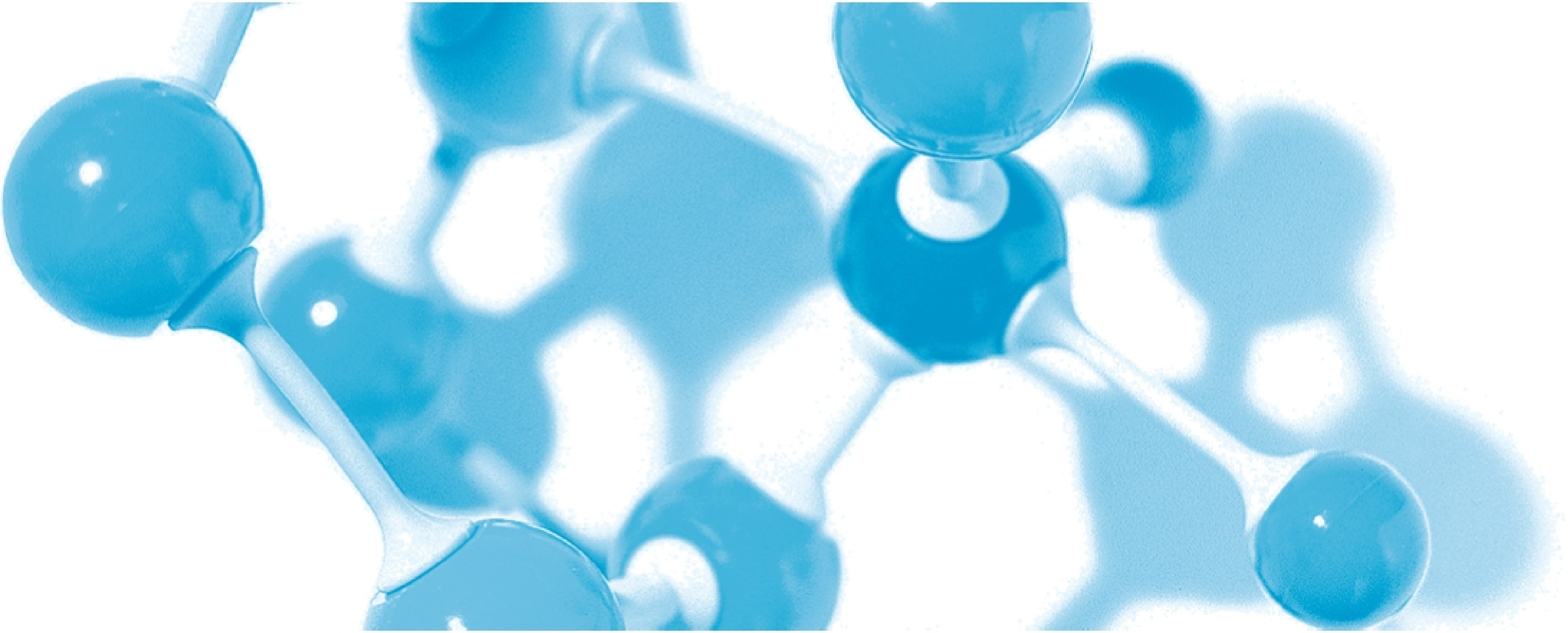


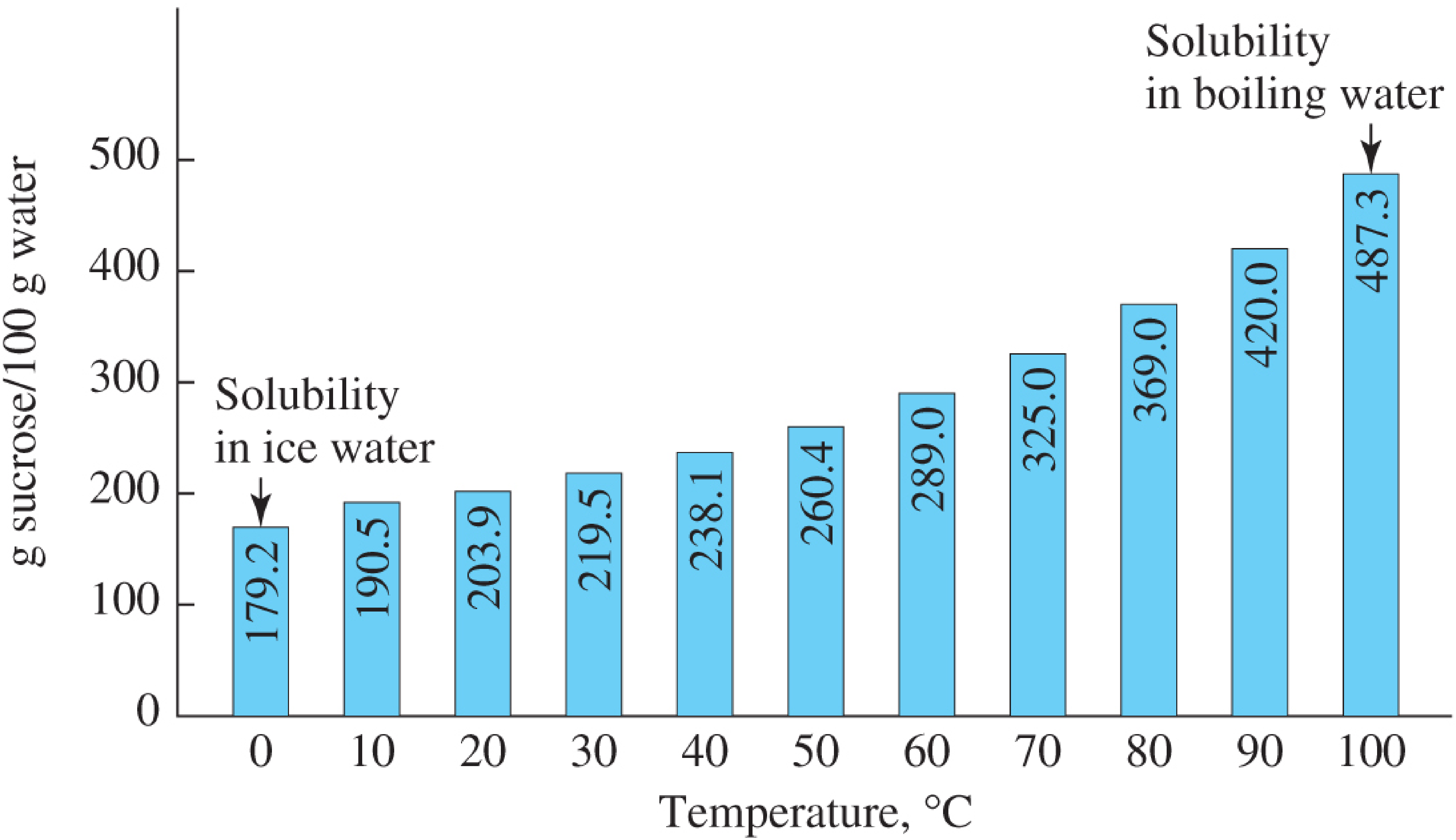
- Solution – a homogeneous mixture of two or more substances (a physical mixture)
  - Solvent – the major component of the solution
  - Solute – the minor component(s) of the solution

TABLE 13.1 Examples of Various Types of Solutions

Solution Type (solute listed first)		Example
Gaseous Solutions		
	Gas dissolved in gas	Dry air (oxygen and other gases dissolved in nitrogen)
	Liquid dissolved in gas*	Wet air (water vapor in air)
	Solid dissolved in gas*	Moth repellent (or mothballs) sublimed into air
Liquid Solutions		
	Gas dissolved in liquid	Carbonated beverage (carbon dioxide in water)
	Liquid dissolved in liquid	Vinegar (acetic acid dissolved in water)
	Solid dissolved in liquid	Salt water
Solid Solutions		
	Gas dissolved in solid	Hydrogen in platinum
	Liquid dissolved in solid	Dental filling (mercury dissolved in silver)
	Solid dissolved in solid	Sterling silver (copper dissolved in silver)

**An alternative viewpoint is that liquid-in-gas and solid-in-gas solutions do not actually exist as true solutions. From this viewpoint, water vapor or moth repellent in air is considered to be a gas-in-gas solution since the water or moth repellent must evaporate or sublime first in order to enter the air.*

- Solubility – the maximum amount of solute that will dissolve in a standard amount of solvent at a given temperature and pressure
  - Saturated solution – when the concentration of solute is at the solubility limit.
  - Unsaturated solution – when the solute concentration is below the solubility limit.
  - Supersaturated solution – when the solute concentration is above the solubility limit.
    - This is an unstable solution.
    - Some solute will precipitate to reduce the concentration to a saturated solution.



Solubility
in boiling water
Solubility
in ice water
g sucrose/100 g water
500
400
300
200
100
0
179.2
190.5
203.9
219.5
238.1
260.4
289.0
325.0
369.0
420.0
487.3
0
10
20
30
40
50
60
70
80
90
100
Temperature, °C


Saturated
solution
Undissolved solute

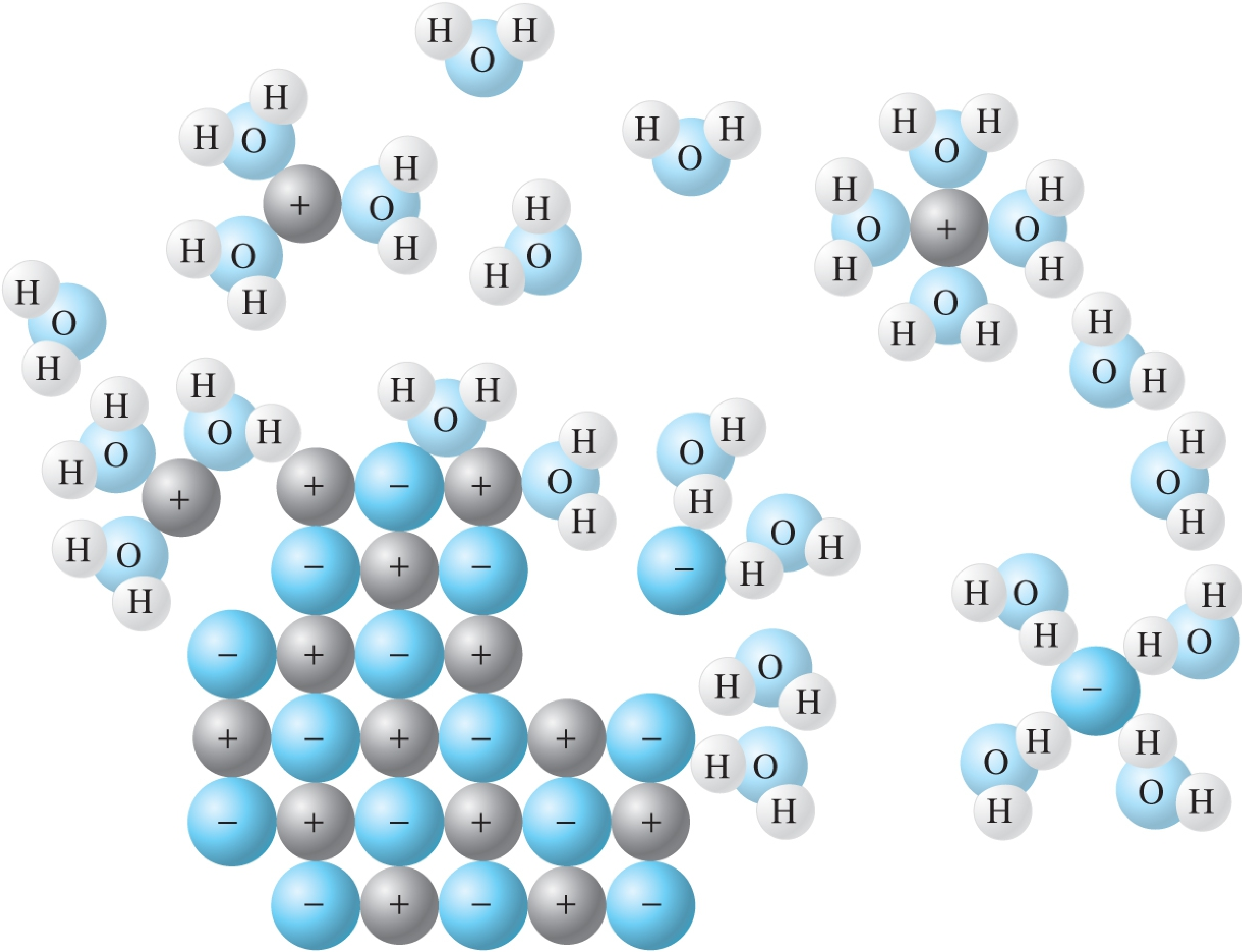
TABLE 13.2 Solubilities of Various Compounds in Water at 0°C, 50°C, and 100°C

	Solubility (g solute/100 g H_2O)		
Solute	**0°C**	**50°C**	**100°C**
Lead(II) bromide ($PbBr_2$)	0.455	1.94	4.75
Silver sulfate (Ag_2SO_4)	0.573	1.08	1.41
Copper(II) sulfate ($CuSO_4$)	14.3	33.3	75.4
Sodium chloride (NaCl)	35.7	37.0	39.8
Silver nitrate ($AgNO_3$)	122	455	952
Cesium chloride (CsCl)	161.4	218.5	270.5

Solution Formation (dissolving solids in liquids)

- Solvent molecules bring solute into solution one molecule/ion at a time.

- Increased rate of dissolving with:
  - agitation – brings low conc. solvent to solid
  - larger surface area – more solvent is able to interact with solid surface
  - higher temperatures – solid more easily dissolved with higher kinetic energy



Solubility Rules

- General guidelines to indicate if an ionic compound is soluble in water at room temp.

- 'Solubility' is a qualitative description.

  - Compounds listed as 'soluble' all have different solubility limits.

  - Even 'insoluble' compounds are very slightly soluble.

TABLE 13.3 Qualitative Solubility Terms

Solute Solubility (g solute/100 g solvent)	Qualitative Solubility Description
Less than 0.1	insoluble
0.1–1	slightly soluble
1–10	soluble
Greater than 10	very soluble

TABLE 13.4 Solubility Guidelines for Ionic Compounds in Water

Soluble Compounds	Important Exceptions
Compounds containing the following ions are soluble, with exceptions as noted.	
Group IA (Li^+, Na^+, K^+, etc.)	none
Ammonium (NH_4^+)	none
Acetate ($C_2H_3O_2^-$)	none
Nitrate (NO_3^-)	none
Chloride (Cl^-), bromide (Br^-), and iodide (I^-)	Ag^+, Pb^{2+}, Hg_2^{2+}
Sulfate (SO_4^{2-})	Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}
Insoluble Compounds	**Important Exceptions**
Compounds containing the following ions are insoluble, with exceptions as noted.	
Carbonate (CO_3^{2-})	group IA and NH_4^+
Phosphate (PO_4^{3-})	group IA and NH_4^+
Sulfide (S^{2-})	groups IA and IIA and NH_4^+
Hydroxide (OH^-)	group IA, Ca^{2+}, Sr^{2+}, Ba^{2+}

Concentration

- Ratio of solute to solvent
- Independent of the volume of the sample
- Measured in many different units

- Percent by mass = (mass solute/mass solution)*100%
- Percent by volume = (vol. solute/vol. solution)*100%
- Percent by number = (mole solute/mole solution)*100%
- parts per million (ppm) = parts solute/million parts solution
- parts per billion (ppb) = parts solute/billion parts solution

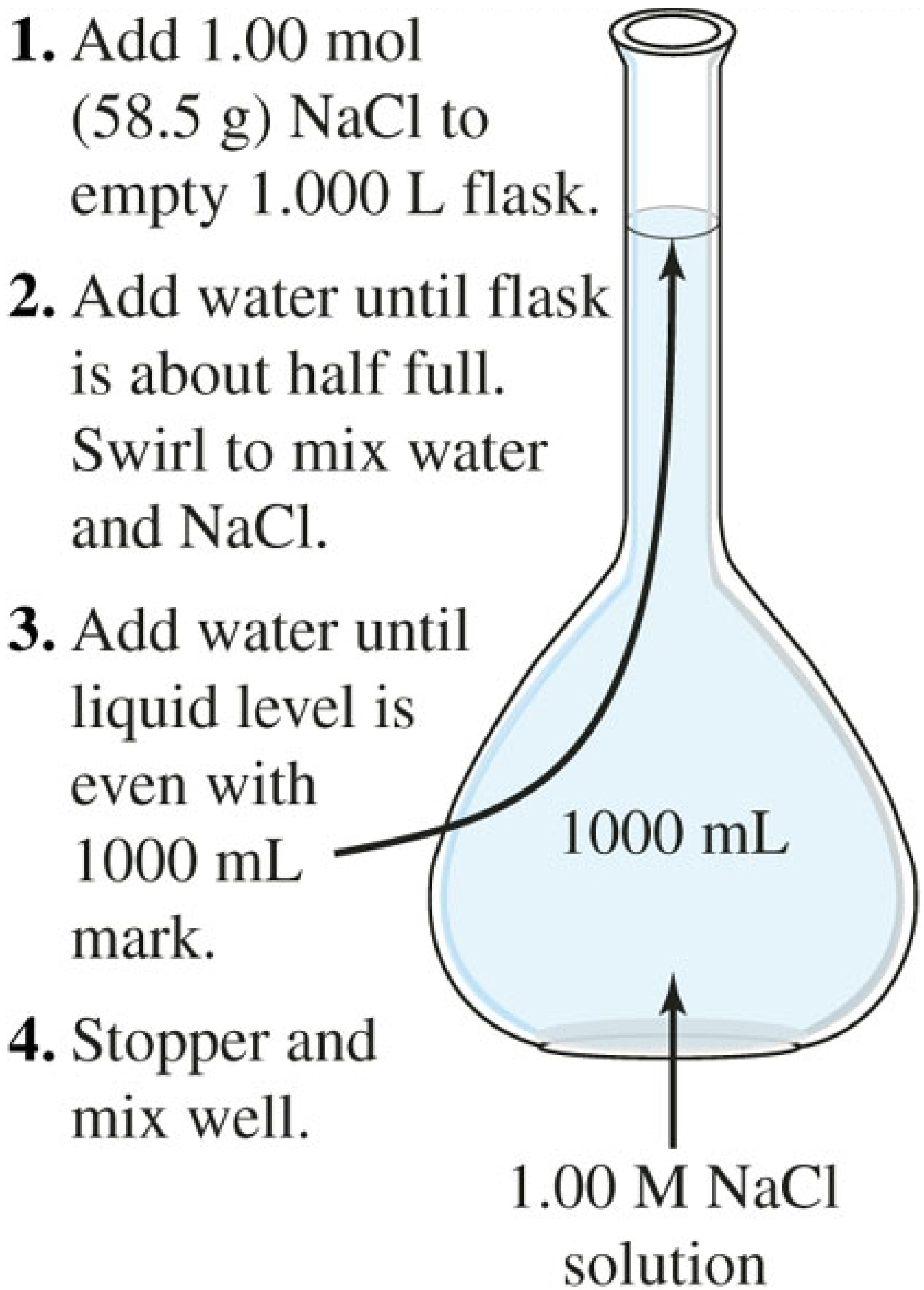
Molarity (M)

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

$$M = \frac{\text{n}}{\text{V}}$$

Molarity makes conversion between moles and volume easy to calculate.

How to prepare a 1.00 M NaCl solution:



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

Note – you do NOT add 58.5 g NaCl to 1.00 L of water.

The 58.5 g will take up some volume, resulting in slightly more than 1.00 L of solution – and the molarity would be lower.

Dilutions

- Solvent added to a solution to reduce concentration
- Moles of solute is constant – dilution has the same amount of solute as original solution
- $M_1V_1 = M_2V_2$

where M_1V_1 are the original molarity and volume and M_2V_2 are the dilution molarity and volume

Stoichiometry of Aqueous Reactions

- Convert volume and molarity to moles.

- Treat as standard stoichiometric problem to determine the desired moles.

- Convert this value into molarity, if needed.

- Note: if solutions of two reactants are mixed, the remember that the final volume is the sum of the initial volumes.

If the concentrations of both reactants in the *reactant mixture* are known, then:

- Molarity may be used in place of moles for stoichiometric calculations.
- The unknown value calculated will also be in molarity.
- If two solutions are combined for the reaction, the molarities in this mixture must be used, not the initial molarities before mixing.