Chemistry, The Central Science, 10th edition,
AP version
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Chapter 13 (part I of II)Properties of Solutions

(N.B. aspects of this topic were seen in chapter 4)

(This ppt is a modified file from our Textbook publisher with additional slides taken from ppt file found at http://www.chemistrygeek.com/chem2.htm)

Resources and Activities

- Textbook - chapter 13 & ppt file

(AP, SAT II and regents exams)

- Online practice quiz
- Lab activities
- POGIL activities:
  - Molarity
  - Saturated and Unsaturated Solutions
  - Solubility
  - Common Ion Effect on Solubility
  - Fractional Precipitation
  - Solution Stoichiometry
- Chem guy video-lectures at

http://www.cosmolearning.com/courses/ap-chemistry-with-chemguy/video-lectures/

Chemtour videos from W.W. Norton chapter 11 :

http://www.wwnorton.com/college/chemistry/chemistry3/ch/11/chemtours.aspx

Chapter 12 Animations from glencoe website for Chang’s book:

http://glencoe.mcgraw-hill.com/sites/0023654666/student_view0/chapter12/animations_center.html#

Activities and Problem set for chapter 13 (due date________)

TextBook ch. 13 – content required for regents (in part), SAT II and AP exams

Lab activities:

- Solubility of KNO_3
- Colligative properties lab

POGILS (6 – Some of these were done with Chapter 4 work)

- Molarity
- Saturated and Unsaturated Solutions
- Solubility
- Common Ion Effect on Solubility
- Fractional Precipitation
- Solution Stoichiometry

Chapter 13 – sample and practice exercises, GIST, VC problems, selection of end of chapter exercises.

Independent work - students to view animations & interactive activities (from Norton and from the Glencoe site for Chang's book) and write summary notes on each. These summaries are to be included in your portfolio.

Animation to view

- http://www.wwnorton.com/college/chemistry/chemistry3/ch/11/chemtours.aspx

http://glencoe.mcgraw-hill.com/sites/0023654666/student_view0/chapter12/animations_center.html#

(dissolution of an ionic and a covalent compound; osmosis)

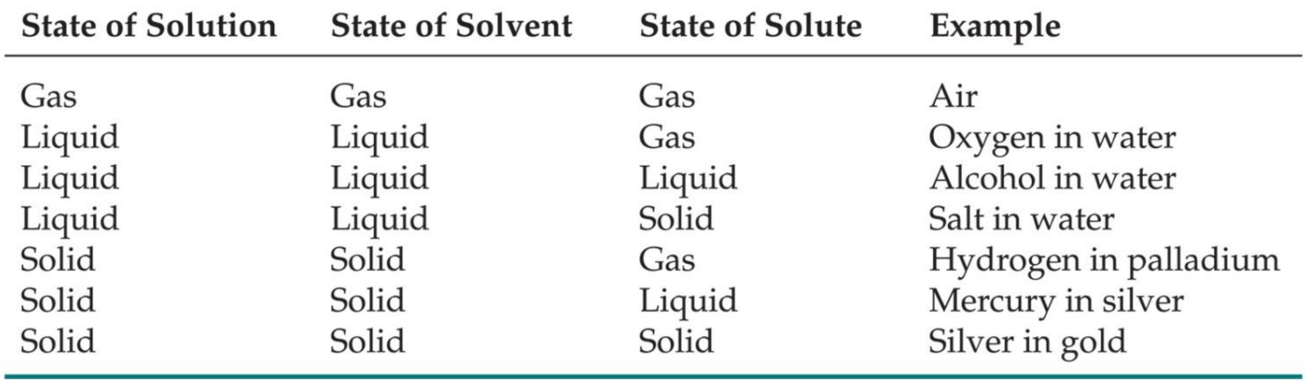
Student, Beware!

Just because a substance disappears when it comes in contact with a solvent, it doesn't mean the substance dissolved.

- Dissolution is a physical change—you can get back the original solute by evaporating the solvent.
- If you can't, the substance didn't dissolve, it reacted.

Solutions

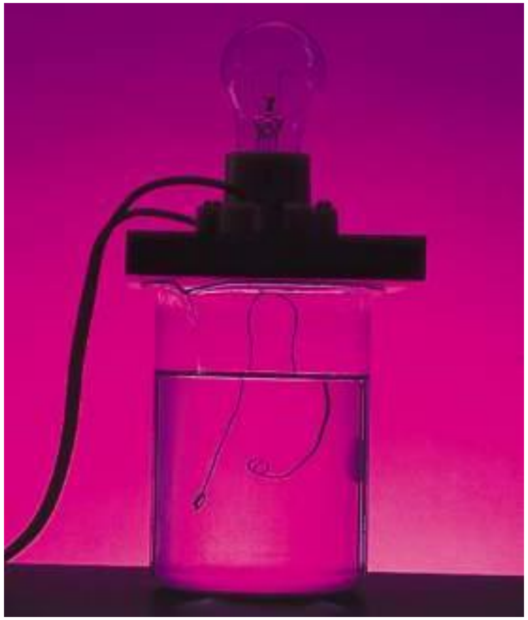
- Solutions are homogeneous mixtures of two or more pure substances.
- In a solution, the solute is dispersed uniformly throughout the solvent. (view Glencoe animation)



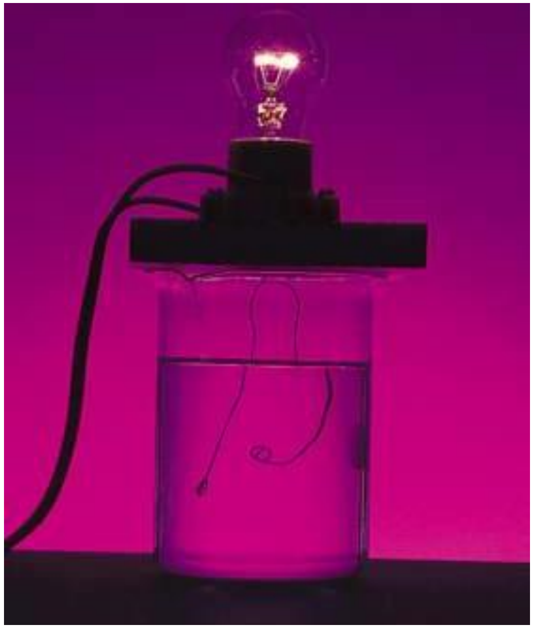
State of Solution	State of Solvent	State of Solute	Example
Gas	Gas	Gas	Air
Liquid	Liquid	Gas	Oxygen in water
Liquid	Liquid	Liquid	Alcohol in water
Liquid	Liquid	Solid	Salt in water
Solid	Solid	Gas	Hydrogen in palladium
Solid	Solid	Liquid	Mercury in silver
Solid	Solid	Solid	Silver in gold

An ***electrolyte*** is a substance that, when dissolved in water, results in a solution that can conduct electricity.

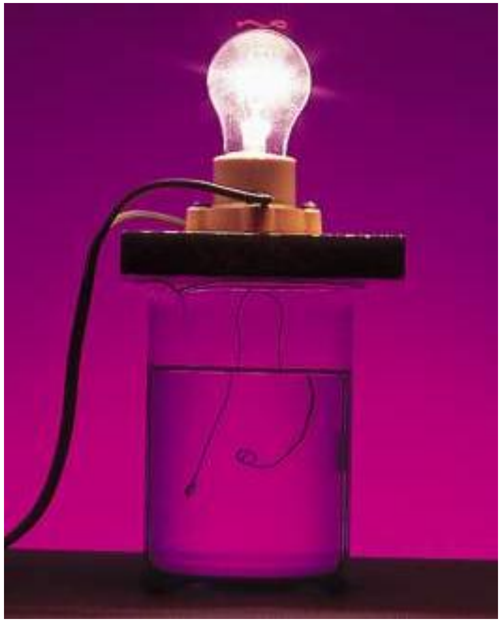
A ***nonelectrolyte*** is a substance that, when dissolved, results in a solution that does not conduct electricity.



nonelectrolyte



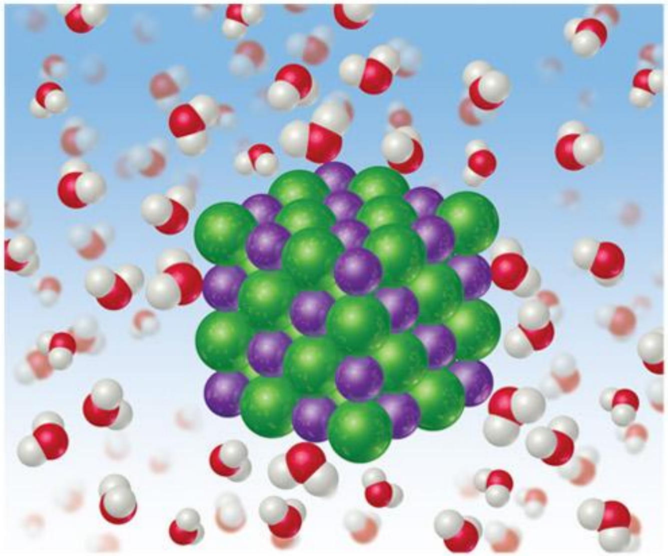
weak electrolyte



strong electrolyte

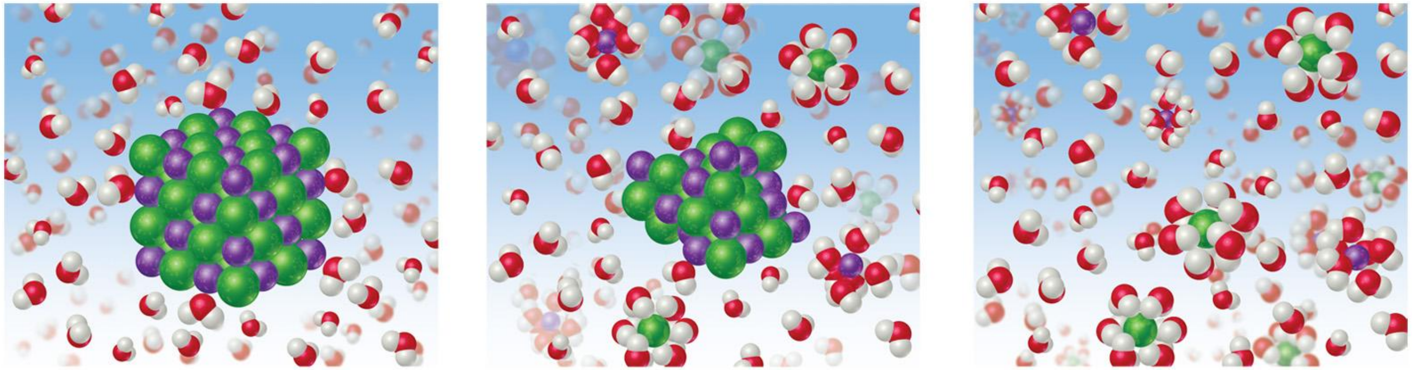
Solutions

The intermolecular forces between solute and solvent particles must be strong enough to compete with those between solute particles and those between solvent particles.



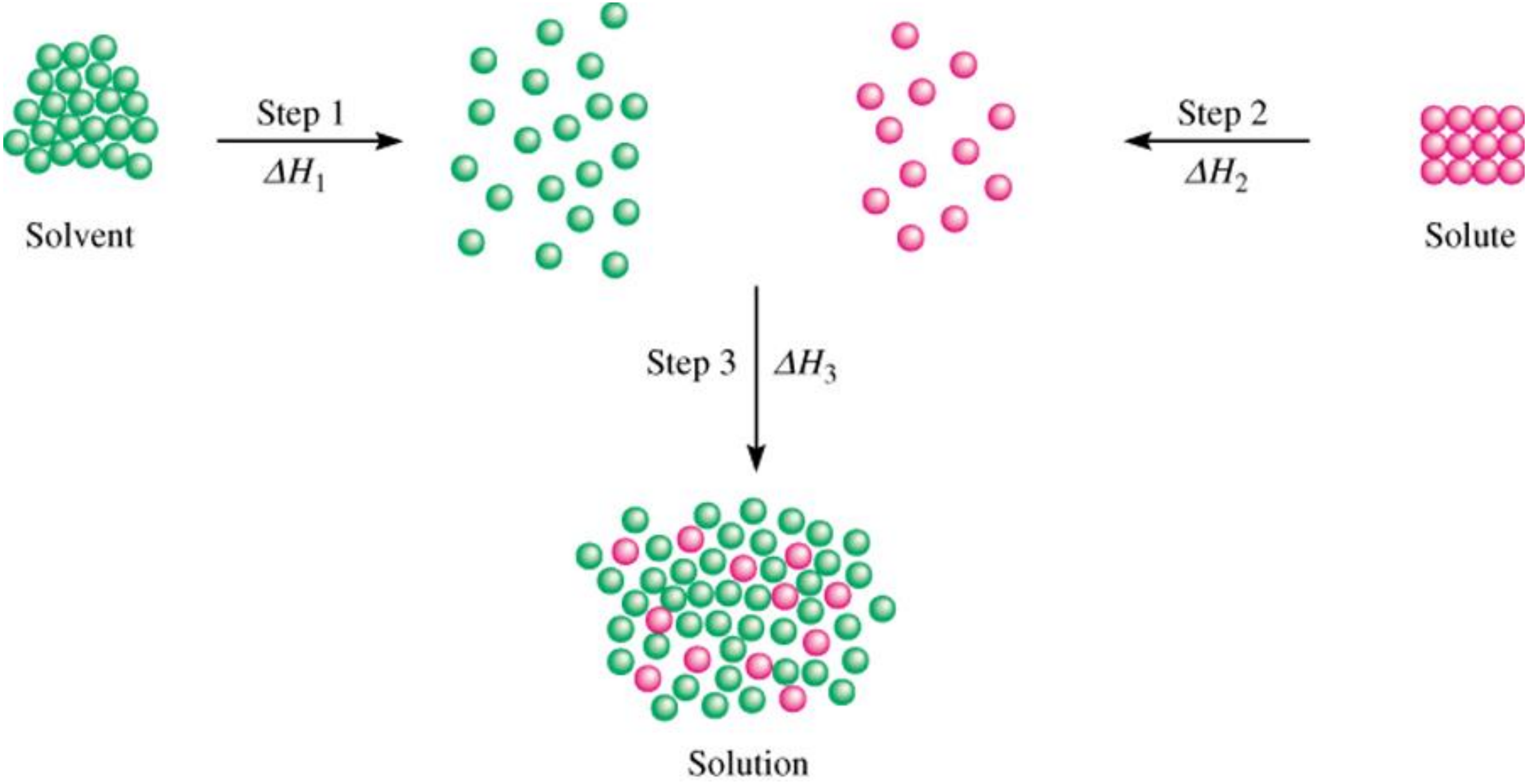
How Does a Solution Form?

As a solution forms, the solvent pulls solute particles apart and surrounds, or solvates, them.



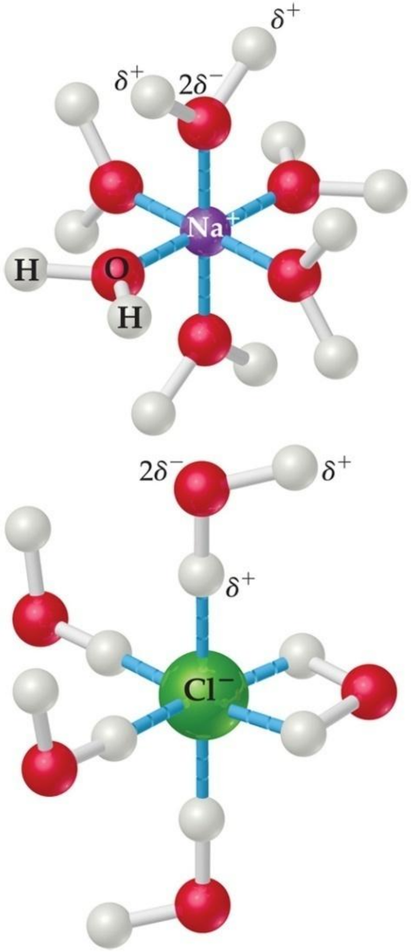
Three types of interactions in the solution process:

- solvent-solvent interaction
- solute-solute interaction
- solvent-solute interaction



$$\Delta H_{soln} = \Delta H_1 + \Delta H_2 + \Delta H_3$$

How Does a Solution Form

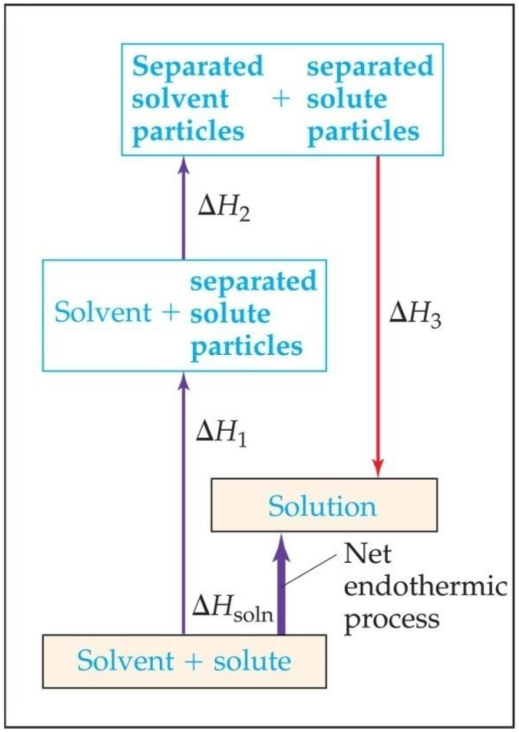


If an ionic salt is soluble in water, it is because the ion-dipole interactions are strong enough to overcome the lattice energy of the salt crystal.

Why Do Endothermic Processes Occur?

Things do not tend to occur spontaneously (i.e., without outside intervention) unless the energy of the system is lowered.

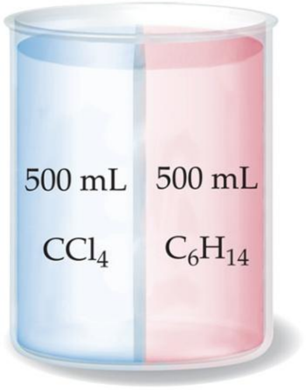
Yet we know that in some processes, like the dissolution of NH_4NO_3 in water, heat is absorbed, not released.



Enthalpy Is Only Part of the Picture

The reason is that increasing the disorder or randomness (known as entropy) of a system tends to lower the energy of the system.

So even though enthalpy may increase, the overall energy of the system can still decrease if the system becomes more disordered.



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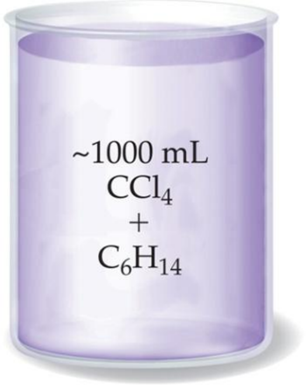


(b)

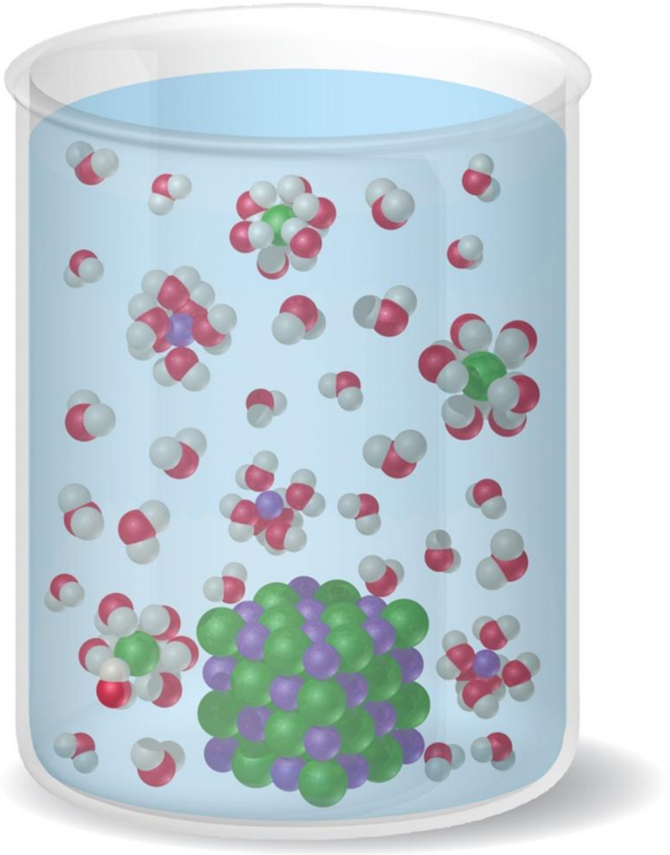
A ***saturated solution*** contains the maximum amount of a solute that will dissolve in a given solvent at a specific temperature.

An ***unsaturated solution*** contains less solute than the solvent has the capacity to dissolve at a specific temperature.

A ***supersaturated solution*** contains more solute than is present in a saturated solution at a specific temperature.

Sodium acetate crystals rapidly form when a seed crystal is added to a supersaturated solution of sodium acetate.

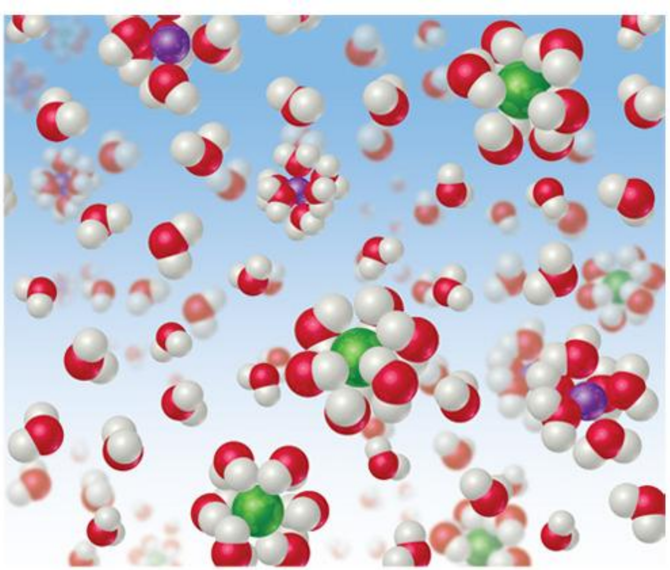
Types of Solutions



- Saturated
  - Solvent holds as much solute as is possible at that temperature.
  - Dissolved solute is in dynamic equilibrium with solid solute particles.

Types of Solutions

- Unsaturated
  - Less than the maximum amount of solute for that temperature is dissolved in the solvent.



Types of Solutions

- Supersaturated
  - Solvent holds more solute than is normally possible at that temperature.
  - These solutions are unstable; crystallization can usually be stimulated by adding a “seed crystal” or scratching the side of the flask.

Factors Affecting Solubility

- Chemists use the axiom “like dissolves like”:
  - Polar substances tend to dissolve in polar solvents.
  - Nonpolar substances tend to dissolve in nonpolar solvents.

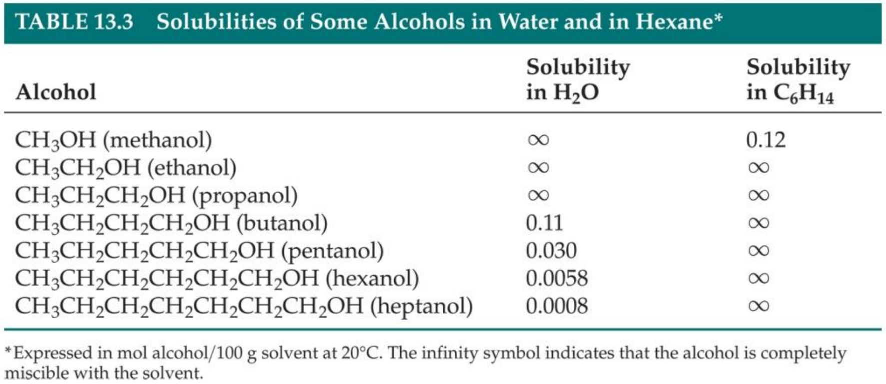
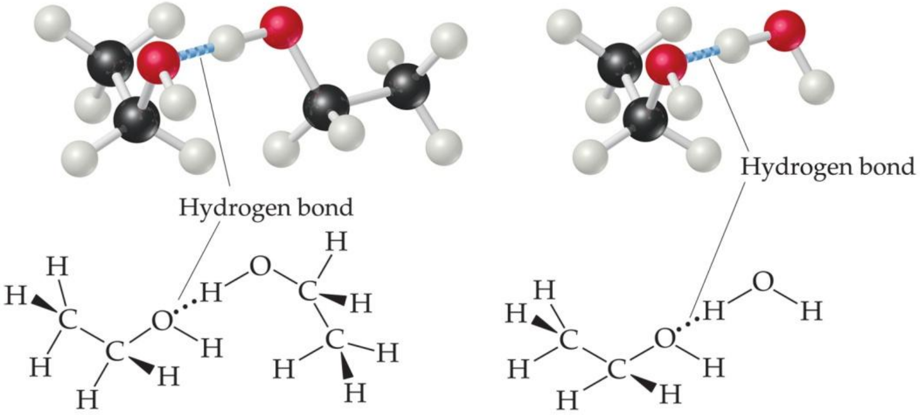


TABLE 13.3 Solubilities of Some Alcohols in Water and in Hexane*

Alcohol	Solubility in H_2O	Solubility in C_6H_{14}
CH_3OH (methanol)	∞	0.12
CH_3CH_2OH (ethanol)	∞	∞
$CH_3CH_2CH_2OH$ (propanol)	∞	∞
$CH_3CH_2CH_2CH_2OH$ (butanol)	0.11	∞
$CH_3CH_2CH_2CH_2CH_2OH$ (pentanol)	0.030	∞
$CH_3CH_2CH_2CH_2CH_2CH_2OH$ (hexanol)	0.0058	∞
$CH_3CH_2CH_2CH_2CH_2CH_2CH_2OH$ (heptanol)	0.0008	∞

*Expressed in mol alcohol/100 g solvent at 20°C. The infinity symbol indicates that the alcohol is completely miscible with the solvent.

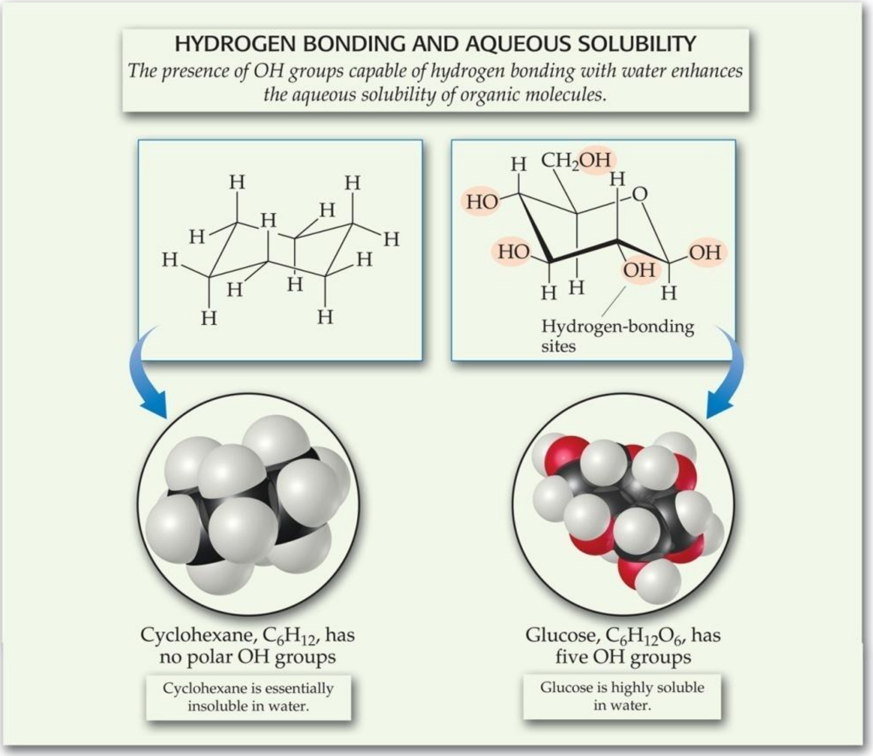
Factors Affecting Solubility



The more similar the intermolecular attractions, the more likely one substance is to be soluble in another.

Factors Affecting Solubility

Glucose (which has hydrogen bonding) is very soluble in water, while cyclohexane (which only has dispersion forces) is not.



HYDROGEN BONDING AND AQUEOUS SOLUBILITY

The presence of OH groups capable of hydrogen bonding with water enhances the aqueous solubility of organic molecules.

Cyclohexane, C_6H_{12}, has no polar OH groups

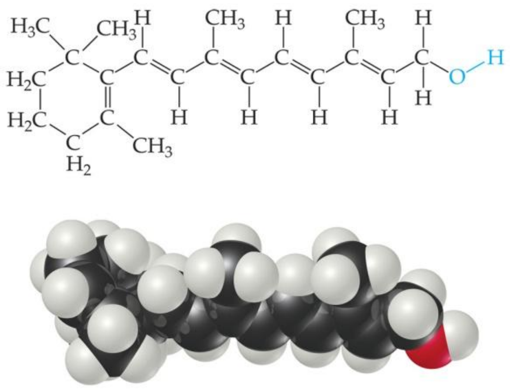
Cyclohexane is essentially insoluble in water.

Glucose, $C_6H_{12}O_6$, has five OH groups

Glucose is highly soluble in water.

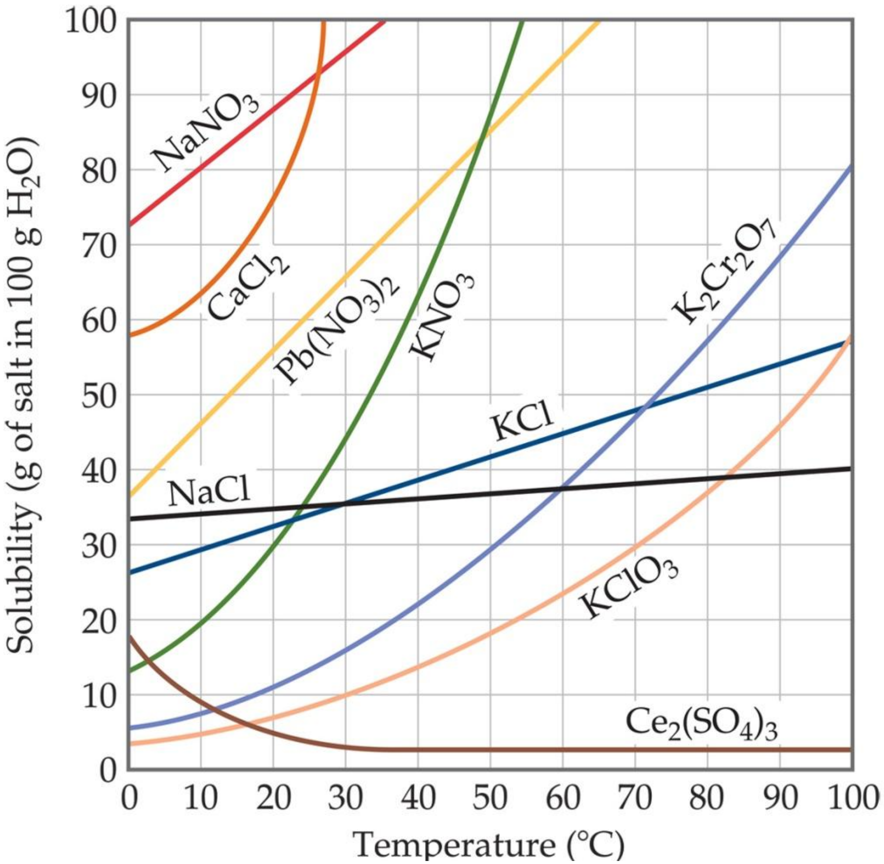
Factors Affecting Solubility

- Vitamin A is soluble in nonpolar compounds (like fats).
- Vitamin C is soluble in water.



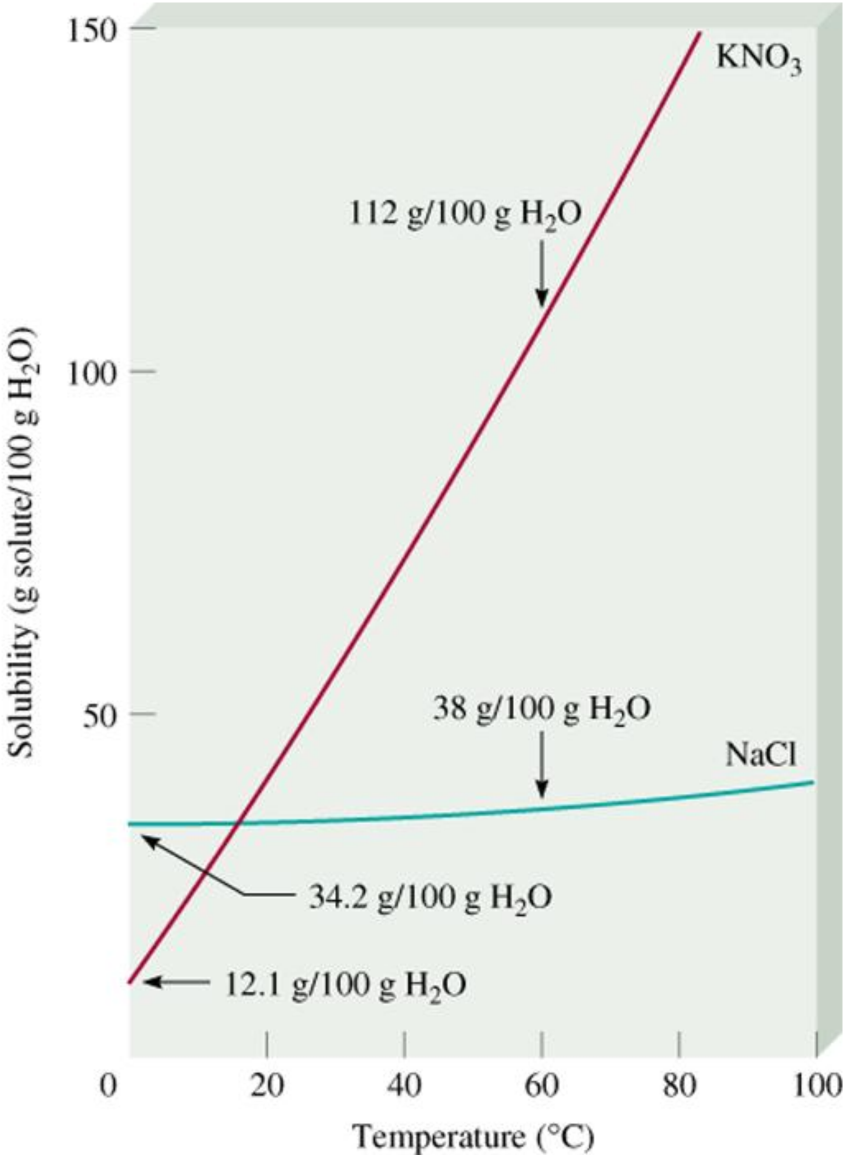
Vitamin A

Temperature



Generally, the solubility of solid solutes in liquid solvents increases with increasing temperature.

Fractional crystallization is the separation of a mixture of substances into pure components on the basis of their differing solubilities.



Suppose you have 90 g KNO_3 contaminated with 10 g NaCl.

Fractional crystallization:

1. Dissolve sample in 100 mL of water at 60^0C
2. Cool solution to 0^0C
3. All NaCl will stay in solution (s = 34.2g/100g)
4. 78 g of PURE KNO_3 will precipitate (s = 12 g/100g). 90 g – 12 g = 78 g

Gases in Solution

- In general, the solubility of gases in water increases with increasing mass.
- Larger molecules have stronger dispersion forces.

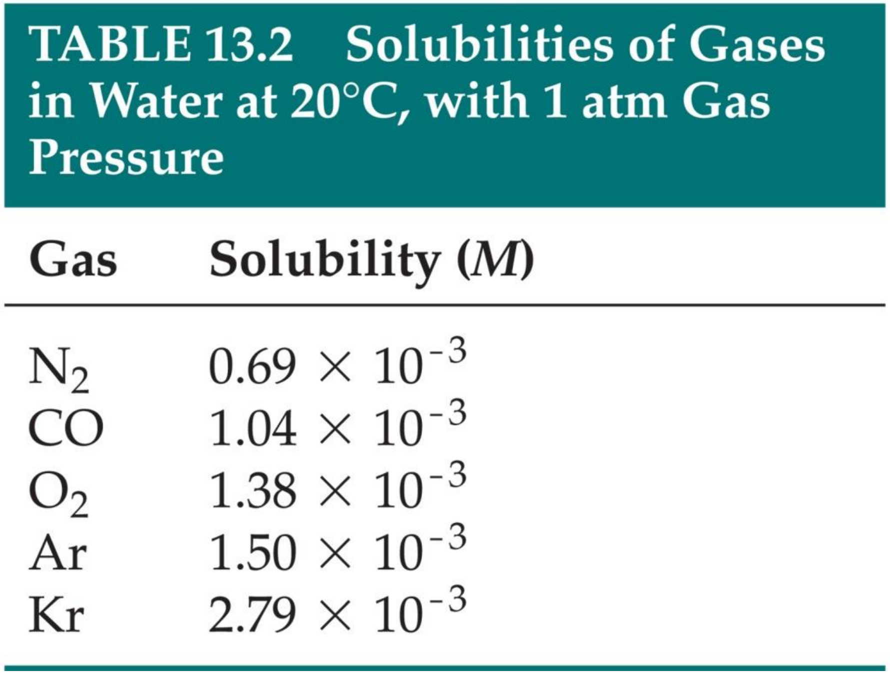
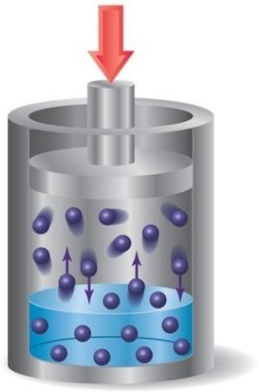


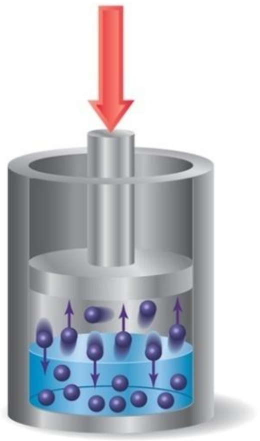
TABLE 13.2 Solubilities of Gases in Water at 20°C, with 1 atm Gas Pressure

Gas	Solubility (*M*)
N_2	0.69×10^{-3}
CO	1.04×10^{-3}
O_2	1.38×10^{-3}
Ar	1.50×10^{-3}
Kr	2.79×10^{-3}

Gases in Solution



(a)



(b)

- The solubility of liquids and solids does not change appreciably with pressure.
- The solubility of a gas in a liquid is directly proportional to its pressure.

Henry's Law

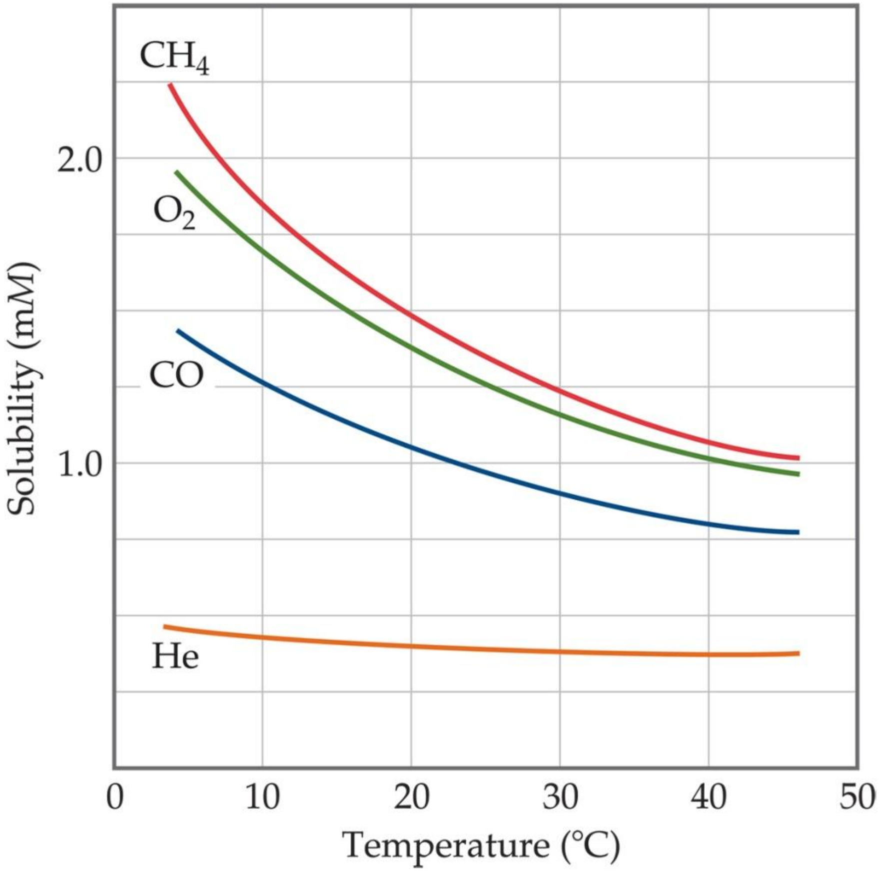
$$S_g = kP_g$$

where

- S_g is the solubility of the gas;
- k is the Henry's law constant for that gas in that solvent;
- P_g is the partial pressure of the gas above the liquid.

Temperature

- The opposite is true of gases:
  - Carbonated soft drinks are more “bubbly” if stored in the refrigerator.
  - Warm lakes have less O_2 dissolved in them than cool lakes.



Ways of Expressing Concentrations of Solutions

- mass percentage
- parts per million (ppm)
- parts per billion (ppb)
- Mole fraction (X)
- molarity (M)
- molality (m)

- The concentration of a solution is the amount of solute present in a given quantity of solvent or solution

Mass Percentage

$$\text{Mass \% of A} = \frac{\text{mass of A in solution}}{\text{total mass of solution}} \times 100$$

Mole Fraction (*X*)

$$X_A = \frac{\text{moles of A}}{\text{total moles in solution}}$$

- In some applications, one needs the mole fraction of *solvent*, not solute—make sure you find the quantity you need!

Parts per Million and Parts per Billion

Parts per Million (ppm)

$$\text{ppm} = \frac{\text{mass of A in solution}}{\text{total mass of solution}} \times 10^6$$

Parts per Billion (ppb)

$$\text{ppb} = \frac{\text{mass of A in solution}}{\text{total mass of solution}} \times 10^9$$

Molarity (*M*)

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

- You will recall this concentration measure from Chapter 4.
- Because volume is temperature dependent, molarity can change with temperature.

Molality (*m*)

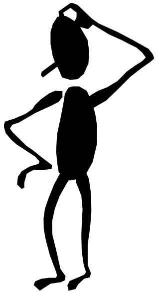
$$m = \frac{\text{mol of solute}}{\text{kg of } \boxed{\text{solvent}}}$$

Because both moles and mass do not change with temperature, molality (unlike molarity) is *not* temperature dependent.

Solution Stoichiometry (Chapter 4)

The ***concentration*** of a solution is the amount of solute present in a given quantity of solvent or solution.

$$\boldsymbol{M} = \textbf{\textit{molarity}} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

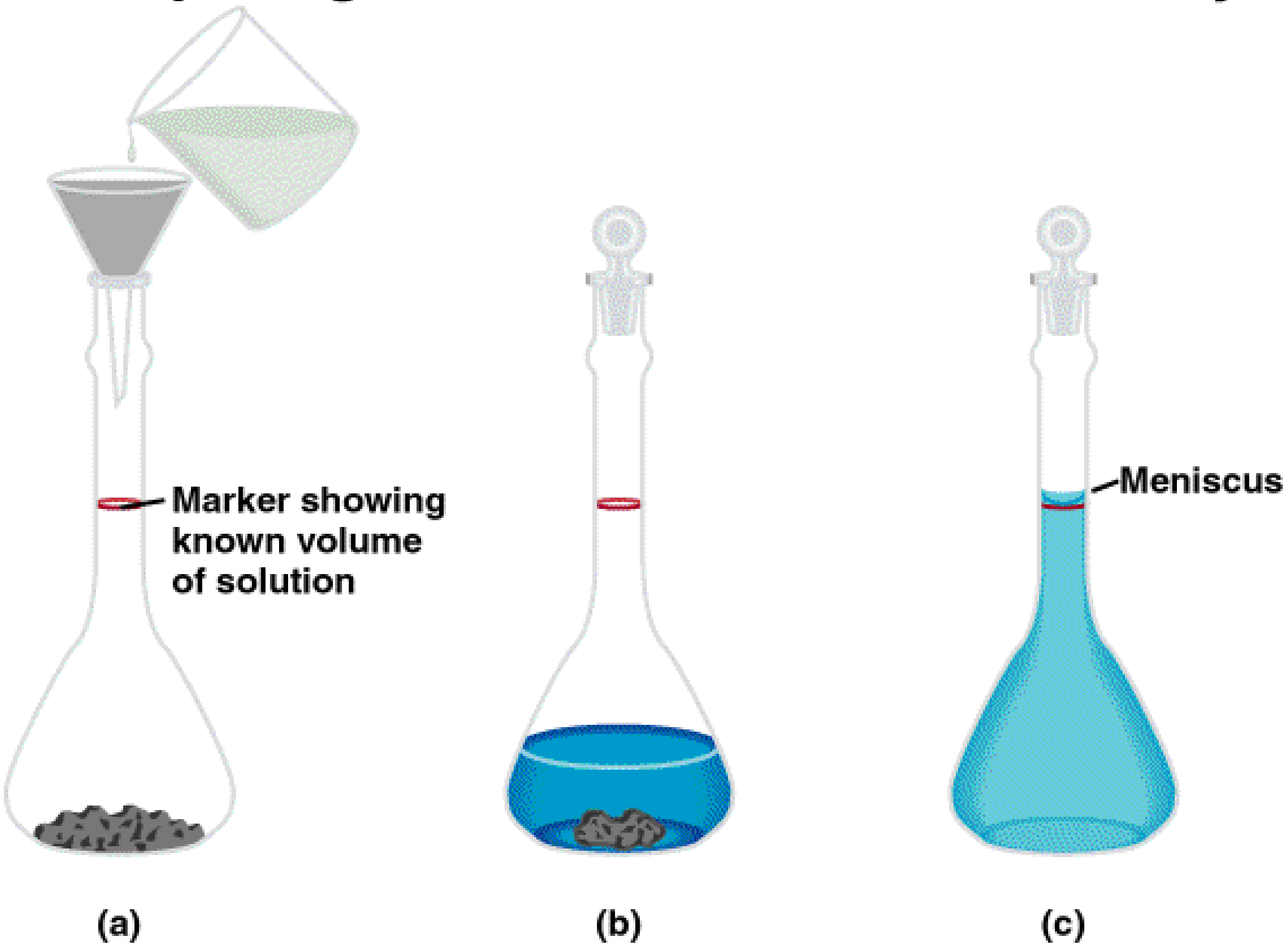


What mass of KI is required to make 500. mL of a 2.80 *M* KI solution?

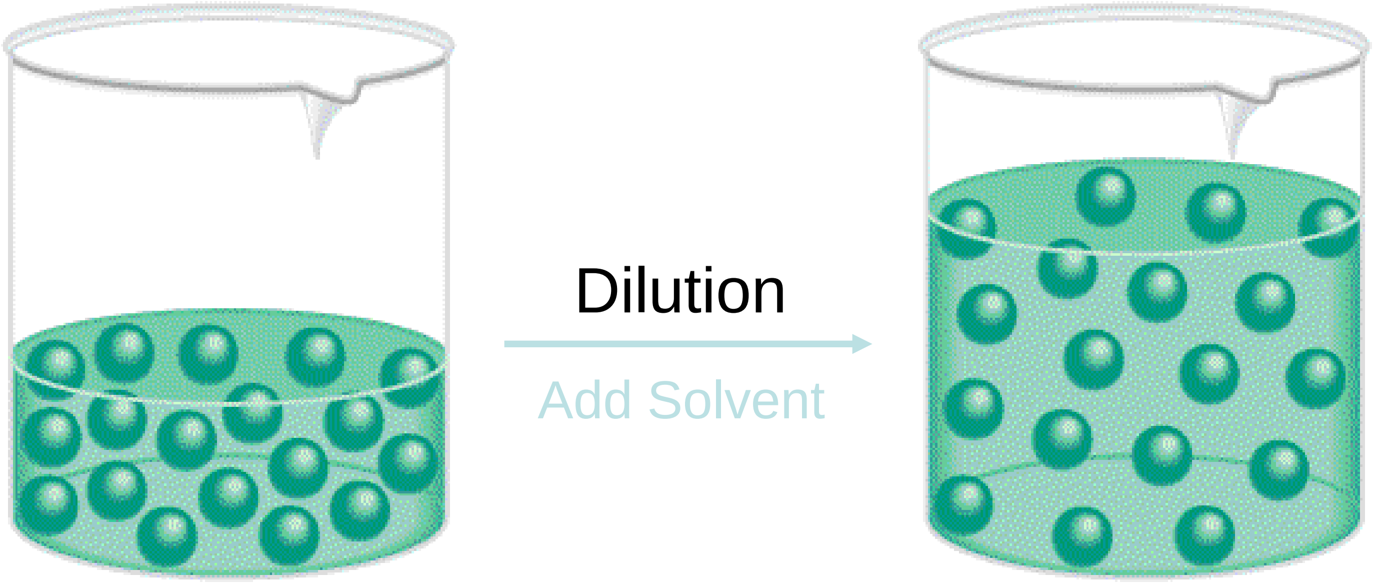
$$\text{volume KI} \xrightarrow{M\ \text{KI}} \text{moles KI} \xrightarrow{\mathcal{M}\ \text{KI}} \text{grams KI}$$

$$500.\ \cancel{\text{mL}} \times \frac{1\ \cancel{\text{L}}}{1000\ \cancel{\text{mL}}} \times \frac{2.80\ \cancel{\text{mol}}\ \text{KI}}{1\ \cancel{\text{L}}\ \text{soln}} \times \frac{166\ \text{g KI}}{1\ \cancel{\text{mol}}\ \text{KI}} = 232\ \text{g KI}$$

Preparing a Solution of Known Molarity

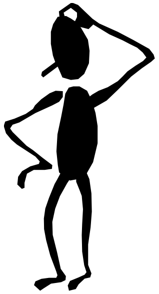


Dilution is the procedure for preparing a less concentrated solution from a more concentrated solution.



Moles of solute before dilution (i) = Moles of solute after dilution (f)

$$M_i V_i = M_f V_f$$



How would you prepare 60.0 mL of 0.2 *M* HNO_3 from a stock solution of 4.00 *M* HNO_3?

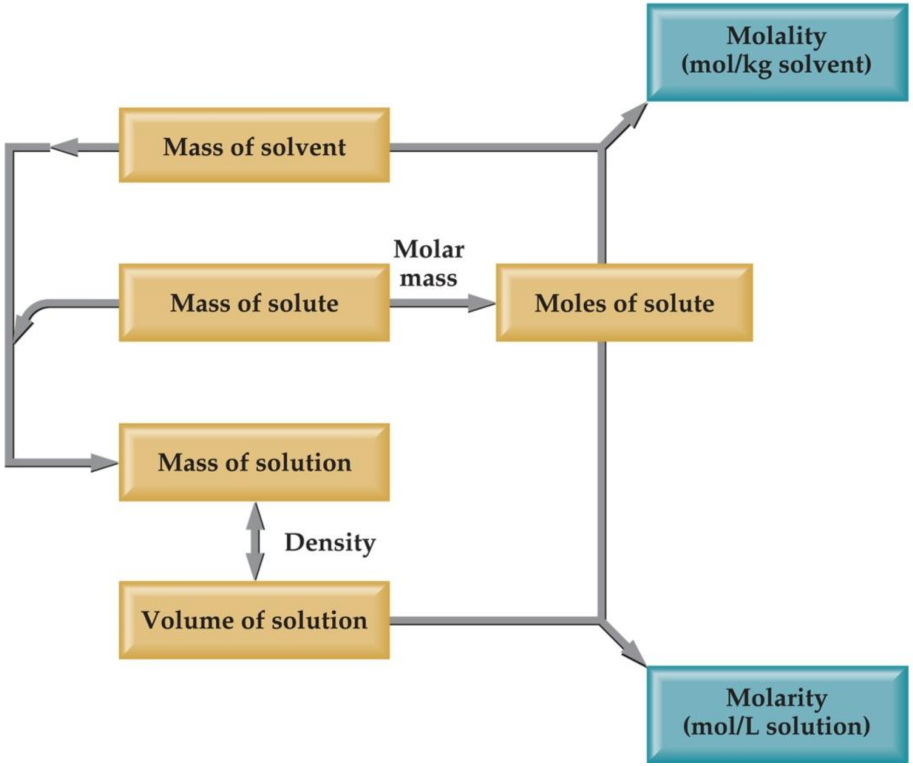
$$M_iV_i = M_fV_f$$

$M_i = 4.00$ $M_f = 0.200$ $V_f = 0.06\ L$ $V_i = ?\ L$

$$V_i = \frac{M_fV_f}{M_i} = \frac{0.200 \times 0.06}{4.00} = 0.003\ L = 3\ mL$$

3 mL of acid + 57 mL of water = 60 mL of solution

Changing Molarity to Molality



If we know the density of the solution, we can calculate the molality from the molarity, and vice versa.

What is the molality of a 5.86 *M* ethanol (C_2H_5OH) solution whose density is 0.927 g/mL?

$$\boldsymbol{m} = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}} \qquad \boldsymbol{M} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Assume 1 L of solution:

5.86 moles ethanol = 270 g ethanol

927 g of solution (1000 mL x 0.927 g/mL)

mass of solvent = mass of solution – mass of solute

= 927 g – 270 g = 657 g = 0.657 kg

$$\boldsymbol{m} = \frac{\text{moles of solute}}{\text{mass of solvent (kg)}} = \frac{5.86 \text{ moles } C_2H_5OH}{0.657 \text{ kg solvent}} = 8.92\ m$$

Gravimetric Analysis

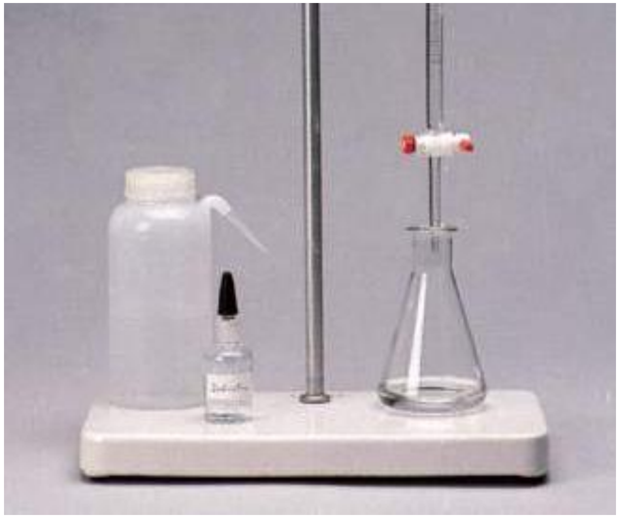
1. Dissolve unknown substance in water
2. React unknown with known substance to form a precipitate
3. Filter and dry precipitate
4. Weigh precipitate
5. Use chemical formula and mass of precipitate to determine amount of unknown ion

Titrations

In a ***titration*** a solution of accurately known concentration is added gradually added to another solution of unknown concentration until the chemical reaction between the two solutions is complete.

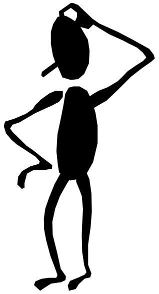
Equivalence point – the point at which the reaction is complete

Indicator – substance that changes color at (or near) the equivalence point



Slowly add base to unknown acid UNTIL

the indicator changes color



What volume of a 1.420 *M* NaOH solution is Required to titrate 25.00 mL of a 4.50 *M* H_2SO_4 solution?

WRITE THE CHEMICAL EQUATION!

$$H_2SO_4 + 2NaOH \longrightarrow 2H_2O + Na_2SO_4$$

$$\text{volume acid} \xrightarrow[\text{acid}]{M} \text{moles acid} \xrightarrow[\text{coef.}]{\text{rx}} \text{moles base} \xrightarrow[\text{base}]{M} \text{volume base}$$

$$25.00\ \cancel{\text{mL}} \times \frac{4.50\ \cancel{\text{mol H}_2\text{SO}_4}}{1000\ \cancel{\text{mL soln}}} \times \frac{2\ \cancel{\text{mol NaOH}}}{1\ \cancel{\text{mol H}_2\text{SO}_4}} \times \frac{1000\ \text{ml soln}}{1.420\ \cancel{\text{mol NaOH}}} = 158\ \text{mL}$$