Chemistry 3A

Introductory General Chemistry

S. M. Halloran

Fresno City College

- Mixtures: Heterogeneous & Homogeneous
- Solutions & Aqueous Solutions
- Solubility of Solids and Gases
- Concentration: Percent of Dissolution, Molarity, Calculations
- Colligative Properties, Molality, Freezing Point Depression, Boiling Point Elevation, Osmosis

Solutions

- **Solutions** are homogeneous mixtures which are usually visibly liquid but can also be gaseous
  - Mixture: A combination of two or more substances that are physically mixed (not chemically bonded). Each substance retains its own chemical identity and can often be separated by physical means

    Trail mix, sand and water are heterogeneous; saltwater and air are homogeneous
  - Homogeneous: Such mixtures are look like a single substance but are composed of two or more substances that are indistinguishable by eye

    Saltwater, air, brass (a metal alloy).

Solutions

- The **solvent** is the major component of the solution
- The **solute** is the minor component of the solution

“Major” here means its mass or particle count (moles) is far larger than the mass or particle count (moles) of what is “minor”

- Salt water has the minor component sodium chloride ($NaCl$) dissolved in the major component H_2O as its liquid (water)
- Even air (the atmosphere) is considered a solution in which the solute diatomic oxygen (O_2) is dissolved in a solvent of diatomic nitrogen (N_2)

Air is 78% N_2 and 21% O_2

Solutions: Quiz Time

- **1.00 g** sucrose ($C_{12}H_{22}O_{11}$) is dissolved in **100.0 g** of H_2O liquid (water). What is the solvent and what is the solute?

  The sucrose is clearly a minor component by mass compared to water. Sucrose is the **solute**, water the **solvent**

- **3.33 g** HCl gas is dissolved in **40.0 g** of CH_3OH (methanol, methyl alcohol). Identify solvent and solute

  HCl (*g*) **solute**, CH_3OH (*l*) **solvent**

Table 9.1.1: Types of Solutions

Solvent Phase	Solute Phase	Example
gas	gas	air
liquid	gas	carbonated beverages
liquid	liquid	ethanol (C_2H_5OH) in H_2O (alcoholic beverages)
liquid	solid	salt water
solid	gas	H_2 gas absorbed by Pd metal
solid	liquid	Hg(ℓ) in dental fillings
solid	solid	steel alloys

Figure 9.1.1: Making a saline water solution by dissolving table salt (NaCl) in water. The salt is the solute and the water the solvent. (CC-BY-SA 3.0; Chris 73).

Will It Dissolve?

- "Like dissolves like" generally
- This means polar solutes in polar solvents. But ionic solutes also dissolve in polar solvents
- Nonpolar solutes dissolve in nonpolar solvents
- EXCEPTIONS: a very small amount (fraction) of nonpolar solutes mix polar solvents (O_2 and CO_2 dissolving in H_2O), and same for polar solutes in nonpolar solvents (water in oil)

Table 9.1.1.2: Summary of Solubilities

Solute (Polarity of Compound)	**Solvent (Polarity of Compound)**	**Dominant Intermolecular Force**	**Is Solution Formed?**
Polar	Polar	Dipole-Dipole Force and/or Hydrogen Bond	yes
Non-polar	Non-polar	Dispersion Force	yes
Polar	Non-polar		no
Non-polar	Polar		no
Ionic	Polar	Ion-Dipole	yes
Ionic	Non-polar		no

Figure 9.1.1.2: Water (clear liquid) and oil (yellow) do not form liquid solutions. (CC BY-SA 1.0 Generic; Victor Blacus)

Practice: Will It Dissolve?

- Would I_2 (diatomic molecule iodine) be more soluble in CCl_4 (carbon tetrachloride) or H_2O (water)?

  I_2 is **nonpolar**. CCl_4 is nonpolar and H_2O is polar, so expect I_2 to be much more soluble in CCl_4 than H_2O

- Would C_3H_7OH (any of isomers of propanol) be more soluble in CCl_4 (carbon tetrachloride) or H_2O (water)?

  C_3H_7OH is **polar** particularly because it is capable of hydrogen bonding, and H_2O also does hydrogen bonding. Expect C_3H_7OH to be much more soluble in H_2O than CCl_4

Practice: Will It Dissolve?

- Which of these will be solutes in the solvent water?

a. Methanol (CH_3OH)

Both water and methanol have hydroxyl (–OH) functional groups that make these molecules **polar**. So they will mix

Note methanol (methyl alcohol) is a solvent in its own right: so two solvents that are mixed to form solution are said to be **miscible**

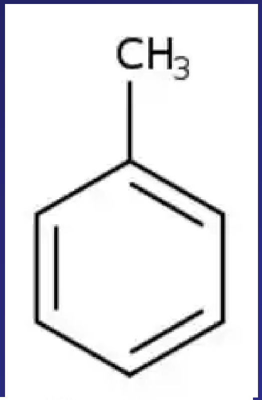
b. Sodium sulfate (Na_2SO_4)

Na_2SO_4 is an ionic compound. Ionic compounds have atoms (or polyatomic molecules as ions) that have permanent electric charge character, either positive (cationic) or negative (anionic) charge. These always prefer a **polar** than a nonpolar solvent, so will dissolve in water

c. Octane (C_8H_{18})

C_8H_{18} is a **nonpolar** compound. Compounds of just carbon and hydrogen will generally be nonpolar. So it will not dissolve

Practice: Will It Dissolve?



Which of these will be substances will dissolve in **nonpolar** solvent toluene ($C_6H_5CH_3$)

a. H_2O (water)

Water is quite **polar**. And following the rule of "like dissolves like" which means "polar dissolves polar" and "nonpolar dissolves nonpolar", these will not mix (see image)

toluene - water

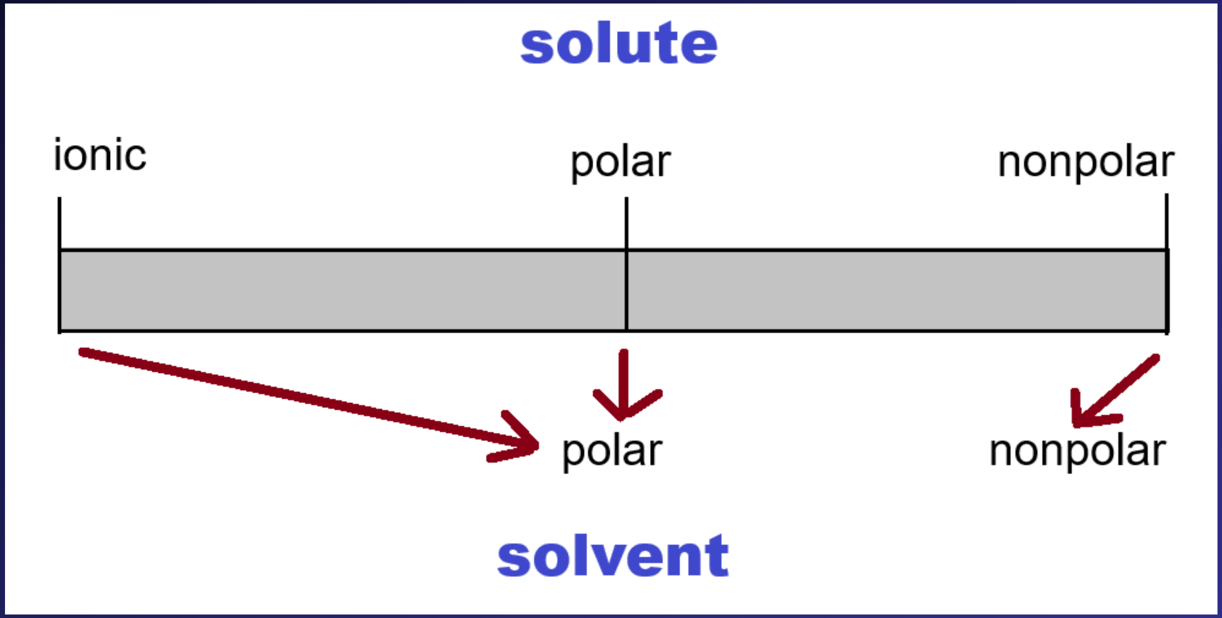
b. Na_2SO_4 (sodium sulfate)

It was noted that Na_2SO_4 is an ionic compound. Ionic compounds will interact with polar solvents ONLY

c. C_8H_{18} (octane)

As a **nonpolar** compound, C_8H_{18} will have no problem dissolving in a nonpolar solvent like toluene

“Like Prefers to Dissolve in Like”



What about ionic solvents?

Yes there are **ionic solvents**

- Molten salts (melted NaCl, KNO3, etc)
- Room temperature ionic liquids (RTILs)

Electrolytes / Nonelectrolytes

- Ionic compounds dissolving in a solvent create a solution with **electrolytes**
- Substances dissolving but not being ions (ionic) in solution are **nonelectrolytes**
- If the ionic solute dissolves (almost) entirely in the solvent, it is a **strong** electrolyte and a good conductor
- If ionic solute dissolves only partially, it is likely a **weak** electrolyte and likely not a good conductor

Carrying Electric Current

- Aqueous solutions with dissolved electrolytes conduct electricity
- The dissolved ions move freely in the solvent when an external voltage is applied

Dissolved ions of the one charge move to the electrode of opposite charge, creating the continuity that is the electric current

If solute does not produce ions, there is no electric current

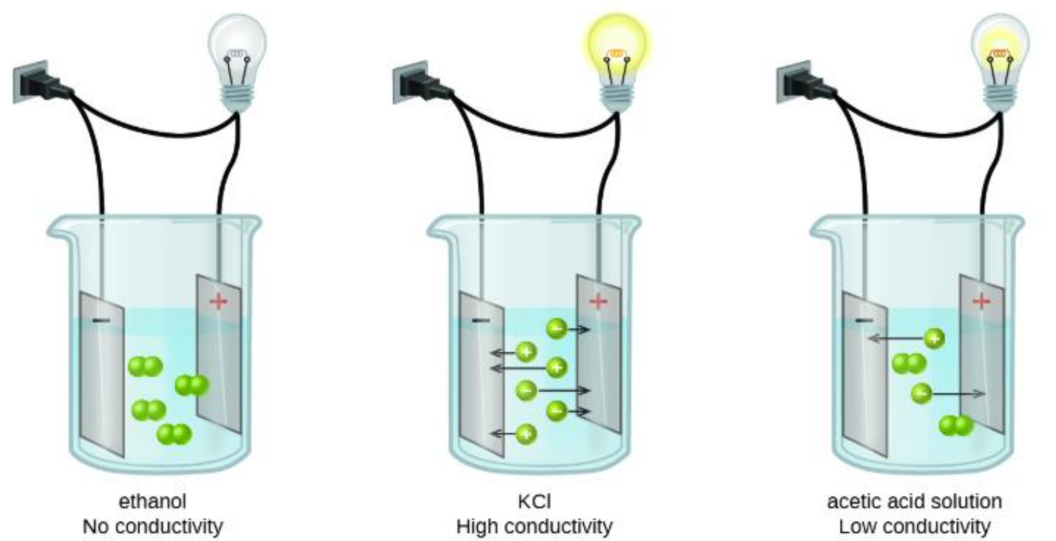
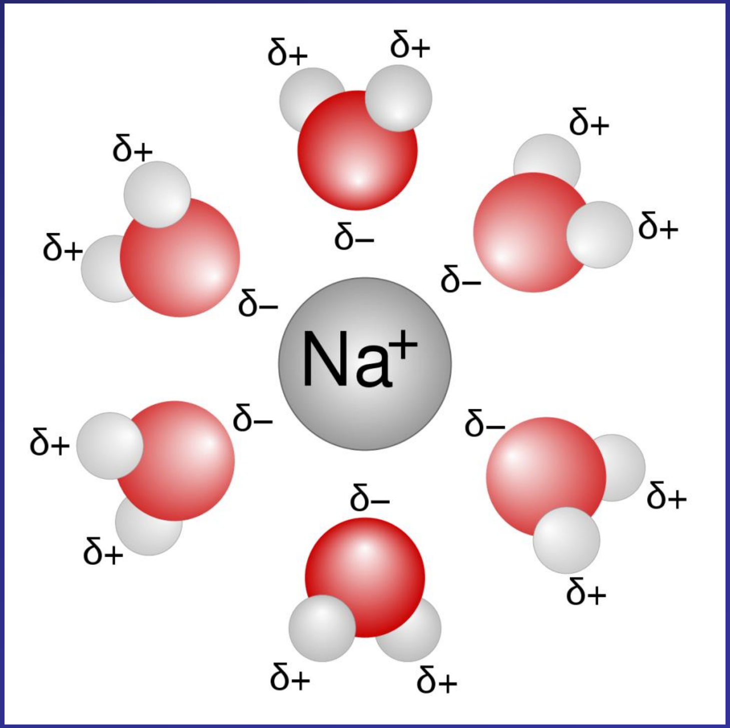


Figure 9.1.2.1: Solutions of nonelectrolytes, such as ethanol, do not contain dissolved ions and cannot conduct electricity. Solutions of electrolytes contain ions that permit the passage of electricity. The conductivity of an electrolyte solution is related to the strength of the electrolyte.

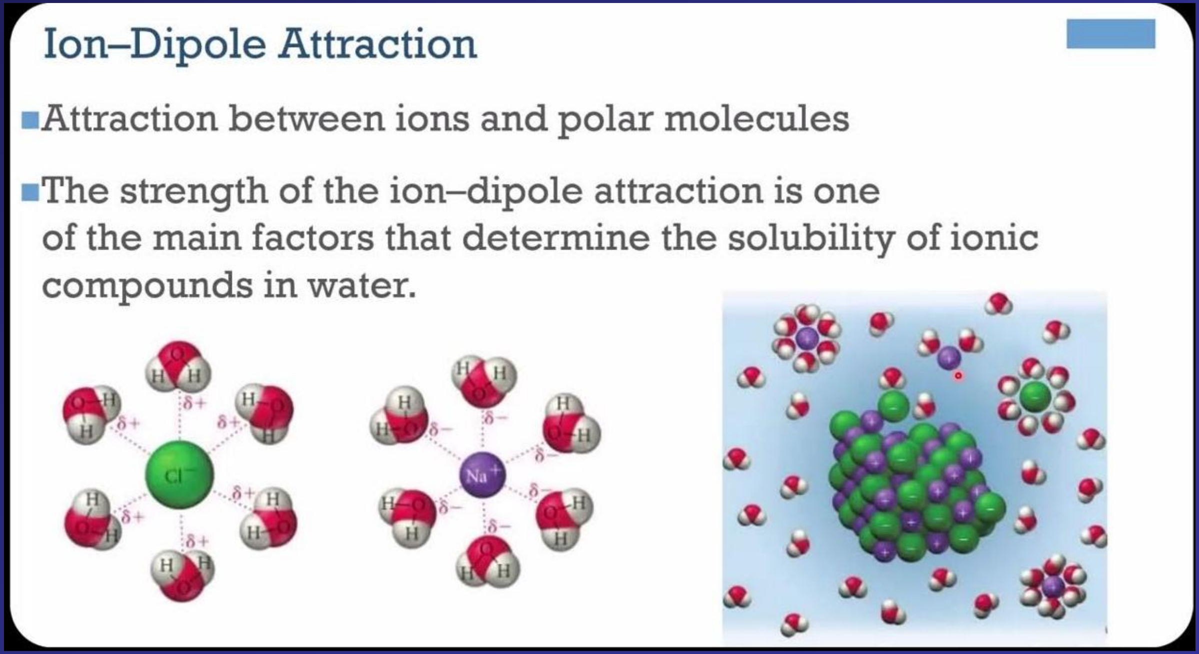
Ion-Dipole Attraction

- Ionic compounds (salts) dissolve as ions in polar solvents like water
- Water has no electric charge like a dissolved ion
- But it DOES have a **dipole** (recall from previous unit)
- That dipole in the H_2O molecule formed by the electron group & molecular geometry of the molecule enables interactions with dissolved ions
- For NaCl (sodium chloride), note how the oxygen (O) atom with its negative dipole orients itself to Na^+ (sodium) ions
- And the hydrogen (H) atoms with their positive dipole will orient themselves to the Cl^- (chlorine/ "chloride") ions



Ion-Dipole Attraction

- The ion-dipole attraction actually helps to dissolve the salt crystal/solid
- During the dissolution process, ions leave the solid (crystal) and are solvated in a physical process called **dissociation**, forming the electrolytic solution



Which Become Ions in Solution?

- It needs to be an ion (an ionic compound)
- Except for polyatomic molecules forming ions, molecules whether polar or nonpolar will not form ions!

❖ LiF (lithium fluoride)

yes

❖ P_2F_5 (diphosphorus pentafluoride)

no: all bonded are covalent and don't break apart in a solvent

❖ C_2H_5OH (CH_3CH_2OH) (ethanol)

no, a covalent molecule

❖ $C_6H_{12}O_6$ (glucose)

no, a covalent molecule

❖ CCl_4 (carbon tetrachloride)

no a covalent molecule and also nonpolar

❖ $CaCl_2$ (calcium chloride)

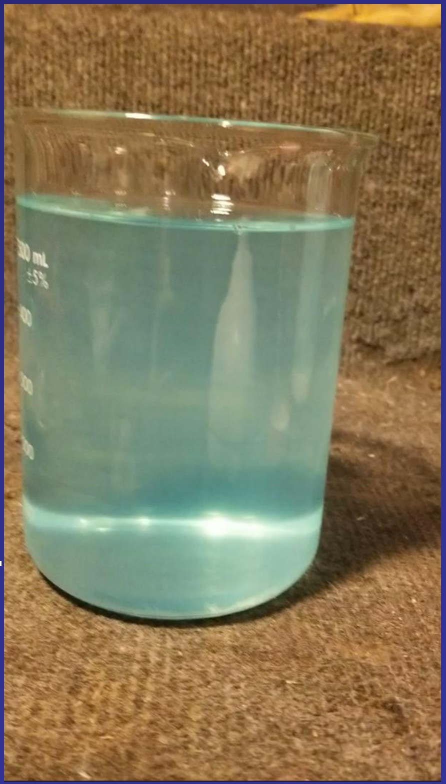
yes

❖ $AgNO_3$ (silver nitrate)

yes

Solubility in Aqueous Solutions

- Some ions (a cation and an anion) that are brought together in solution will form a complex quickly and ***precipitate*** out of solution, forming a solid in the solution
- A classic example is silver (Ag^+) cation binding or combining (not bonding) with chloride (Cl^-) anion to form the silver chloride ($AgCl$) precipitate
- Several ion combinations occurring in solution and having no or low solubility—thus forming precipitates—are known (table next page)



Video shows HCl being added to $AgNO_3$ (*aq*) to form AgCl (*s*)

Solubility in Aqueous Solutions

- These rules for solubility concerning the various ion combinations in aqueous solutions are part of the Green Sheet

If you are asked about whether the mixing of solutions with soluble ions form a precipitate, be prepared to say yes or no by understanding the table

Table 9.1.3.1: Solubility Rules for Soluble Substances

Compounds containing these ions are generally soluble (aq)	**… except combinations described below are insoluble (s)**
Group 1 (Li^+, Na^+, K^+, etc.), NH_4^+	**Except** Li^+ is slightly soluble with CO_3^{2-}, PO_4^{3-}, and F^-.
ClO_4^-, ClO_3^-, NO_3^-, $C_2H_3O_2^-$ / CH_3COO^-	None.
Cl^-, Br^-, I^-	**Except** for those containing Ag^+, Hg_2^{2+}, and Pb^{2+}.
F^-	**Except** for those containing Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}, and Pb^{2+}.
SO_3^{2-}, SO_4^{2-}	**Except** for those containing Ca^{2+}, Sr^{2+}, Ba^{2+}, Ag^+, and Pb^{2+}.

Compounds containing these ions are generally insoluble (s)	**… except combinations described below are soluble (aq)**
CO_3^{2-}, PO_4^{3-}	**Except** those of Group 1 and NH_4^+.
CrO_4^{2-}, $C_2O_4^{2-}$	**Except** those of Group 1 and NH_4^+.
O^{2-}, S^{2-}	**Except** those of Group 1, NH_4^+, Ca^{2+}, Sr^{2+}, and Ba^{2+}.
OH^-	**Except** those of Group 1, NH_4^+. **Except** OH^- is slightly soluble with Ca^{2+}, Sr^{2+}, and Ba^{2+}.

Practice: Is It Soluble?

Refer to the Green Sheet

a. $Zn(NO_3)_2$ (zinc nitrate)

The solubility table indicates ALL nitrates are soluble

b. $PbBr_2$ (lead(II) bromide)

Table states all bromides are soluble ***except*** when combined with Pb^{2+} lead ion

c. $Sr_3(PO_4)_2$ (strontium phosphate)

Table states all phosphates are insoluble (not soluble), so $Sr_3(PO_4)_2$ is

d. $Mg(OH)_2$ (magnesium hydroxide)

The Group 2 ion is insoluble with hydroxide ion

e. KBr (potassium bromide)

Soluble: Group 1 ions like K^+ are soluble except for Li^+ in some combination

f. $Pb(NO_3)_2$ (lead(II) nitrate)

Table states all nitrates are soluble with no exceptions

Solubility and Saturation

- Some salt (cation-anion) combinations like silver chloride ($AgCl$) are not really soluble at all (see table below)
- But many salts, while quite soluble, have a maximum limit of dissolution such that addition of more salt will NOT dissolve the added salt
- The solute has reached a maximum **saturation**
- Sodium chloride ($NaCl$) added above **36.1 g** to **100 g** H_2O will not dissolve the added $NaCl$.

The solution is said to be "saturated" with NaCl solute

Table 9.2.1: Solubilities of Some Ionic Compounds

Solute	Solubility (g per 100 g of H_2O at 25°C)
$AgCl$	0.00019
$CaCO_3$	0.0006
KBr	70.7
$NaCl$	36.1
$NaNO_3$	94.6

Saturated / Unsaturated

- When a solute has not reached its maximum saturation, it is called unsaturated

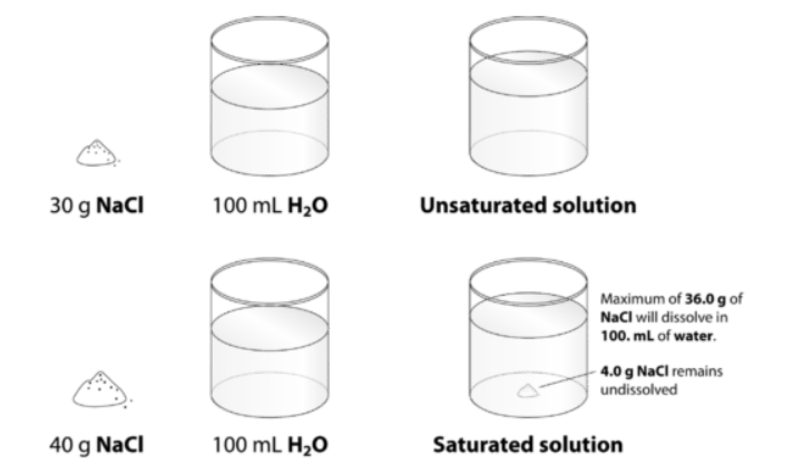


Figure 9.2.1: When $30.0\ g$ of $NaCl$ is added to $100\ mL$, it all dissolves, forming an unsaturated solution. When $40.0\ g$ is added, $36.0\ g$ dissolves and $4.0\ g$ remains undissolved, forming a saturated solution.

Supersaturation

Want to see something cool?

- Add 175 g sodium acetate trihydrate ($NaOOCCH_3 \bullet 3\ H_2O$) to a 50 mL distilled water in a 500 mL Erlenmeyer flask
- Heat flask inside a 2 L beaker containing 1.5 L water on hot plate until it dissolves
- Invert 100 mL beaker over mouth of flask and allow to cool (a couple of hours)
- Add a seed crystal or "scratch" the flask glass inside and watch the fun

Supersaturation

- Maximum solubility (limits of saturation) are temperature-dependent
- Solubility usually increases as the temperature goes up
- Suppose you heat a solution and add solute up to its maximum solubility while make sure all solute dissolves
- Now you cool the solution and solute mass is far above its maximum solubility for the temperature
- Yet solute remains dissolved (not precipitating out)
- This is **supersaturation**

Solids Dissolved in Water

- The **solubility** of a substance is the mass representing the upper limit of a solute dissolving in a solvent before the substance becomes saturating
- It is usually measured as g of solute per 100 g solvent
- Up to **36.0 g** NaCl can be added to **100 g** H_2O at 20°C
- Solubility is temperature-dependent
  Up to **32 g** KCl will dissolve in **100 g** H_2O at 20°C,
  but **55 g** KCl will dissolve in **100 g** H_2O at 100°C

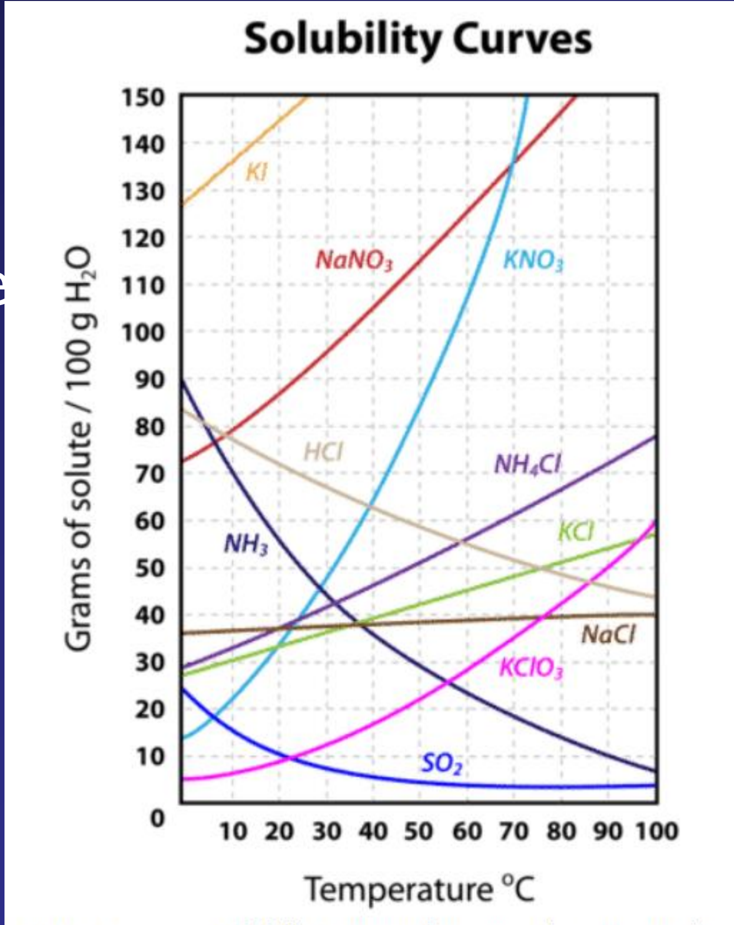


Figure 9.2.1.4: Solubility curves for several compounds.

Solids Solubility vs Temp

Note the trends in the solubility curves

- NaCl is flat
- KNO_3 goes steeply up with temp

Why?

- Dissolving NaCl is both endothermic and exothermic:
  - The endothermic part breaks the crystal lattice & this requires energy
  - The exothermic part is the hydration of Na^+ and Cl^-, and this release energy
- Dissolving KNO_3 requires a LOT of energy and is HIGHLY endothermic and the hydration of K^+ and NO_3^- does not release enough energy
- Increasing temperature would add energy to processes that are highly endothermic like dissolving solids needing the energy

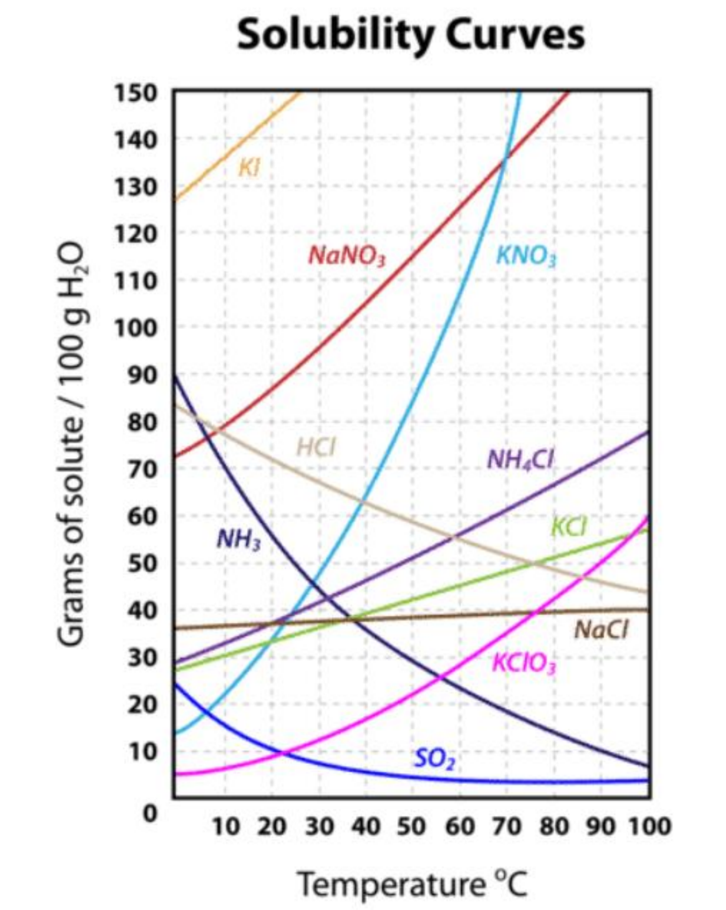


Figure 9.2.1.4: Solubility curves for several compounds.

Gas Solubility vs Temp

- Look at the solubility of HCl, NH_3 and SO_2

  Solubility of these goes down with increasing temperature

- He < O_2 << $CHCl_3$ in H_2O

  in this case, intermolecular forces such as molecular polarity and potential hydrogen bonding affect solubility

- If water temperature goes up, O_2 solubility is severely affected and could deny water life the gas it needs to survive (fish kills)

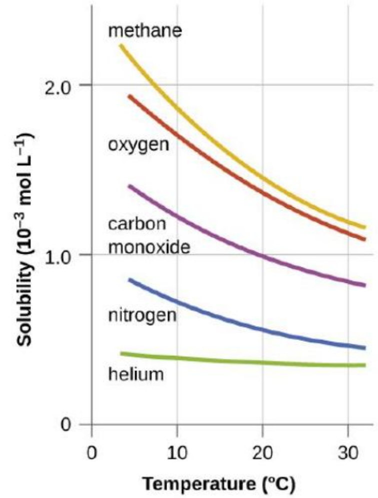


Figure 9.2.2.1: The solubilities of these gases in water decrease as the temperature increases. All solubilities were measured with a constant pressure of 101.3 kPa (1 atm) of gas above the solutions.

Figure 9.2.2.2: (a) The small bubbles of air in this glass of chilled water formed when the water warmed to room temperature and the solubility of its dissolved air decreased. (b) The decreased solubility of oxygen in natural waters subjected to thermal pollution can result in large-scale fish kills. (Credit a: modification of work by Liz West; credit b: modification of work by U.S. Fish and Wildlife Service.)

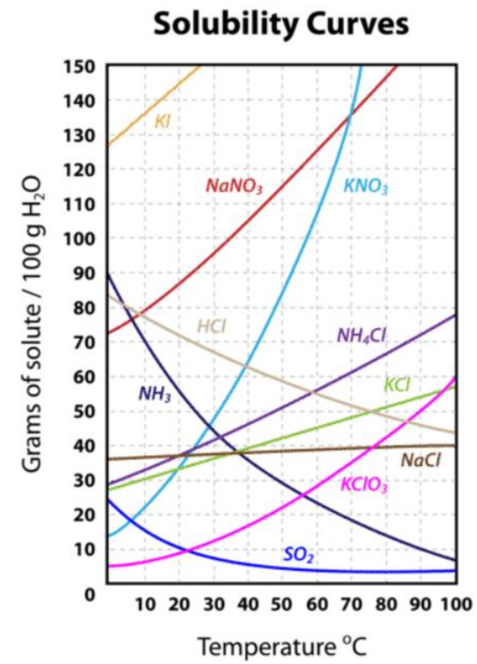


Figure 9.2.1.4: Solubility curves for several compounds.

Partial Pressure of Gas Solute

- Gas solutes are affected by partial pressure of the solute in the gas to which a solution is exposed
- If total gas pressure is increased, the solubility of gas in a solution increases

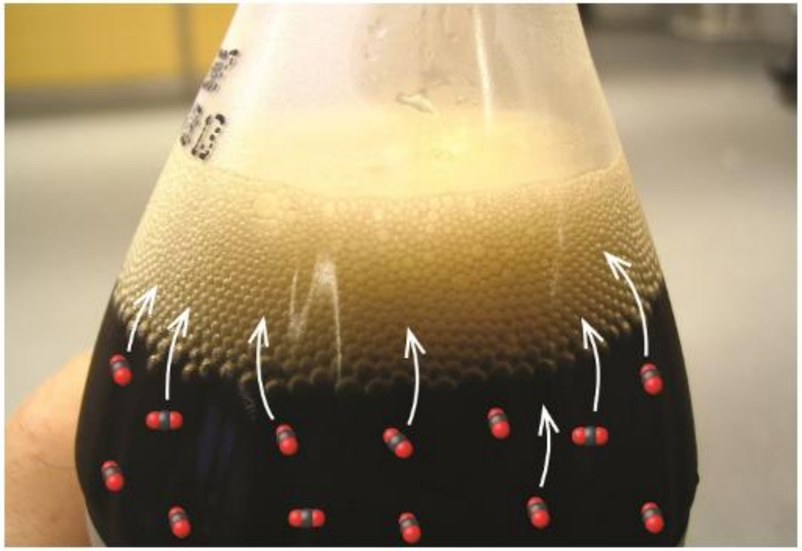


Figure 9.2.2.3: Opening the bottle of carbonated beverage reduces the pressure of the gaseous carbon dioxide above the beverage. The solubility of CO_2 is thus lowered, and some dissolved carbon dioxide may be seen leaving the solution as small gas bubbles. (Credit: modification of work by Derrick Coetzee.)

- Carbonated beverages are capped with high pressures of CO_2 gas
- When the cap is released, CO_2 rapidly leaves solution to enter atmosphere: fizzing, small bubbles
- This is a supersaturation effect of CO_2

Concentration

- A solution is composed of a solute and a solvent, and the amount of solute dissolved in the solvent is called a **concentration**

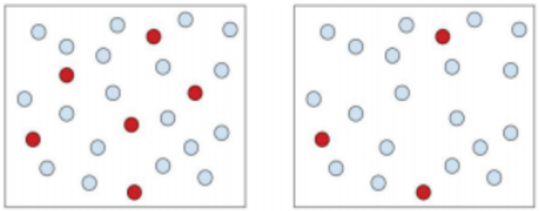


Figure 9.3.1: The solution on the left is more concentrated than the solution on the right because there is a greater ratio of solute (red balls) to solvent (blue balls) particles. The solution particles are closer together. The solution on the right is more dilute (less concentrated). (CC-SA-BY-3.0 Tracy Poulsen).

- A **5%** acetic acid (vinegar) in water is a concentration
- A **10%** acetic acid is *more* concentrated than a **5%** solution. The **5%** solution is *more* **dilute**

Percent Solutions

- Preliminary: a solution = solute + solvent

$$\text{solution mass }(g) = \text{solute mass }(g) + \text{solvent mass }(g)$$

- A concentration of a solution can be indicated by a mass of solute divided by mass of solution as a percentage

$$\text{percent by mass }(m/m)\% = \frac{\text{solute mass }(g)}{\text{solution mass }(g)} \times 100\%$$

- **25.0 g** sugar dissolved in **100 g** water is:

$$\text{percent by mass }(m/m)\% = \frac{25.0\text{ g}}{(25.0+100)\text{g}} \times 100\%$$

$$= 20.0\%\text{ sugar}$$

Mass Percent Solutions

- Compute the **mass** of NaCl solute to prepare **3000 g** of **5%** NaCl solution

$$\text{solute mass } (g) = \frac{\text{percent by mass } \left(\frac{m}{m}\right) \%}{100\%} \times \text{solution mass}$$

$$\frac{5\%}{100\%} \times 3000 \text{ g} = 150 \text{ g NaCl}$$

Volume Percent Solutions

- The **volume percent** of a solution is used when the solute and the solvent are **liquids**

$$\text{percent by volume } (v/v)\% = \frac{\text{solute volume } (mL)}{\text{solution volume } (mL)} \times 100\%$$

- Compute the **percent by volume** of **40 mL** ethanol (CH_3CH_2OH) solute to prepare **240 mL** ethanol solution

$$\text{percent by volume } (m/m)\% = \frac{40 \text{ mL}}{240 \text{ mL}} \times 100\%$$

$$= 16.7\% \text{ ethanol}$$

Molarity of Solutions

- The **molarity** of a solution is the **number of moles [mol]** of solute per **liter [L]** of the solution

$$\text{molarity } (M) = \frac{\text{solute moles (mol)}}{\text{solution volume (L)}}$$

- The molarity of **1.5 mol** NaCl in a final dissolution in **500. mL** solution is:

$$\frac{1.5 \text{ mol}}{500.\text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 3.0\ M \text{ NaCl}$$

Your book notes that phase indicator “(aq)” is added to “3.0 M NaCl” but this is unnecessary since “M” indicates molarity which is a concentration, implying a solute dissolved in a suitable usually water [aqueous] solvent

Practice: Molarity

- **42.23 g** NH_4Cl solute is dissolved in **500.0 mL** solution. Calculate its molarity.
- Find applicable mathematical relationship: $\text{molarity } (M) = \dfrac{\text{solute moles (mol)}}{\text{solution volume (L)}}$
- Identify associated inputs, determine needed conversions:
  - Solute is NH_4Cl; amount in grams (mass), but need moles. Have to use molar mass of NH_4Cl

$$42.23 \text{ g NH}_4\text{Cl} \times \frac{1 \text{ mol NH}_4\text{Cl}}{53.50 \text{ g NH}_4\text{Cl}} = 0.7893 \text{ mol NH}_4\text{Cl}$$

  - Solution volume is found but in units “mL” (milliliters), need to convert to “L” (liters)

$$500.0 \text{ mL} \times \frac{1 \text{ L solution}}{1000 \text{ mL solution}} = 0.5000 \text{ L solution}$$

Substitute values & solve:

$$42.23 \text{ g NH}_4\text{Cl} \times \frac{1 \text{ mol NH}_4\text{Cl}}{53.50 \text{ g NH}_4\text{Cl}} \times \frac{1}{500.0 \text{ mL}} \times \frac{1000 \text{ mL solution}}{1 \text{ L solution}}$$

$$= 1.579 \text{ M solution}$$

Note that the volume expression was inverted to show how conversions can be included in the entire calculation

Practice: Molarity

- **66.2 g** $C_6H_{12}O_6$ (glucose) solute is dissolved to prepare a **235 mL** solution. Calculate its molarity.
- Find applicable mathematical relationship: $\text{molarity } (M) = \dfrac{\text{solute moles (mol)}}{\text{solution volume (L)}}$
- Identify associated inputs, determine needed conversions:
  - Solute is $C_6H_{12}O_6$; amount in grams (mass), but need moles. Have to use molar mass of $C_6H_{12}O_6$

$$66.2\text{ g } C_6H_{12}O_6 \times \frac{1\text{ mol } C_6H_{12}O_6}{180.16\text{ g } C_6H_{12}O_6} = 0.367\text{ mol } C_6H_{12}O_6$$

  - Solution volume is found but in units "mL" (milliliters), need to convert to "L" (liters)

$$235\text{ mL} \times \frac{1\text{ L solution}}{1000\text{ mL solution}} = 0.235\text{ L solution}$$

Substitute values & solve:

$$\frac{0.367\text{ mol } C_6H_{12}O_6}{0.235\text{ L solution}} = 1.56\text{ M } C_6H_{12}O_6\text{ solution}$$

Note that the volume expression was inverted to show how conversions can be included in the entire calculation

Practice: Molarity

- **137 g** NaCl (sodium chloride) solute is dissolved to prepare a **500 mL** solution. Calculate its molarity.
- Find applicable mathematical relationship: $\text{molarity } (M) = \frac{\text{solute moles (mol)}}{\text{solution volume (L)}}$
- Identify associated inputs, determine needed conversions:
  - Solute is NaCl; amount in grams (mass), but need moles. Have to use molar mass of NaCl

$$137 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} = 2.34 \text{ mol NaCl}$$

  - Solution volume is found but in units "mL" (milliliters), need to convert to "L" (liters)

$$500 \text{ mL} \times \frac{1 \text{ L solution}}{1000 \text{ mL solution}} = 0.5 \text{ L solution}$$

Substitute values & solve:

$$137 \text{ g NaCl} \times \frac{1 \text{ mol NaCl}}{58.44 \text{ g NaCl}} \times \frac{1}{500 \text{ mL}} \times \frac{1000 \text{ mL solution}}{1 \text{ L solution}}$$

$$= 4.69 \text{ M NaCl solution} \quad \rightarrow \text{"5 M" according to significant digits}$$

Note that the volume expression was inverted to show how conversions can be included in the entire calculation

Mass = Concentration × Volume

- Just as $\text{concentration} = \dfrac{\text{mass or amount (mol)}}{\text{volume (L)}}$, algebra can be done to take any 2 of the 3 values to solve for the third value
- What **volume** (in liters) must be used of stock **2.35 M** $CuSO_4$ [copper(II) sulfate] solution to obtain **4.88 mol** solute?
- Find applicable mathematical relationship:

$$\text{molarity } (M) = \frac{\text{solute moles (mol)}}{\text{solution volume (L)}}$$

- Is algebra needed? $\text{solution volume (L)} = \dfrac{\text{solute moles (mol)}}{\text{molarity } (M)}$
- Substitute values & solve:

$$\text{solution volume (L)} = \frac{4.88\text{ mol CuSO}_4}{2.35\ M\ (\text{mol/L})\ \text{CuSO}_4} = 2.08\text{ L solution}$$

Mass = Molar Mass × Molarity × Volume

- A chemist must prepare **3.00 L** of a **0.250 M** $KMnO_4$ [potassium permanganate] solution. What **mass** of solute is needed
- Find applicable mathematical relationship:

$$\text{molarity } (M) = \frac{\text{solute moles (mol)}}{\text{solution volume (L)}}$$

- Is algebra needed? $\text{solute moles (mol)} = \text{solution volume (L)} \times \text{molarity } (M)$
- Conversions? need a mass in grams, so must relate moles ←→ grams, so a molar mass is required: for $KMnO_4$, **158.04 g/mol**
- Substitute values & solve:

$$\text{solute mass (g)} = 3.00 \text{ L solution} \times \frac{0.250 \text{ mol } KMnO_4}{1 \text{ L solution}} \times \frac{158.04 \text{ g } KMnO_4}{1 \text{ mol } KMnO_4}$$

$$= 119 \text{ g } KMnO_4$$

Volume = Moles (Mass) / Concentration

- What is the volume (L) required to make **0.0444 M** HCHO [formaldehyde] solution. Using **0.0773 mol** HCHO
- Find applicable mathematical relationship:

$$\text{molarity } (M) = \frac{\text{solute moles (mol)}}{\text{solution volume (L)}}$$

- Is algebra needed? $\text{solution volume (L)} = \frac{\text{solute moles (mol)}}{\text{molarity } (M)}$
- Conversions? none needed
- Substitute values & solve:

$$\text{solution volume (L)} = \frac{0.0773 \text{ mol HCHO}}{0.0444 \text{ M HCHO solution}} = 1.74 \text{ L solution}$$

Volume = Moles (Mass) / Concentration

- What is solute mass of **1.08 L** of a **0.0578 M** H_2SO_4 [sulfuric acid] solution?
- Find applicable mathematical relationship: $\text{molarity } (M) = \frac{\text{solute moles (mol)}}{\text{solution volume (L)}}$
- Is algebra needed? $\text{solute moles (mol)} = \text{molarity } (M) \times \text{solution volume (L)}$
- Conversions? Convert mole H_2SO_4 to grams H_2SO_4
- Substitute values & solve:

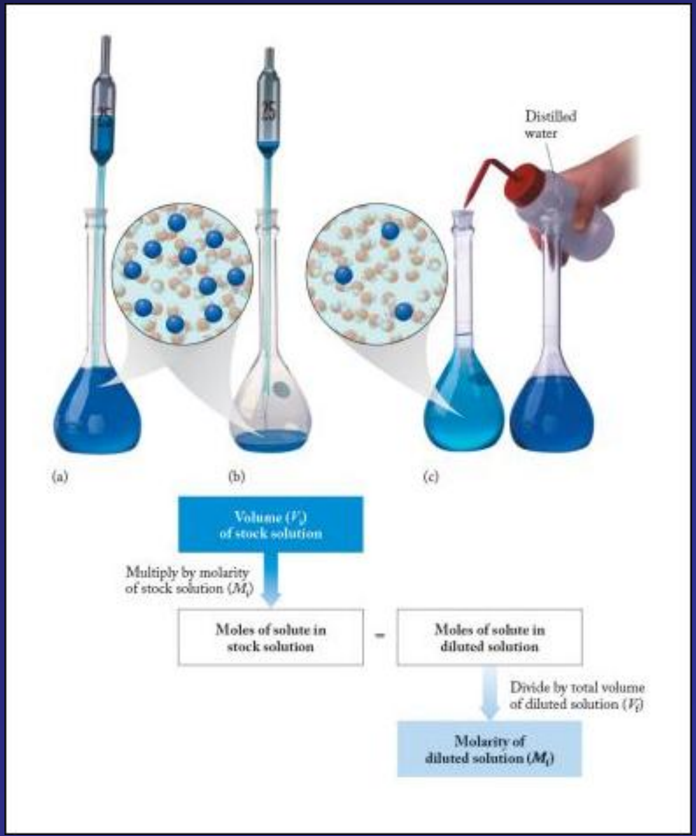
$$\text{solute mass (g)} = \frac{0.0578 \text{ mol } H_2SO_4}{1 \text{ L solution}} \times 1.08 \text{ L solution} \times \frac{98.09 \text{ g } H_2SO_4}{1 \text{ mol } H_2SO_4}$$
$$= 6.12 \text{ g } H_2SO_4$$

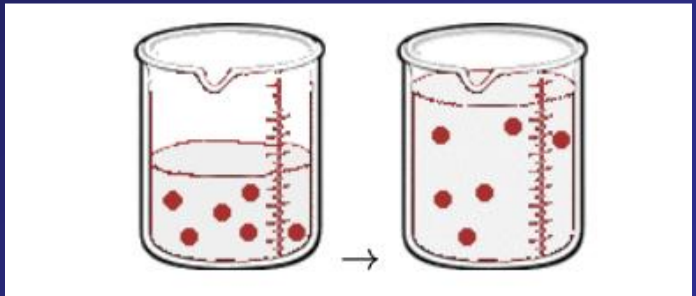
- What is volume of a **1.50 M** HCl [sulfuric acid] that contains **10.0 g** HCl?
- Find applicable mathematical relationship: $\text{molarity } (M) = \frac{\text{solute moles (mol)}}{\text{solution volume (L)}}$
- Is algebra needed? $\text{solution volume (L)} = \frac{\text{solute moles (mol)}}{\text{molarity } (M)}$
- Conversions? Must convert mass in grams HCl to moles HCl
- Substitute values & solve:

$$\text{solution volume (L)} = 10.0 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \times \frac{1 \text{ L solution}}{1.50 \text{ mol HCl}} = 0.183 \text{ L} = 183 \text{ mL}$$

Dilutions of Solutions

- Making stock solutions in aqueous solvent by dissolving solids in water is one thing
- Another very common practice is to make **dilute solutions** from the stock solutions
- The concentration (molarity) of the dilute solution must be less than the stock from which it is made





Before Dilution and After Dilution

The Math of Dilutions

- The math of making dilutions is a conservation of matter ("mass") equation actually

  moles of concentrated solution ***aliquot*** =
  moles of solute in diluted solution

- The mass of solute (in mol) of the solution = Concentration (in mol/L) × Volume (in L)
- So we relate $C_1V_1 = C_2V_2$ where
  C_1 = stock concentration
  V_1 = stock concentration volume used (aliquot)
  C_2 = (final) dilution concentration desired
  V_2 = (final) diluted volume desired
- In a dilution, diluted volume ALWAYS greater than stock volume ($V_2 > V_1$)

Practice: The Math of Dilutions

- Using a purchased **16 M** HNO_3 (nitric acid) stock solution, what volume is needed to prepare **8.00 L** of a (diluted) **0.50 M** HNO_3 solution?
- Find applicable mathematical relationship: $C_1V_1 = C_2V_2$
- Any algebra? $V_1 = \frac{C_2V_2}{C_1}$
- Substitute values & solve: This is a 16 M / 0.50 M = 32-fold dilution

$$V_1 = \frac{0.50\ M\ \times 8.00\ \text{L}}{16\ M} = 0.25\ \text{L} = 250\ \text{mL}$$

- A **76.5 mL** volume of **0.885 M** KBr is in a beaker. What volume water must be added using the appropriate graduated cylinder to make it **0.500 M** KBr?
- Find applicable mathematical relationship: $C_1V_1 = C_2V_2$
- Any algebra? $V_2 = \frac{C_1V_1}{C_2}$
- Substitute values & solve: This is a 0.885 M / 0.500 M = 1.77-fold dilution

$$V_2 = \frac{0.885\ M\ \times 76.5\ \text{mL}}{0.500\ M} = 135\ \text{mL}$$

It was not necessary to convert V_1 to liters since concentrations cancel

Colligative Properties: Molality

- **Colligative properties** are about what happens to a solvent at different solute concentrations

not related to the compound [type] itself

- This is about number of particles of solute to particles of solvent
  - When molecules dissolve, molecule number = particle number
  - When ionic compounds dissolve, ion number is twice or more the formula unit (particle) number
- A concentration (solute in the solution) called **molality** is useful here

$$\textbf{molality}\ (m) = \frac{\text{moles solute}}{\text{kilograms solvent}}$$

Note that the **mass** in KILOGRAMS of the solvent and NOT the solution is utilized in the definition

Practice: Molality

- **42.23 g** NH_4Cl (ammonium chloride) is dissolved in **500.0 g** of H_2O (water). What is the molality?
- Find applicable mathematical relationship:

$$\text{molality } (m) = \frac{\text{mole solute}}{\text{kg solvent}}$$

- Any algebra? Nope
- Any conversions?
  - Convert grams NH_4Cl to moles (molar mass = 53.50 g/mol)
  - Convert grams solvent to kilograms
- Substitute values & solve:

$$\frac{42.23 \text{ g NH}_4\text{Cl}}{500.0 \text{ g solvent}} \times \frac{1000 \text{ g solvent}}{1 \text{ kg solvent}} \times \frac{1 \text{ mol NH}_4\text{Cl}}{53.50 \text{ g NH}_4\text{Cl}} = 1.579 \ m$$

Boiling Point and Freezing Point

- Two properties solutes in a solvent affect are
  - increase (elevation) of the **boiling point** of the solvent
  - decrease (depression) of the **freezing point** of solvent
- Pure H_2O (water) at $\boldsymbol{P}$ = 1 atm
  - boils at 100°C
  - freezes at 0°C
- When table salt is added
  - the boiling point goes above 100°C
  - the freezing point falls below 0°C
- It does not matter what the solute is though
- 0.20 m NaCl and 0.20 m HCl would cause the boiling point and freezing point changes in same way

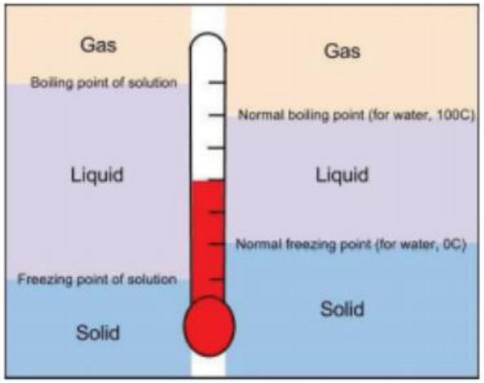


Figure 9.5.1.1: Comparison of boiling and freezing points of a pure liquid (right side) with a solution (left side).

Solute "Particle Number"

- Colligative properties are about number of solute particles
- When a molecule like glucose dissolves in water, glucose is a particle not different when dissolved than when a dry solid
- But when an ionic compound like NaCl dissolves, it creates two particles as ions: Na^+ and Cl^-
- When $CaCl_2$ dissolves, it creates three particles as ions: one Ca^{2+} and two Cl^- ions
- The molality of the solution and whether ionic or not ionic (i.e. molecular/covalent) is the indicator

1. Look for the number of ions produced when dissolved
2. Multiply molality (m) by number of particles (ions or molecular) to get total particle concentrations
3. Compare values of the solutions: higher total particle concentration will given higher b.p. and lower f.p. values

Effective Particle Molality

- Rank these solutions from higher b.p./lower f.p. values.
- 0.1 *m* NaCl
- 0.1 *m* $C_6H_{12}O_6$
- 0.1 *m* CaI_2

- NaCl → 2 particles (ions: Na^+, Cl^-)

  Effective particle molality = 2 x 0.1 *m* = ***0.2 m***
- $C_6H_{12}O_6$ → 1 particle (the molecule)

  Effective particle molality = 1 x 0.1 *m* = ***0.1 m***
- CaI_2 → 3 particles (ions: Ca^{2+}, two I^-)

  Effective particle molality = 3 x 0.1 *m* = ***0.3 m***

Computing B.P. / F.P. Changes

- The math for computing the changes to b.p. and f.p. values are known

$$\Delta T_b = kb \times \mathrm{m} \times \mathrm{i}$$
$$\Delta T_f = kf \times \mathrm{m} \times \mathrm{i}$$

ΔT_b, ΔT_f = change in boiling, freezing points, respectively

m = solution molality

i = number of particles formed in solution

i = 1 for molecules, i <= 2 for ions

k_b, k_f = a constant relating ΔT to molality and particle number

Table 9.5.1: Boiling and Freezing Constants of Selected Solvents

Substance	Chemical Formula	Normal Melting Point (°C)	Normal Boiling Point (°C)	K_f (°C/*m*)	K_b (°C/*m*)
Water	H_2O	0.0	100.0	1.86	0.512
Diethyl Ether	$C_4H_{10}O$	-116.3	34.5	1.76	2.02
Ethanol	C_2H_5OH	-114.1	78.3	1.99	1.22
Benzene	C_6H_6	5.50	80.1	5.10	2.53

Osmosis

- **Osmosis** is net movement of solvent molecules (typically water) across a selectively ("semi-") permeable membrane because solution concentrations differ across the membrane
- Solute is not permeable (does not pass through the membrane) but solvent molecules do pass
- The movement of solvent molecules is because a force (as a **pressure**) to equalize or the solute concentrations on both sides
- This can cause the volume of solution to be greater on one side of the membrane compared to the other

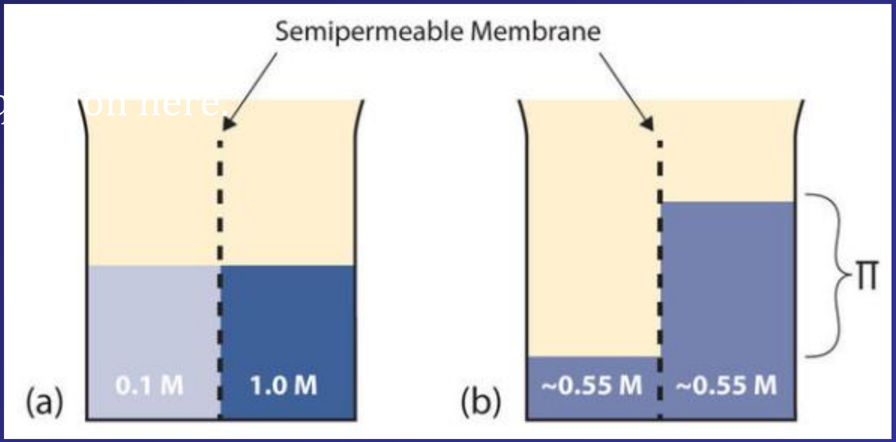


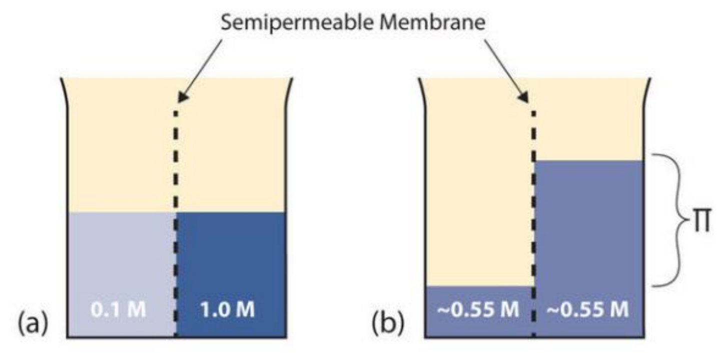
Figure: Osmosis. (a) Two solutions of differing concentrations are placed on either side of a semipermeable membrane. (b) When osmosis occurs, solvent molecules selectively pass through the membrane from the dilute solution to the concentrated solution, diluting it until the two concentrations are the same. The pressure exerted by the different height of the solution on the right is called the osmotic pressure. (CC BY-SA-NC 3.0; anonymous)

Osmotic Pressure

- **Osmotic pressure** (symbolized by a capital Greek letter **pi** Π, standing for a pressure) can be measured as a function of the difference in height difference of the levels of solution on both sides of the membrane
- Mathematically, osmotic pressure has the equation:
- $\Pi = MRT$

  Where ***M*** = solution molarity, ***R*** = ideal gas constant, and ***T*** = temperature in K

Osmosis explains what happens to red blood cells whether they remain normal, shrink/shrivel or swell & explode ("hemolyze") in saline solutions that are isotonic, hypertonic or hypotonic, respectively



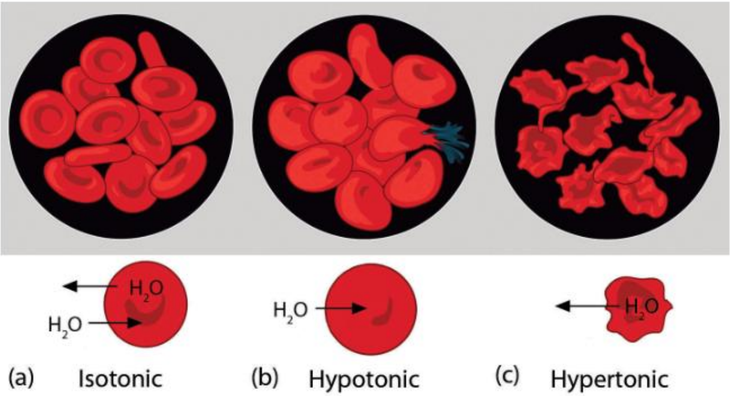


Figure 9.5.2.3: Osmotic Pressure and Red Blood Cells. (a) This is what a normal red blood cell looks like. (b) When a red blood cell is exposed to a hypotonic solution, solvent goes through the cell membrane and dilutes the inside of the cell. (c) When a red blood cell is exposed to a hypertonic solution, solvent goes from the cell to the surrounding solution, diluting the hypertonic solution and collapsing the cell. Neither of these last two cases is desirable, so IV solutions must be isotonic with blood serum to not cause deleterious effects. (Public Domain; Mariana Ruiz Villareal)

Practice: Osmotic Pressure

- What is the osmotic pressure of **0.333 M** $C_6H_{12}O_6$ (glucose) at 25°C
- Find applicable mathematical relationship: $\Pi = MRT$
- Algebra? Conversions?
  - Temperature must be converted from °C to K
- Substitute values & solve:

$$\Pi = MRT = 0.333\ \mathrm{M}\ \times 0.08205\ \mathrm{L\ atm/mol\ K} \times (25 + 273)\mathrm{K} = 8.14\ \mathrm{atm}$$

Note that L/mol is canceled by the molarity value (mol/L)

- What is the osmotic pressure of **0.0522 M** $C_{12}H_{22}O_{11}$ (glucose) at 25°C
- Find applicable mathematical relationship: $\Pi = MRT$
- Algebra? Conversions?
  - Temperature must be converted from °C to K
- Substitute values & solve:

$$\Pi = MRT = 0.0522\ \mathrm{M}\ \times 0.08205\ \mathrm{L\ atm/mol\ K} \times (55 + 273)\mathrm{K} = 1.40\ \mathrm{atm}$$