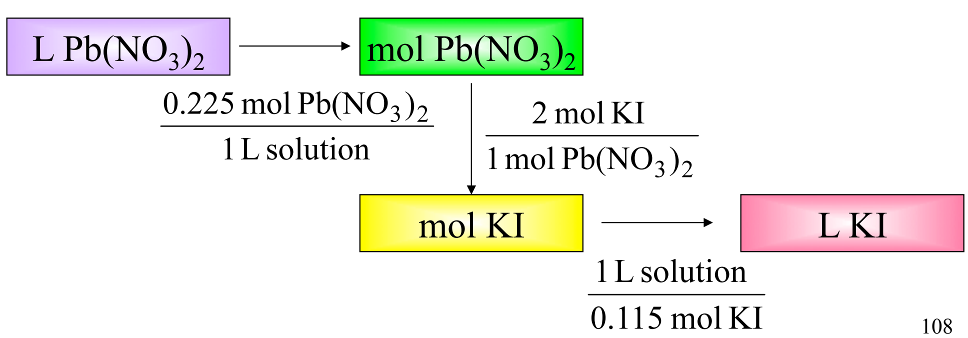
$$\text{Chemical equation} \therefore 2\text{ mol KI} \equiv 1\text{ mol } Pb(NO_3)_2.$$



Example:
How much 0.115 M KI solution, in liters, is required to completely precipitate all the $Pb^{2+}$ in 0.104 L of 0.225 M $Pb(NO_3)_2$?
$2\,KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2\,KNO_3(aq)$

Information:
Given: 0.104 L $Pb(NO_3)_2$
Find: L KI
Conversion Factors:
0.115 mol KI ≡ 1 L solution
0.225 mol $Pb(NO_3)_2$ ≡ 1 L solution
2 mol KI ≡ 1 mol $Pb(NO_3)_2$

- Design a solution map:



Example:
How much 0.115 M KI solution, in liters, is required to completely precipitate all the $Pb^{2+}$ in 0.104 L of 0.225 M $Pb(NO_3)_2$?
$2\ KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2\ KNO_3(aq)$

Information:
Given: 0.104 L $Pb(NO_3)_2$
Find: L KI
Conversion Factors:
0.115 mol KI ≡ 1 L solution
0.225 mol $Pb(NO_3)_2$ ≡ 1 L solution
2 mol KI ≡ 1 mol $Pb(NO_3)_2$
Solution Map:
L $Pb(NO_3)_2$ → mol $Pb(NO_3)_2$ → mol KI → L KI

- Apply the solution map:

$$0.104\ \text{L}\ Pb(NO_3)_2\ \text{sol'n} \times \frac{0.225\ \text{mol}\ Pb(NO_3)_2}{1\ \text{L sol'n}} \times \frac{2\ \text{mol KI}}{1\ \text{mol}\ Pb(NO_3)_2} \times \frac{1\ \text{L KI sol'n}}{0.115\ \text{mol KI}}$$

$= 0.40696$ L

$= 0.407$ L



Example:
How much 0.115 M KI solution, in liters, is required to completely precipitate all the $Pb^{2+}$ in 0.104 L of 0.225 M $Pb(NO_3)_2$?
$2\ KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2\ KNO_3(aq)$

Information:
Given: 0.104 L $Pb(NO_3)_2$
Find: L KI
Conversion Factors:
0.115 mol KI ≡ 1 L solution
0.225 mol $Pb(NO_3)_2$ ≡ 1 L solution
2 mol KI ≡ 1 mol $Pb(NO_3)_2$
Solution Map:
L $Pb(NO_3)_2$ → mol $Pb(NO_3)_2$ → mol KI → L KI

- Check the solution:

Volume of KI solution required = 0.407 L.

The units of the answer, L KI solution, are correct.
The magnitude of the answer makes sense since the molarity of $Pb(NO_3)_2$ is larger than KI and it takes 2x as many moles of KI as $Pb(NO_3)_2$, the volume of KI solution should be larger than the volume of $Pb(NO_3)_2$.

## Practice—How Many Liters of 0.0623 M $Ba(OH)_2(aq)$ Are Needed to React with 0.438 L of 0.107 M HCl?

$$Ba(OH)_2(aq) + 2\ HCl(aq) \rightarrow BaCl_2(aq) + 2\ H_2O(l)$$



## Practice—How Many Liters of 0.0623 M $Ba(OH)_2(aq)$ Are Needed to React with 0.438 L of 0.107 M HCl? $Ba(OH)_2(aq) + 2\ HCl(aq) \rightarrow BaCl_2(aq) + 2\ H_2O(l)$, Continued

**Given:** 0.0.438 L HCl

**Find:** L $Ba(OH)_2$

**Solution Map:**

**L HCl** → **mol HCl** → **mol $Ba(OH)_2$** → **L $Ba(OH)_2$**

$\frac{0.107\ \text{mol}}{1\ \text{L}}$ ; $\frac{1\ \text{mol}\ Ba(OH)_2\ \text{l}}{2\ \text{mol HCl}}$ ; $\frac{1\ \text{L}}{0.0623\ \text{mol}}$

**Relationships:** 0.0623 mol $Ba(OH)_2$ = 1 L; 2 mol HCl= 1 mol $Ba(OH)_2$; 0.107 mol HCl = 1 L

**Solve:**

$$0.438\ \text{L HCl} \times \frac{0.107\ \text{mol HCl}}{1\ \text{L}} \times \frac{1\ \text{mol}\ Ba(OH)_2}{2\ \text{mol HCl}} \times \frac{1\ \text{L}}{0.0623\ \text{mol}\ Ba(OH)_2} = 0.376\ \text{L}\ Ba(OH)_2$$

**Check:** The unit is correct.

# Why Do We Do That?

- We spread salt on icy roads and walkways to melt the ice.
- We add antifreeze to car radiators to prevent the water from boiling or freezing.
  - ✓ Antifreeze is mainly ethylene glycol.
- When we add solutes to water, it changes the freezing point and boiling point of the water.





# Colligative Properties

- The properties of the solution are different from the properties of the solvent.
- Any property of a solution whose value depends only on the number of dissolved solute particles is called a **colligative property**.
  - ✓ It does not depend on what the solute particle is.
- The freezing point, boiling point, and osmotic pressure of a solution are colligative properties.

# Solution Concentration Molality, *m*

- Moles of solute per 1 kilogram of solvent.
  - ✓Defined in terms of amount of solvent, not solution.
  - ✓Does not vary with temperature.
    - ➢Because based on masses, not volumes.

$$\text{molality} = \frac{\text{mole of solute}}{\text{kg of solvent}}$$

Mass of solution = volume of solution x density.

Mass of solution = mass of solute + mass of solvent.



# Example 13.8—What Is the Molality of a Solution Prepared by Mixing 17.2 g of $C_2H_6O_2$ with 0.500 kg of $H_2O$?

| | |
|---|---|
| **Given:** | 17.2 g $C_2H_6O_2$, 0.500 kg $H_2O$ |
| **Find:** | *m* |
| **Concept Plan:** | **g $C_2H_6O_2$** → **mol $C_2H_6O_2$**; **mol $C_2H_6O_2$** + **kg $H_2O$** → ***m*** ($m = \frac{\text{mol}}{\text{kg}}$) |
| **Relationships:** | *m* = mol/kg, 1 mol $C_2H_6O_2$ = 62.07 g |
| **Solve:** | $17.2 \text{ g } C_2H_6O_2 \times \frac{1 \text{ mol } C_2H_6O_2}{62.07 \text{ g } C_2H_6O_2} = 0.2771 \text{ mol}$ <br> $m = \frac{0.2771 \text{ mol } C_2H_6O_2}{0.500 \text{ kg } H_2O}$ <br> $m = 0.554 \text{ M}$ |
| **Check:** | The unit is correct, the magnitude is reasonable. |

Example 13.8:

- Calculate the molality of a solution containing 17.2 g of ethylene glycol ($C_2H_6O_2$) dissolved in 0.500 kg of water.



Example:
Calculate the molality of a solution containing 17.2 g of ethylene glycol ($C_2H_6O_2$) dissolved in 0.500 kg of water.

- Write down the given quantity and its units.

Given: 17.2 g $C_2H_6O_2$

0.500 kg $H_2O$

Example:
Calculate the molality of a solution containing 17.2 g of ethylene glycol ($C_2H_6O_2$) dissolved in 0.500 kg of water.

Information:
Given: 17.2 g $C_2H_6O_2$; 0.500 kg $H_2O$

- Write down the quantity to find and/or its units.

Find: molality (*m*)



Example:
Calculate the molality of a solution containing 17.2 g of ethylene glycol ($C_2H_6O_2$) dissolved in 0.500 kg of water.

Information:
Given: 17.2 g $C_2H_6O_2$; 0.500 kg $H_2O$
Find: molality, *m*

- Collect needed equations and conversion factors:

$$\text{molality} = \frac{\text{moles solute}}{\text{kg solvent}}$$

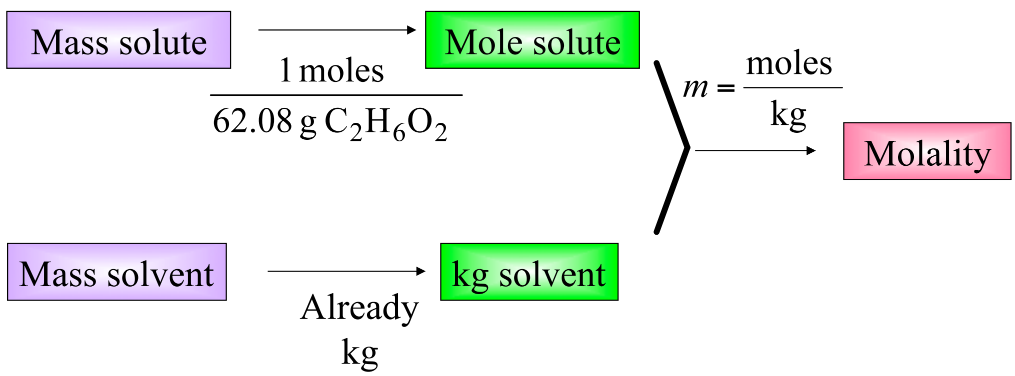
Molar Mass $C_2H_6O_2$ = 62.08 g/mol ∴ 62.08 g $C_2H_6O_2$ = 1 mol.

Example:
Calculate the molality of a solution containing 17.2 g of ethylene glycol ($C_2H_6O_2$) dissolved in 0.500 kg of water.

Information:
Given: 17.2 g $C_2H_6O_2$; 0.500 kg $H_2O$
Find: molality, *m*
Conversion Factors:
62.08 g $C_2H_6O_2$ = 1 mol

$$\text{Molality} = \frac{\text{moles solute}}{\text{kg solvent}}$$

- Design a solution map:






Example:
Calculate the molality of a solution containing 17.2 g of ethylene glycol ($C_2H_6O_2$) dissolved in 0.500 kg of water.

Information:
Given: 17.2 g $C_2H_6O_2$; 0.500 kg $H_2O$
Find: molality, *m*
Conversion Factors:
62.08 g $C_2H_6O_2$ = 1 mol

$$\text{Molality} = \frac{\text{moles solute}}{\text{kg solvent}}$$

- Apply the solution map:

$$17.2\,\cancel{\text{g}}\,C_2H_6O_2 \times \frac{1\text{ mol } C_2H_6O_2}{62.08\,\cancel{\text{g}}} = 0.27\underline{7}1\text{ mol } C_2H_6O_2$$

$$\text{Molality} = \frac{\text{moles solute}}{\text{kg solvent}} = \frac{0.27\underline{7}1\text{ mol } C_2H_6O_2}{0.500\text{ kg } H_2O}$$

$$= 0.554\ m\ C_2H_6O_2$$

| Example: Calculate the molality of a solution containing 17.2 g of ethylene glycol ($C_2H_6O_2$) dissolved in 0.500 kg of water. | Information:<br>Given: 17.2 g $C_2H_6O_2$; 0.500 kg $H_2O$<br>Find: molality, *m*<br>Conversion Factors:<br>62.08 g $C_2H_6O_2$ = 1 mol<br>$\text{Molality} = \frac{\text{moles solute}}{\text{kg solvent}}$ |
|---|---|

- Check the solution:

Molality of solution = 0.554 *m*.

The units of the answer, m, are correct.
The magnitude of the answer makes sense since the mass of solute is less than the ½ mole and the mass of solvent is ½ kg.



## Practice—What Is the Molality of a Solution that Is Made by Dissolving 3.4 g of $NH_3$ (MM 17.03) in 1500 mL of $H_2O$ ($d$ =1.00 g/mL).

## Practice—What Is the Molality of a Solution that Is Made by Dissolving 3.4 g of $NH_3$ (MM 17.03) in 1500 mL of $H_2O$ ($d$ =1.00 g/mL), Continued.

| | |
|---|---|
| **Given:** | 3.4 g $NH_3$, 1500 mL $H_2O$ |
| **Find:** | $m$ |
| **Solution Map:** | **g $NH_3$** → **mol $NH_3$**; **mL $H_2O$** → **g $H_2O$** → **kg $H_2O$**; → ***m***, $m = \frac{\text{mol}}{\text{kg}}$ |
| **Relationships:** | $m$ = mol/kg, 1 mol $NH_3$ = 17.03 g, 1 kg = 1000g, 1.00 g = 1 mL |
| **Solve:** | $3.4\ \text{g NH}_3 \times \frac{1\ \text{mol NH}_3}{17.03\ \text{g NH}_3} = 0.20\ \text{mol}$ <br> $1500\ \text{mL} \times \frac{1.00\ \text{g H}_2\text{O}}{1\ \text{mL}} \times \frac{1\ \text{kg}}{1000\ \text{g}} = 1.5\ \text{kg}$ <br> $m = \frac{0.20\ \text{mol NH}_3}{1.5\ \text{kg}}$ <br> $m = 0.13\ m$ |
| **Check:** | The unit is correct, the magnitude is reasonable. |



# Freezing Points of Solutions

- The freezing point of a solution is always lower than the freezing point of a pure solvent.
  - ✓ **Freezing point depression.**
- The difference between the freezing points of the solution and pure solvent is directly proportional to the **molal** concentration.
- $\Delta T_f = m \times K_f$
  - ✓ $K_f$ = freezing point constant.
- Used to determine molar mass of compounds.

# Freezing and Boiling Point Constants

| Solvent | $K_f$ °C/$m$ | FP °C | $K_b$ °C/$m$ | BP °C |
|---|---|---|---|---|
| Water, $H_2O$ | 1.86 | 0.00 | 0.512 | 100.0 |
| Benzene, $C_6H_6$ | 5.12 | 5.53 | 2.53 | 80.1 |
| Cyclohexane,$C_6H_{12}$ | 20.0 | 6.47 | 2.79 | 80.7 |
| Naphthalene, $C_{10}H_8$ | 6.9 | 80.2 | 5.65 | 218 |
| Ethanol, $C_2H_5OH$ | 1.99 | -115 | 1.22 | 78.4 |
| t-butanol, $(CH_3)_3COH$ | 8.3 | 25.6 | | 82.4 |
| Carbon tetrachloride,$CCl_4$ | 29.8 | -22.3 | 5.02 | 76.8 |
| Methanol, $CH_3OH$ | | -97.8 | 0.80 | 64.7 |
| Acetic acid, $HC_2H_3O_2$ | 3.9 | 16.7 | 3.07 | 118 |



## Example 13.9—What Is the Freezing Point of a 1.7 *m* Aqueous Ethylene Glycol Solution, $C_2H_6O_2$?

**Given:** 1.7 $m$ $C_2H_6O_2(aq)$

**Find:** $T_f$, °C

**Solution Map:** ***m*** → ***ΔT<sub></sub>*** → ***FP***

$\Delta T_f = m \bullet K_f$ $\quad FP_{solv} - FP_{sol'n} = \Delta T$

**Relationships:** $\Delta T_f = m \cdot K_f$, $K_f$ for $H_2O$ = 1.86 °C/$m$, $FP_{H2O}$ = 0.00 °C

**Solve:**

$$\Delta T_f = m \bullet K_{f,H_2O}$$
$$= (1.7\ m)\left(1.86 \frac{°C}{m}\right)$$
$$\Delta T_f = 3.2\ °C$$

$$FP_{H_2O} - FP_{sol'n} = \Delta T_f$$
$$0.00°C - FP_{sol'n} = 3.2°C$$
$$FP_{sol'n} = -3.2°C$$

**Check:** The unit is correct, the freezing point being lower than the normal freezing point makes sense.

Example 13.9:

- Calculate the freezing point of a 1.7 *m* ethylene glycol solution.



Example:
Calculate the freezing point of a 1.7 *m* ethylene glycol solution.

- Write down the given quantity and its units.

Given: 1.7 *m* $C_2H_6O_2$

Example:
Calculate the freezing point of a 1.7 *m* ethylene glycol solution.

Information:
Given: 1.7 *m* $C_2H_6O_2$ in $H_2O$

- Write down the quantity to find and/or its units.

Find: freezing point (°C)



Example:
Calculate the freezing point of a 1.7 *m* ethylene glycol solution.

Information:
Given: 1.7 *m* $C_2H_6O_2$ in $H_2O$
Find: *FP* (°C)

- Collect needed equations:

$$\Delta T_f = m \times K_f$$

$$FP_{solution} = FP_{solvent} - \Delta T_f$$

Example:
Calculate the freezing point of a 1.7 *m* ethylene glycol solution.

Information:
Given: 1.7 *m* $C_2H_6O_2$ in $H_2O$
Find: *FP* (°C)
Equations:

$$\Delta T_f = m \cdot K_f;$$
$$FP_{\text{solution}} = FP_{\text{solvent}} - \Delta T_f$$

- Design a solution map:

Molality → $\Delta T_f$ → $FP_{\text{solution}}$

$$\Delta T_f = m \cdot K_f \qquad FP_{\text{solution}} = FP_{\text{solvent}} - \Delta T_f$$



Example:
Calculate the freezing point of a 1.7 *m* ethylene glycol solution.

Information:
Given: 1.7 *m* $C_2H_6O_2$ in $H_2O$
Find: *FP* (°C)
Equation: $\Delta T_f = m \cdot K_f;$

$$FP_{\text{solution}} = FP_{\text{solvent}} - \Delta T_f$$

Solution Map: $m \rightarrow \Delta T_f \rightarrow FP_{\text{solution}}$

- Apply the solution map:

$$\Delta T_f = m \times K_f$$
$$\Delta T_f = 1.7 \frac{\text{mol solute}}{\text{kg solvent}} \times 1.86 \frac{{}^\circ\text{C kg solvent}}{\text{mol solute}}$$
$$\Delta T_f = 3.2\ {}^\circ\text{C}$$
$$FP_{\text{solution}} = FP_{\text{solvent}} - \Delta T_f$$
$$FP_{\text{solution}} = 0.00\ {}^\circ\text{C} - 3.2\ {}^\circ\text{C}$$
$$FP_{\text{solution}} = -3.2\ {}^\circ\text{C}$$

| Example: Calculate the freezing point of a 1.7 $m$ ethylene glycol solution. | Information:<br>Given: 1.7 $m$ $C_2H_6O_2$ in $H_2O$<br>Find: $FP$ (°C)<br>Equation: $\Delta T_f = m \cdot K_f$;<br>$FP_{solution} = FP_{solvent} - \Delta T_f$<br>Solution Map: $m \rightarrow \Delta T_f \rightarrow FP_{solution}$ |
|---|---|

- Check the solution:

Freezing point of solution = −3.2 °C.

The units of the answer, °C, are correct. The magnitude of the answer makes sense since the freezing point of the solution is less than the freezing point of $H_2O$.



## Practice—What Is the Freezing Point of a Solution that Has 0.20 moles of Sulfur Dissolved in 0.10 kg of Cyclohexane?

($FP_{cyclohexane}$ = 6.5 °C, $K_f$ = 20.0 °C/$m$)

## Practice—What Is the Freezing Point of a Solution that Has 0.20 moles of Sulfur Dissolved in 0.10 kg of Cyclohexane?, Continued

| | |
|---|---|
| **Given:** | 0.20 mol S, 0.10 kg cyclohexane |
| **Find:** | $T_f$, °C |
| **Solution Map:** | **mol S, kg solvent** → ***m*** → $\boldsymbol{\Delta T_f}$ → ***FP***<br>$m = \frac{\text{mol}}{\text{kg}}$ ; $\Delta T_f = m \bullet K_f$ ; $FP_{solv} - FP_{sol'n} = \Delta T$ |
| **Relationships:** | $\Delta T_f = m \cdot K_f$, $K_f = 20.0$ °C/$m$, FP = 6.5 °C, $m$ = mol/kg |
| **Solve:** | $m = \frac{0.20\text{ mol S}}{0.10\text{ kg}} = 2.0\ m$<br>$\Delta T_f = m \bullet K_f = (2.0\ m)\left(20.0\ \frac{°C}{m}\right)$<br>$\Delta T_f = 4\underline{0}$ °C<br>$FP_{cyclohexane} - FP_{sol'n} = \Delta T_f$<br>$6.5°C - FP_{sol'n} = 4\underline{0}°C$<br>$FP_{sol'n} = -33°C$ |
| **Check:** | The unit is correct, the freezing point being lower than the normal freezing point makes sense. |



# Boiling Points of Solutions

- The boiling point of a solution is always higher than the boiling point of a pure solvent.
  - ✓**Boiling point elevation.**
- The difference between the boiling points of the solution and pure solvent is directly proportional to the **molal** concentration.
- $\boldsymbol{\Delta T_b = m \times K_b}$
  - ✓$K_b$ = boiling point constant.

## Example 13.10—What Is the Boiling Point of a 1.7-*m* Aqueous Ethylene Glycol Solution, $C_2H_6O_2$?

| | |
|---|---|
| **Given:** | 1.7 *m* $C_2H_6O_2(aq)$ |
| **Find:** | $T_b$, °C |
| **Solution Map:** | **m** → **$\Delta T_b$** → **BP**<br>$\Delta T_b = m \bullet K_b$    $BP_{sol'n} - BP_{solv} = \Delta T$ |
| **Relationships:** | $\Delta T_b = m \cdot K_b$, $K_b$ $H_2O$ = 0.512 °C/*m*, $BP_{H2O}$ = 100.00 °C |
| **Solve:** | $\Delta T_b = m \bullet K_{b,H_2O}$<br>$= (1.7\ m)\left(0.512 \frac{°C}{m}\right)$<br>$\Delta T_f = 0.87\ °C$<br>$BP_{solution} - BP_{solvent} = \Delta T_b$<br>$BP_{solution} - 100.00°C = 0.87°C$<br>$BP_{sol'n} = 100.87°C$ |
| **Check:** | The unit is correct, the boiling point being higher than the normal boiling point makes sense. |



Example 13.10:

- Calculate the boiling point of a 1.7 *m* ethylene glycol solution.

Example:
Calculate the boiling point of a 1.7-*m* ethylene glycol solution.

- Write down the given quantity and its units.

Given: 1.7 *m* $C_2H_6O_2$



Example:
Calculate the boiling point of a 1.7 *m* ethylene glycol solution.

Information:
Given: 1.7 *m* $C_2H_6O_2$ in $H_2O$

- Write down the quantity to find and/or its units.

Find: boiling point (°C)

Example:
Calculate the boiling point of a 1.7 *m* ethylene glycol solution.

Information:
Given: 1.7 *m* $C_2H_6O_2$ in $H_2O$
Find: *BP* (°C)

- Collect needed equations:

$$\Delta T_b = m \times K_b$$

$$BP_{\text{solution}} = BP_{\text{solvent}} + \Delta T_b$$



Example:
Calculate the boiling point of a 1.7 *m* ethylene glycol solution.

Information:
Given: 1.7 *m* $C_2H_6O_2$ in $H_2O$
Find: *BP* (°C)
Equations: $\Delta T_b = m \cdot K_b$;
$BP_{\text{solution}} = BP_{\text{solvent}} + \Delta T_b$

- Design a solution map:

Molality → $\Delta T_b$ → $BP_{\text{solution}}$

$\Delta T_b = m \cdot K_b$ $\quad$ $BP_{\text{solution}} = BP_{\text{solvent}} + \Delta T_b$

| Example: Calculate the boiling point of a 1.7 *m* ethylene glycol solution. | Information:<br>Given: 1.7 m $C_2H_6O_2$ in $H_2O$<br>Find: *BP* (°C)<br>Equation: $\Delta T_b = m \cdot K_b$; $BP_{\text{solution}} = BP_{\text{solvent}} + \Delta T_b$<br>Solution Map: $m \rightarrow \Delta T_b \rightarrow BP_{\text{solution}}$ |
|---|---|

- Apply the solution map:

$$\Delta T_b = m \times K_b$$

$$\Delta T_b = 1.7 \frac{\text{mol } \cancel{\text{solute}}}{\text{kg } \cancel{\text{solvent}}} \times 0.512 \frac{°\text{C } \cancel{\text{kg solvent}}}{\cancel{\text{mol}} \text{ solute}}$$

$$\Delta T_f = 0.87\ °\text{C}$$

$$BP_{\text{solution}} = BP_{\text{solvent}} + \Delta T_b$$

$$BP_{\text{solution}} = 100.00\ °\text{C} + 0.87\ °\text{C}$$

$$BP_{\text{solution}} = 100.87\ °\text{C}$$



| Example: Calculate the boiling point of a 1.7 *m* ethylene glycol solution. | Information:<br>Given: 1.7 m $C_2H_6O_2$ in $H_2O$<br>Find: *BP* (°C)<br>Equation: $\Delta T_b = m \cdot K_b$; $BP_{\text{solution}} = BP_{\text{solvent}} + \Delta T_b$<br>Solution Map: $m \rightarrow \Delta T_b \rightarrow BP_{\text{solution}}$ |
|---|---|

- Check the solution:

Boiling point of solution = 100.87 °C.

The units of the answer, °C, are correct.
The magnitude of the answer makes sense since the boiling point of the solution is higher than the boiling point of $H_2O$.

## Practice—What Is the Boiling Point of a Solution that Has 0.20 moles of Sulfur Dissolved in 0.10 kg of Cyclohexane?

($BP_{cyclohexane}$ = 80.7 °C, $K_b$= 2.79 °C/$m$)



## Practice—What Is the Boiling Point of a Solution that Has 0.20 moles of Sulfur Dissolved in 0.10 kg of Cyclohexane?, Continued

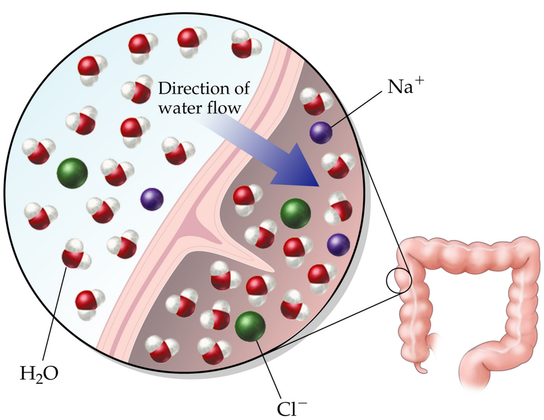
| | |
|---|---|
| **Given:** | 0.20 mol S, 0.10 kg cyclohexane |
| **Find:** | $T_b$, °C |
| **Solution Map:** | **mol S, kg solvent** → ***m*** → **$\Delta T_b$** → ***BP***<br>$m = \frac{\text{mol}}{\text{kg}}$ $\Delta T_b = m \bullet K_b$ $BP_{sol'n} - BP_{solv} = \Delta T$ |
| **Relationships:** | $\Delta T_b = m \cdot K_b$, $K_b$= 2.79 °C/$m$, $BP$ =80.7 °C, $m$ = mol/kg |
| **Solve:** | $m = \frac{0.20\text{ mol S}}{0.10\text{ kg}} = 2.0\ m$<br>$\Delta T_b = m \bullet K_b = (2.0\ m)\left(2.79\ \frac{°C}{m}\right)$<br>$\Delta T_b = 5.\underline{5}8\ °C$<br>$BP_{solution} - BP_{cyclohexane} = \Delta T_b$<br>$BP_{solution} - 80.7°C = 5.\underline{5}8°C$<br>$BP_{solution} = 86.3°C$ |
| **Check:** | The unit is correct, the boiling point being higher than the normal boiling point makes sense. |

# Osmosis and Osmotic Pressure

- Osmosis is the process in which solvent molecules pass through a semipermeable membrane that does not allow solute particles to pass.
  - ✓ Solvent flows to try to equalize concentration of solute on both sides.
  - ✓ Solvent flows from side of low concentration to high concentration.
- Osmotic pressure is pressure that is needed to prevent osmotic flow of solvent.
- Isotonic, hypotonic, and hypertonic solutions.
  - ✓ Hemolysis.



# Drinking Seawater

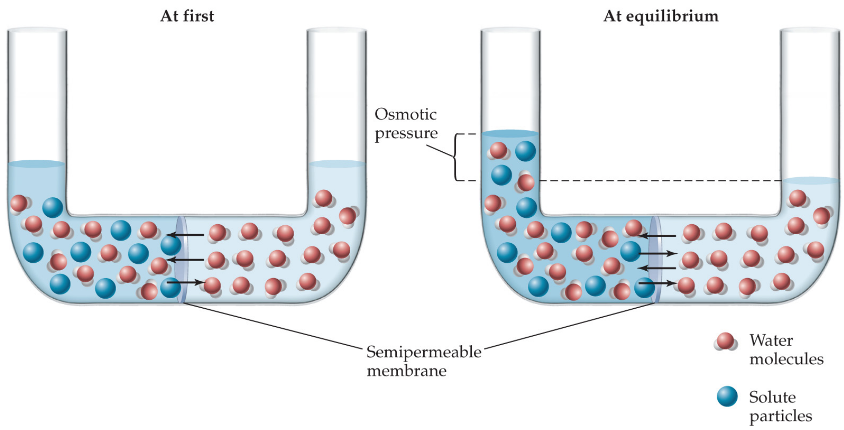





Because seawater has a higher salt concentration than your cells, water flows out of your cells into the seawater to try to decrease its salt concentration.

The net result is that, instead of quenching your thirst, you become dehydrated.

# Osmotic Pressure

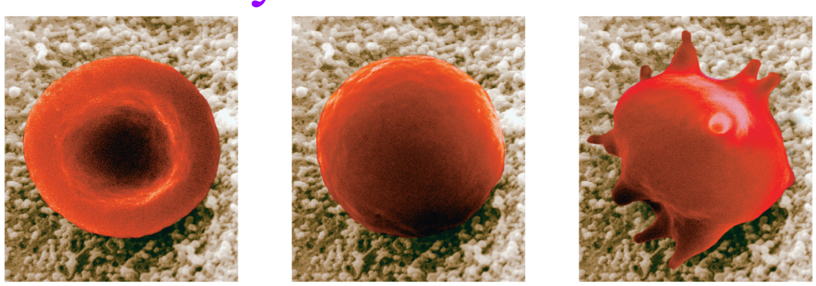





Solvent flows through a semipermeable membrane to make the solution concentration equal on both sides of the membrane. The pressure required to stop this process is **osmotic pressure**.



# Hemolysis and Crenation



(a) (b) (c)



Normal red blood cell in an isotonic Solution.

Red blood cell in a hypotonic solution. Water flows into the cell, eventually causing the cell to burst.

Red blood cell in hypertonic solution. Water flows out of the cell, eventually causing the cell to distort and shrink.