Dissolving LiCl (s):

\[
\text{LiCl (s)} \rightarrow \text{Li}^+ (g) + \text{Cl}^- (g) \quad \Delta H = 853 \text{ kJ/mol}
\]

\[
\text{Li}^+ (g) + \text{Cl}^- (g) + \text{H}_2\text{O (l)} \rightarrow \text{Li}^+ (aq) + \text{Cl}^- (aq) \quad \Delta H = -890 \text{ kJ/mol}
\]

\[
\text{LiCl (s)} + \text{H}_2\text{O (l)} \rightarrow \text{Li}^+ (aq) + \text{Cl}^- (aq) \quad \Delta H_{\text{soln}} = -37 \text{ kJ/mol}
\]

Note: some soluble salts have an endothermic heat of solution, like KCl (\(\Delta H_{\text{soln}} = +17.2 \text{ kJ/mol}\)). CHEM 2 explains why KCl is soluble despite having a positive \(\Delta H_{\text{soln}}\).
MOLECULAR COMPOUNDS

Like Dissolves Like Principle:

If solute and solvent are “like” each other in polarity, they dissolve.

If they are “unlike” each other, they don’t dissolve.
Which of these should be soluble in water?

- Hexachlorobenzene, probable human carcinogen, previously used fungicide on wheat
- Furan, possible human carcinogen, small # in roasted coffee
- Ascorbic acid, Vitamin C
- Vitamin E
GAS SOLUBILITY IN LIQUIDS

Henry’s law
\[ C = k_H \cdot P \]
- \( C \) = concentration of dissolved gas
- \( k_H \) = Henry’s law constant for a solution
- \( P \) = partial pressure of gas solute above the solution

Sample Problem:

Calculate the concentration of \( \text{CO}_2 \)(aq) in 25 °C water, when the atmosphere has a \( P_{\text{CO}_2} \) of \( 4.0 \times 10^{-4} \) atm. The Henry’s law constant for \( \text{CO}_2 \) in water is 0.031 \( M/\text{atm} \) at 25 °C.

Water Solubility of Gases at 25 °C

<table>
<thead>
<tr>
<th>Gas</th>
<th>Water Soluble</th>
<th>g / 100 g water</th>
<th>Gas</th>
<th>Water Soluble</th>
<th>g / 100 g water</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{H}_2 )</td>
<td>N</td>
<td>0.00016</td>
<td>( \text{CO}_2 )</td>
<td></td>
<td>0.169</td>
</tr>
<tr>
<td>( \text{O}_2 )</td>
<td>N</td>
<td>0.0043</td>
<td>( \text{SO}_2 )</td>
<td></td>
<td>11.28</td>
</tr>
<tr>
<td>( \text{N}_2 )</td>
<td>N</td>
<td>0.0019</td>
<td>( \text{NH}_3 )</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

What can you do to increase the concentration of dissolved \( \text{CO}_2 \)?

Sample Problem:

Calculate the solubility of \( \text{CO}_2 \)(aq) at 25 °C, when bottled with a \( \text{CO}_2 \) pressure of 5.0 atm over the liquid.
Sample Problems:

At 1 atm, the partial pressure of N$_2$ is 0.7809 atm (as 78.09% of air is N$_2$). The solubility of N$_2$ at 25 °C is 0.00055 M.

a. Calculate the Henry’s law constant for N$_2$ at 25 °C.

b. At 3 atm, the P$_{N2} = 3 \times 0.7809$ atm = 2.34 atm. Calculate the concentration of N$_2$ (aq) at this pressure.

Scuba Diving and “The Bends”

GAS SOLUBILITY + TEMPERATURE

- Gas solubility
- Pouring cold vs. warm soda
- Lake cooling of nuclear power plants
CONCENTRATION UNITS

DEFINITIONS

Concentration = \frac{\text{quantity of smaller \# (e.g. solute)}}{\text{larger quantity present (e.g. total)}}

<table>
<thead>
<tr>
<th>Molarity (M)</th>
<th>Molality (m)</th>
<th>Mass Percent (%)</th>
<th>Mole Fraction (\chi)</th>
</tr>
</thead>
<tbody>
<tr>
<td>M = \frac{\text{moles (solute)}}{L (solution)}</td>
<td>m = \frac{\text{moles (solute)}}{kg (solvent)}</td>
<td>% = \frac{\text{mass}_A}{\text{total mass}} \times 100</td>
<td>\chi_A = \frac{\text{moles}_A}{\text{total moles}}</td>
</tr>
</tbody>
</table>

Sample Problem:

Industrial strength hydrogen peroxide is 30.0\% \text{H}_2\text{O}_2\text{ by mass and has a density of 1.11 g/mL. Calculate the molarity, molality, and mole fraction of \text{H}_2\text{O}_2 in this aqueous solution.}

a. Molarity

b. Molality

c. Mole fraction of \text{H}_2\text{O}_2
**COLLIGATIVE PROPERTIES OF SOLUTIONS**

**Colligative Properties:** only depend on the concentration of solute, not their identity.

**VAPOR PRESSURE OF SOLUTIONS**

All liquids / solutions at 20 °C:

<table>
<thead>
<tr>
<th>Solution</th>
<th>Vapor Pressure (torr)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td>17.5</td>
</tr>
<tr>
<td>2.97 M sugar (aq)</td>
<td>16.6</td>
</tr>
<tr>
<td>2.97 M ethylene glycol (aq)</td>
<td>16.6</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Raoult’s law</th>
<th>P_{soln} = \chi_{solvent} \cdot P^*_{solvent}</th>
</tr>
</thead>
<tbody>
<tr>
<td>( P_{soln} )</td>
<td>vapor pressure of the solution ( (\text{meaning \ # \ of \ solvent \ in \ gas \ phase}) )</td>
</tr>
<tr>
<td>( \chi_{solvent} )</td>
<td>mole fraction of solvent</td>
</tr>
<tr>
<td>( P^*_{solvent} )</td>
<td>vapor pressure of the pure solvent</td>
</tr>
</tbody>
</table>

**Sample Problems:**

Calculate the vapor pressure of a 2.97 \( M \) sugar solution (mole fraction = 0.0516 sugar) at 20.0 °C. The vapor pressure of water at 20.0 °C is 17.552 torr.

Which has the lowest vapor pressure?
Sample Problem:

Calculate the vapor pressure of an aqueous solution that contains 25.2 g urea (CH₄N₂O, a non-volatile substance, MM = 60.07) in 15.0 mol water at 25 °C. The vapor pressure of water at 25 °C is 23.76 mmHg.

**VAN’T HOFF FACTOR**

All liquids / solutions at 25 °C:

- Water: $P_{\text{water}} = 23.76$ torr
- $3.7 \text{ M glucose (aq)}$: $P_{\text{water}} = 22.3$ torr
- $3.7 \text{ M NaCl (aq)}$: $P_{\text{water}} = 21.0$ torr

Van’t Hoff Factor ($i$) = Number of actual particles in solution

<table>
<thead>
<tr>
<th>Solution</th>
<th>$i$ (expected)</th>
<th>$i$ (actual)</th>
</tr>
</thead>
<tbody>
<tr>
<td>C₆H₁₂O₆ (aq)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaCl (aq)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>MgCl₂ (aq)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>MgSO₄ (aq)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Boiling occurs when…

Which has the highest boiling point?

- Water
- 2.0 M KCl (aq)
- 2.0 M Ca(NO₃)₂ (aq)

**Boiling Point Elevation**

\[ \Delta T_b = i \cdot k_b \cdot m \]

- \( \Delta T_b \) = increase in boiling point of solution (compared to pure solvent)
- \( i \) = Van’t Hoff factor (# solute particles)
- \( k_b \) = boiling point constant for a solvent
- \( m \) = molality of solution (moles solute / kg solvent)

**Sample Problem:**

Benzene \((C_6H_6)\) has a normal boiling point of 80.1 °C. What would be the boiling point of a 0.186 molal solution of Vitamin A \((C_{20}H_{30}O, \text{a non-volatile molecular compound})\) in benzene? The boiling point constant \(k_b\) for benzene is 2.53 °C·kg/mol.
Sample Problem:

How many grams of NaCl (MM = 58.44) must be dissolved in 2.0 L water so that the boiling point of the solution is 101.0 °C?

The density of water is 0.997 g/mL and $k_b$ for water is 0.5121 °C·kg/mol.

FREEZING POINT DEPRESSION

<table>
<thead>
<tr>
<th></th>
<th>Water</th>
<th>$1\ m$ glucose (C$<em>6$H$</em>{12}$O$_6$)</th>
<th>$1\ m$ KCl</th>
<th>$1\ m$ CaCl$_2$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Freezing point</td>
<td>0.00 °C</td>
<td>−1.86</td>
<td>−3.29</td>
<td>−6.35</td>
</tr>
</tbody>
</table>

Salt applied to icy roads melts the ice
Freezing Point Depression

\[ \Delta T_f = i \cdot k_f \cdot m \]

\( \Delta T_f \) = decrease in freezing point of solution (compared to pure solvent)

\( i \) = Van’t Hoff factor (# solute particles)

\( k_f \) = freezing point constant for a solvent

\( m \) = molality of solution (moles solute / kg solvent)

Sample Problem:

Prestone antifreeze is 50% water and 50% ethylene glycol (C\(_2\)H\(_6\)O\(_2\), non-volatile, MM = 62.07).

Calculate the freezing point of a 45.0% ethylene glycol, 55.0% water mixture. The \( k_f \) for water is 1.86 °C·kg/mol.

\[ HO\overset{\cdot}{\backslash}CH_2\overset{\cdot}{\backslash}CH_2\overset{\cdot}{\backslash}OH \]

ethylene glycol

Which has the highest melting point?

- Water
- 0.2 \( M \) CsCl (aq)
- 0.3 \( M \) C\(_{12}\)H\(_{22}\)O\(_{11}\) (aq)

Ice cream softens at –14 °C
Freezer temp is often –18 °C
**Sample Problem:**

Benzene ($C_6H_6$) has a normal freezing point of 5.53 °C. What mass of methyl stearate ($C_{19}H_{38}O_2$, a non-electrolyte, MM = 298.51) must be dissolved in 3.20 kg of benzene to give a solution with a freezing point of –2.00 °C? The freezing point constant $k_f$ for benzene is 5.12 °C·kg/mol.

---

**OSMOSIS + OSMOTIC PRESSURE**

**DEFINITIONS**
Osmotic pressure

\[ \Pi = i \cdot M \cdot R \cdot T \]

- \( \Pi \) = osmotic pressure (in atm)
- \( i \) = Van’t Hoff factor (# solute particles)
- \( M \) = molarity (mol/L) of solute
- \( R \) = gas constant (0.08206 L·atm/mol·K)
- \( T \) = temperature (in K)

**Sample Problem:**

A 0.158 M NaCl \((aq)\) solution is “isotonic” with blood at 25 °C. Calculate the osmotic pressure of blood.

**APPLICATIONS**

- Normal red blood cell
- Red blood cell in pure water: water flows into cell
- Red blood cell in concentrated solution: water flows out of cell
- Drinking seawater
- Preserving meat through salting
**Sample Problem:**

20.0 mg insulin (a non-dissociating protein) is dissolved in 5.00 mL water. The osmotic pressure of the solution is measured to be 12.5 mmHg at 27 °C. Calculate the molar mass of insulin.

---

**REVERSE OSMOSIS**

Seawater has an osmotic P of 25.6 atm