Slide 1

Colligative Properties

Practical uses of solutions

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Slide 2

Units of Concentration

Whatever units you use, the goal is the same: specify the quantity of 1 component (the solute$_s$) relative to the quantity of another component (the solvent).

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Slide 3

Common Units

% by mass

% by volume

Mole %

ppm (parts per million)

Molarity (M)

Molality (m)

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Slide 4

Common Units

$$\% \text{ by mass} = \frac{g\ solute}{100\ g\ solution}$$

$$\% \text{ by volume} = \frac{mL\ solute}{100\ mL\ solution}$$

$$mole\ \% = \frac{mol\ solute}{100\ mol\ solution}$$

$$ppm = \frac{g\ solute}{million\ g\ solution}$$

$$Molarity = \frac{mol\ solute}{L\ solution}$$

$$molality = \frac{mol\ solute}{kg\ SOLVENT}$$

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Slide 5

Solute as part of a solution

□ Note that, with the exception of molality, all of the units of concentration are expressed as some amount of solute compared to some amount of solution.

□ All the units of conversation are easily convertible, although sometimes you may need to know another piece of information (molar mass, density, etc.)

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Slide 6

Molarity

Molarity is probably the most common unit of concentration in chemistry.

Why is that?

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Slide 7

Molarity

Molarity is probably the most common unit of concentration in chemistry.

Why is that?

UNITS! UNITS! UNITS!

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Slide 8

Molarity

Molarity (M) = $\dfrac{\text{moles solute}}{\text{L solution}}$

This is both a chemically relevant unit and a practically relevant unit.

L solution is very easy to measure in the lab!

MOLES! MOLES! MOLES!

Reactions occur based on the relative number of moles.

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Slide 9

% by mass

This is one of the easiest units to use.

Why?

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Slide 10

% by mass

This is one of the easiest units to use.

Why?

UNITS! UNITS! UNITS!

Mass is easy to measure in the lab!

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Slide 11

% by mass

% by mass = $\frac{\text{g solute}}{100 \text{ g solution}}$

This is easy to use, but not as chemically relevant as Molarity. After all, it's all about the MOLES! MOLES! MOLES!

Is this a problem?

Not at all, it is easy to convert these units, but you need to know a few other things.

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Slide 12

Converting M to % by mass

% by mass = $\frac{\text{g solute}}{100 \text{ g solution}}$

M = $\frac{\text{moles solute}}{\text{L solution}}$

What do we need to know to be able to convert from one to the other?

UNITS! UNITS! UNITS! They shall lead the way.

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Slide 13

Converting M to % by mass

Compare the units:

g solute
100 g solution

Moles solute
Liter solution

What do you see?
Both numerators refer only to solute.
Both denominators refer to the whole solution.
PLEASE! PLEASE! PLEASE! Pay attention to the entire
    unit – it's not "grams" it's "grams of something"

Slide 14

Converting M to % by mass

g solute
100 g solution

Moles solute
Liter solution

If I want to convert the numerators, what do I need to do?

Slide 15

Converting M to % by mass

g solute   Moles solute
100 g solution   Liter solution

If I want to convert the numerators, what do I need to do?
(Do I need to say UNITS! UNITS! UNITS! again)?

g solute → moles solute

What's the conversion factor?

g solute   Or   moles solute
moles solute   g solute

What do we call this?

Molar mass!

Slide 16

Converting M to % by mass

g solute   x moles solute = moles solute
100 g solution   g solute   100 g solution

Or, if you're going the other way:

moles solute   x g solute =   g solute
L solution   moles solute   L solution

What about the denominator?

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Slide 17

Converting M to % by mass

moles solute   g solute
100 g solution   Liter solution

100 g solution → L solution

What's the conversion factor?

g solution OR   L solution
L solution   g solution.

What's this called?
Density! We need to know the density of the solution and the molar mass of the solute to be able to convert M to % by mass.

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Slide 18

Sample problem

You have a 1.2 M solution of NaCl in water. What is its % by mass given that the density of 1.2 M NaCl is 1.08 g/mL?

What do we do?

Simply convert the units!

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Slide 19

Sample problem

You have a 1.2 M solution of NaCl in water. What is its % by mass given that the density of 1.2 M NaCl is 1.08 g/mL?

$\dfrac{1.2 \text{ mol NaCl}}{\text{L solution}}$ → $\dfrac{\text{g NaCl}}{100 \text{ g solution}}$

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Slide 20

Sample problem

You have a 1.2 M solution of NaCl in water. What is its % by mass given that the density of 1.2 M NaCl is 1.08 g/mL?

$\dfrac{1.2 \text{ mol NaCl}}{\text{L solution}} \times \dfrac{58.453 \text{ g NaCl}}{\text{mol NaCl}} = \dfrac{70.143 \text{ g NaCl}}{\text{L solution}}$

$\dfrac{70.143 \text{ g NaCl}}{\text{L solution}} = ?\ \dfrac{\text{g NaCl}}{100 \text{ g solution}}$

We have density in g/mL, but we have volume in L, so...?

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Slide 21

Sample problem

You have a 1.2 M solution of NaCl in water. What is its % by mass given that the density of 1.2 M NaCl is 1.08 g/mL?

$\dfrac{1.2 \text{ mol NaCl}}{\text{L solution}} \times \dfrac{58.453 \text{ g NaCl}}{\text{mol NaCl}} = \dfrac{70.143 \text{ g NaCl}}{\text{L solution}}$

$\dfrac{70.143 \text{ g NaCl}}{\text{L solution}} \times \dfrac{1 \text{ L solution}}{1000 \text{ mL solution}} \times \dfrac{1 \text{ mL solution}}{1.08 \text{ g solution}} = \dfrac{0.06495 \text{ g NaCl}}{\text{g solution}}$

To get from g solution to 100 g solution???
Just multiply by 100.

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Slide 22

Sample problem

You have a 1.2 M solution of NaCl in water. What is its % by mass given that the density of 1.2 M NaCl is 1.08 g/mL?

$$\frac{1.2 \ mol \ NaCl}{L \ solution} \times \frac{58.453 \ g \ NaCl}{mol \ NaCl} = \frac{70.143 \ g \ NaCl}{L \ solution}$$

$$\frac{70.143 \ g \ NaCl}{L \ solution} \times \frac{1 \ L \ solution}{1000 \ mL \ solution} \times \frac{1 \ mL \ solution}{1.08 \ g \ solution} = \frac{0.06495 \ g \ NaCl}{g \ solution} * 100$$

6.495 g NaCl/100 g solution OR 6.495% by mass
  (6.5% by mass to the correct number of sig figs.)

Slide 23

Slightly more complicated problem

You have a 2.5 M solution of NaCl in water. What is its % by mass given that the density of 1.0 M NaCl is 1.07 g/mL, the density of 3.0 M NaCl is 1.10 g/mL and the density of 5.0 M NaCl is 1.13 g/mL?

This is a more common real-life example. There are tables of solution densities, but they don't have an infinite number of values. So, what do we do...?

Slide 24

Slightly more complicated problem

You have a 2.5 M solution of NaCl in water. What is its % by mass given that the density of 1.0 M NaCl is 1.07 g/mL, the density of 3.0 M NaCl is 1.10 g/mL and the density of 5.0 M NaCl is 1.13 g/mL?

2 options:
- Use the 3.0 M as being "close enough".
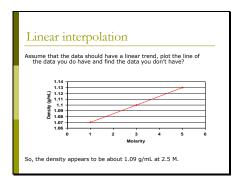- Do a linear interpolation to determine the approximate value at 2.5 M.

What is a "linear interpolation"?

Slide 25

Linear interpolation

Assume that the data should have a linear
trend, plot the line of the data you do have
and find the data you don't have?

Slide 26

Linear interpolation

Assume that the data should have a linear trend, plot the line of
the data you do have and find the data you don't have?

So, the density appears to be about 1.09 g/mL at 2.5 M.

Slide 27

The rest of the problem is the same.

You have a 2.5 M solution of NaCl in water. What is its % by
mass given that the density of 1.0 M NaCl is 1.06 g/mL, the
density of 3.0 M NaCl is 1.09 g/mL and the density of 5.0 M
NaCl is 1.13 g/mL?

$$\frac{2.5\ mol\ NaCl}{L\ solution} \times \frac{58.453\ g\ NaCl}{mol\ NaCl} = \frac{146.13\ g\ NaCl}{L\ solution}$$

$$\frac{146.13\ g\ NaCl}{L\ solution} \times \frac{1\ L\ solution}{1000\ mL\ solution} \times \frac{1\ mL\ solution}{1.09\ g\ solution} = \frac{0.134\ g\ NaCl}{g\ solution} * 100$$

$$= \frac{13.4\ g\ NaCl}{100\ g\ solution} = 13.4\%\ NaCl\ by\ mass$$

Slide 28

Colligative Property

A colligative property is a property that depends ONLY on the amount of the substance present NOT on the identity of the substance.

In other words, it doesn't matter if it is salt, sugar, gasoline, or tennis balls – it will behave the same way!

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Slide 29

Some Examples:

- Vapor Pressure Reduction
  - Related to boiling point
- Freezing Point Depression
  - Salt on the road
  - Anti-freeze in your radiator
- Boiling Point Elevation
  - Anti-freeze in your radiator
- Osmotic Pressure
  - Membrane diffusion
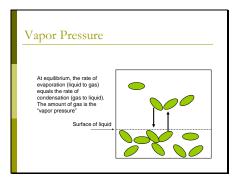 - The Great Sugar Fountain!
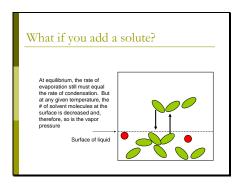
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Slide 30

Vapor Pressure Reduction

What is "vapor pressure"?

Vapor pressure is the amount (P is proportional to n for ideal gases) of gas A that is in equilibrium above the surface of liquid A.

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Slide 31

Vapor Pressure

At equilibrium, the rate of evaporation (liquid to gas) equals the rate of condensation (gas to liquid). The amount of gas is the "vapor pressure"

Surface of liquid

Slide 32

What if you add a solute?

At equilibrium, the rate of evaporation still must equal the rate of condensation. But at any given temperature, the # of solvent molecules at the surface is decreased and, therefore, so is the vapor pressure

Surface of liquid

Slide 33

Boiling Pt. Elevation
Freezing Pt. Depression

Birds of a Feather:

$\Delta T_b = T_{boiling, solution} - T_{boiling, pure solvent} = K_b\,m$

$\Delta T_f = T_{freezing, solution} - T_{freezing, pure solvent} = -K_f\,m$

m = molality of the solution
K_b = boiling constant
K_f = cryoscopic constant

Slide 34

NOTE ABOUT THE SIGN CONVENTION

$\Delta T_b = T_{\text{boiling, solution}} - T_{\text{boiling, pure solvent}} = K_b \, m$

$\Delta T_f = T_{\text{freezing, solution}} - T_{\text{freezing, pure solvent}} = - K_f \, m$

Δ = change
Δ = final state – initial state
Many texts do not put a minus sign before the cryscopic constant and reverse the definition of ΔT_f instead.

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Slide 35

K is a SOLVENT property

Colligative properties don't care what the solute is – they only care how much solute there is.

The boiling point constant and cryoscopic constant are SOLVENT properties only.

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Slide 36

A simple problem

How many grams of sucrose ($C_{11}H_{22}O_{11}$) are needed to lower the freezing point of 100 g of water by 3° C?

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Slide 37

The Answer

$$\Delta T_f = -K_f m$$

We want to decrease the freezing point by 3°C.

$$-3°C = -\left(1.86 \frac{°C}{molal}\right) m$$

$$m = 1.61 \; molal = \frac{1.61 \; mol \; solute}{kg \; water}$$
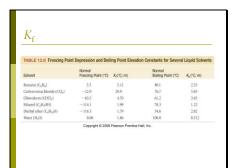
NOTE: K_f is the WATER cryoscopic constant

Slide 38

$C_{11}H_{22}O_{11}$ - sucrose

$$\frac{1.61 \; mol \; solute}{1 \; kg \; water} = \frac{x \; mol \; solute}{0.100 \; kg \; water}$$

Or, my preference is to use concentration as a conversion factor:

$$0.100 \; kg \; water \; \frac{1.61 \; mol \; sol}{1 \; kg \; water} = 0.161 \; mol \; sucrose \; \frac{342 \; g \; sucrose}{mol}$$

$$= 55.1 \; g \; sucrose$$

$$342 \frac{g}{mol} = 11 \times 12.011 \frac{g}{mol} + 22 \times 1.008 \frac{g}{mol} + 11 \times 16 \frac{g}{mol}$$

Slide 39

K_f

TABLE 12.8 Freezing Point Depression and Boiling Point Elevation Constants for Several Liquid Solvents

Solvent	Normal Freezing Point (°C)	K_f(°C/m)	Normal Boiling Point (°C)	K_b(°C/m)
Benzene (C_6H_6)	5.5	5.12	80.1	2.53
Carbon tetrachloride (CCl_4)	−22.9	29.9	76.7	5.03
Chloroform ($CHCl_3$)	−63.5	4.70	61.2	3.63
Ethanol (C_2H_5OH)	−114.1	1.99	78.3	1.22
Diethyl ether ($C_4H_{10}O$)	−116.3	1.79	34.6	2.02
Water (H_2O)	0.00	1.86	100.0	0.512

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Slide 40

Ex 12.8 – 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
Concept Plan:	m $\longrightarrow$ ΔT_f $\Delta T_f = m \bullet K_f$
Relationships	$\Delta T_f = m \cdot K_f$, K_f for H_2O = 1.86 °C/m, FP_{H2O} = 0.00°C
Solve:	$\Delta T_f = m \bullet K_{f,H_2O}$ $FP_{H_2O} - FP_{solh} = \Delta T_f$ $= (1.7\,m)\left(1.86\,\frac{°C}{m}\right)$ $0.00°C - FP_{solh} = 3.2°C$ $\Delta T_f = 3.2\,°C$ $FP_{solh} = -3.2°C$
Check:	the unit is correct, the freezing point lower than the normal freezing point makes sense

Slide 41

Colligative Properties

- **colligative properties** are properties whose value depends only on the number of solute particles, and not on what they are
  - Vapor Pressure Depression, Freezing Point Depression, Boiling Point Elevation, Osmotic Pressure
- the **van't Hoff factor, i,** is the ratio of moles of solute particles to moles of formula units dissolved
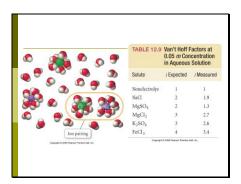- measured van't Hoff factors are often lower than you might expect due to ion pairing in solution

Slide 42

Ionic compounds dissociate…

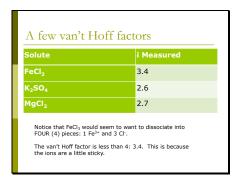…but not always completely.

Ions of opposite charge are always a little "sticky". So depending on just how sticky, an ionic compound may not completely dissociate.

The degree of dissociation is given by the "van't Hoff factor" – abbreviated "i"

Slide 43

TABLE 12.9 Van't Hoff Factors at 0.05 m Concentration in Aqueous Solution

Solute	i Expected	i Measured
Nonelectrolye	1	1
NaCl	2	1.9
MgSO₄	2	1.3
MgCl₂	3	2.7
K₂SO₄	3	2.6
FeCl₃	4	3.4

Ion pairing

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A few van't Hoff factors

Solute	i Measured
FeCl$_3$	3.4
K$_2$SO$_4$	2.6
MgCl$_2$	2.7

Notice that FeCl$_3$ would seem to want to dissociate into FOUR (4) pieces: 1 Fe^{3+} and 3 Cl$^-$.

The van't Hoff factor is less than 4: 3.4. This is because the ions are a little sticky.

Slide 45

An ionic compound in water

What is the freezing point of 1.6 g of MgCl$_2$ dissolved in 500 g of water?

$$\Delta T_f = -iK_f m$$

$$\Delta T_f = -2.7 \left(1.86 \, \frac{°C}{m} \right) \frac{1.6 \, g \, \frac{1 \, mol}{95.21 \, g}}{0.500 \, kg} = -0.169$$

$$T_f - 0°C = -0.169°C$$

$$T_f = -0.169°C$$

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Slide 46

Boiling Point Elevation

What is the "boiling point" of a liquid?

It is the temperature at which the vapor pressure equals the atmospheric pressure. So…

…if you decrease the vapor pressure, you must increase the boiling point – it will take a higher temperature to get enough gas molecules (vapor pressure) to equal the atmospheric pressure.
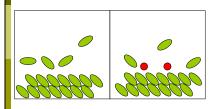
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Slide 47

What is the Microscopic View of Freezing Point Depression?

These phase change properties arise from the attractive interactions between all molecules (Van der Waal's forces).

When you make a solution, you break some of the solvent-solvent interactions and replace them with solvent-solute interactions.
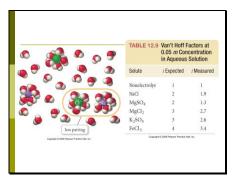
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Slide 48

Freezing is a dynamic equilibrium between melting and freezing.

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Slide 49

Slight Variation on the problem

How many grams of NaCl are needed to lower the freezing point of 100 g of water by 3 °C?

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Slide 50

TABLE 12.9 Van't Hoff Factors at 0.05 *m* Concentration in Aqueous Solution

Solute	*i* Expected	*i* Measured
Nonelectrolye	1	1
NaCl	2	1.9
$MgSO_4$	2	1.3
$MgCl_2$	3	2.7
K_2SO_4	3	2.6
$FeCl_3$	4	3.4

Ion pairing

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Slide 51

The Answer

$$\Delta T_f = -iK_f m$$

We want to decrease the freezing point by 3°C.

$$-3°C = -1.9\left(1.86\ \frac{°C}{molal}\right)m$$

$$m = 0.849\ molal = \frac{0.849\ mol\ solute}{kg\ water}$$

NOTE: K_f is the WATER cryoscopic constant

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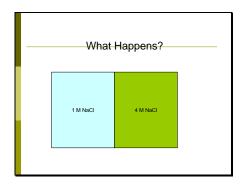
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Slide 52

NaCl

$$0.100 \; kg \; water \; \frac{0.849 \; mol \; sol}{1 \; kg \; water} = 0.0849 \; mol \; NaCl \frac{58.45 \; g \; NaCl}{mol}$$
$$= 4.96 \; g \; NaCl$$

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Slide 53

It's all about the # of particles

NaCl is an electrolyte:

$NaCl \rightarrow Na^+ + Cl^-$

You get 2 moles of solute per mole NaCl (well, 1.9 because of ion pairing)

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Slide 54

Osmotic Pressure

□ Osmotic pressure is the "funky" colligative property, but it is very important biologically

□ Osmotic pressure is the pressure required to overcome the natural pressure exerted by a solution by virtue of having a concentration.

□ Osmotic pressure looks just like the ideal gas law:

$\Pi V = nRT$

where Π is osmotic pressure

$\Pi = (n/V) \; RT = M \; RT$

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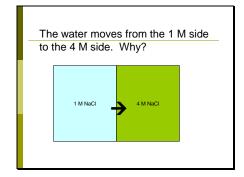
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Slide 55

Why is it important?

Consider two solutions separated by a semi-permeable membrane (meaning water can pass through but nothing else) at room temperature (298 K).

On one side of the membrane is a 1 M NaCl solution. On the other side is a 4 M NaCl solution.

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Slide 56

What Happens?

1 M NaCl | 4 M NaCl

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Slide 57

The water moves from the 1 M side to the 4 M side. Why?

1 M NaCl → 4 M NaCl

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Slide 58

OSMOTIC PRESSURE
   Π= iM RT

   On the 1 M side, the pressure is:
      Π= M RT = (1.9) 1 M (0.0821 Latm/mol K)(298 K)
         Π=48 atm

   On the 4 M side, the pressure is:
      Π= iM RT =(1.9) 4 M (0.0821 Latm/mol K) (298 K)
         Π= 190 atm

Slide 59

Both solutions "push" on the membrane
The bigger push wins!

48 atm

1 M NaCl   4 M NaCl

190 atm

142 atm

Note the direction of the arrows. Osmotic pressure is pushing AGAINST the solution.

Slide 60

It's all about sweating!

□ What happens when you sweat?

Slide 61

It's all about sweating!

What happens when you sweat?

Your body loses water and some dissolved salts through your skin which cools your body.

What is the price you pay for this?

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Slide 62

The cost of sweating

Your cells have lost precious water and salts.

What do you do about it?

Drink water?

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Slide 63

The problem with water

Your cells are slightly dehydrated from sweating. The salt concentration is actually higher than it should be.

Water has no salt concentration. It flows into the cells rapidly (osmotic pressure) and fills the cells. But the cells are now diluted.

Further, since the pure water will always have lower osmotic pressure and flow into the cell until either you run out of water or the cell bursts.

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Slide 64

Gatorade – Man's answer to Osmotic Pressure

Gatorade is flavored sweat!

By adding salts, it matches its osmotic pressure to what the cell should be. This allows for smooth rehydrating of the cell without dilution or lysing.

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Slide 65

EXAMPLE 12.4 Calculating Concentrations

A solution is prepared by dissolving 17.2 g of ethylene glycol ($C_2H_6O_2$) in 0.500 kg of water. The final volume of the solution is 515 mL. For this solution, calculate each of the following:

(a) molarity   (b) molality   (c) percent by mass   (d) mole fraction   (e) mole percent

continued…

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Slide 66

(c) To calculate percent by mass, divide the mass of the solute by the sum of the masses of the solute and solvent and multiply the ratio by 100%.

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Slide 67

EXAMPLE 12.5 Calculating the Vapor Pressure of a Solution Containing a Nonionic and Nonvolatile Solute

Calculate the vapor pressure at 25 °C of a solution containing 99.5 g sucrose ($C_{12}H_{22}O_{11}$) and 300.0 mL water. The vapor pressure of pure water at 25 °C is 23.8 torr. Assume the density of water to be 1.00 g/mL.

Slide 68

Solve Calculate the number of moles of each solution component.

Slide 69

EXAMPLE 12.6 Calculating the Vapor Pressure of a Solution Containing an Ionic Solute

A solution contains 0.102 mol $Ca(NO_3)_2$ and 0.927 mol H_2O. Calculate the vapor pressure of the solution at 55 °C. The vapor pressure of pure water at 55 °C is 118.1 torr.

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Solve The key to this problem is to understand the dissociation of calcium nitrate. Write an equation showing the dissociation.

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EXAMPLE 12.9 Boiling Point Elevation

How much ethylene glycol ($C_2H_6O_2$), in grams, must be added to 1.0 kg of water to produce a solution that boils at 105.0 °C?

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Solve Begin by solving the boiling point elevation equation for molality and substituting the required quantities to compute m.

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EXAMPLE 12.11 Van't Hoff Factor and Freezing Point Depression

The freezing point of an aqueous 0.050 m CaCl$_2$ solution is −0.27 °C. What is the van't Hoff factor (i) for CaCl$_2$ at this concentration? How does it compare to the predicted value of i?