## Solution Chemistry

## Nature of Solutions

- solutions are homogeneous mixtures
- substances in solution are different from their solid, liquid or gas forms
- there should be no observable segregation of component substances


## Types of Solutions

1. Gaseous solutions

Ex. Air - $78 \% \mathrm{~N}_{2}, 21 \% \mathrm{O}_{2}, 1 \% \mathrm{Ar}, 0.03 \% \mathrm{CO}_{2}$
2. Solid solutions (alloys)

Ex. Sterling silver ( $\mathrm{Cu}+\mathrm{Ag}$ ), brass ( $\mathrm{Cu}+\mathrm{Zn}$ ),
Dental amalgam ( $\mathrm{Hg}+\mathrm{Ag}$ ), steel $(\mathrm{Fe}+\mathrm{C})$
3. Liquid solutions:
a. liquid-liquid
Ex. vinegar
b. solid - liquid
Ex. salt water
c. gas - liquid
Ex. carbonated water in soda pop

- The substance present in the greater amount is the solvent
- The solute is the component of lesser abundance


## Polar versus Non-polar Solvents

## Molecular polarity

A dipole results when electrons are displaced to one side of the molecule Ex. $\mathrm{H}_{2} \mathrm{O}$ - polar
$\mathrm{CCl}_{4}$ - non-polar

## Polar versus Non-polar

Polar solvents tend to dissolve polar (and ionic) solutes
Non-polar solvents tend to dissolve non-polar solutes

## Water

Solutions containing water are called aqueous (aq)
Many important reactions (Ex. in living things) take place in aqueous solutions (Probably) the most common solvent

## Saturated and Unsaturated Solutions

- Saturated - solution in which no additional solute can dissolve
- Unsaturated - a solution in which additional solute can dissolve
- Miscible - when there is no apparent limit to the solubility of one liquid in another
- Immiscible - two liquids which do not dissolve in each other; solvation (interaction between solute and solvent particles) does not occur
- Solubility - is a measure of the amount of solute which can dissolve in a given amount of solvent
o $\mathrm{g} / \mathrm{L}, \mathrm{g} / \mathrm{mL}, \mathrm{g} / \mathrm{g}, \mathrm{mL} / \mathrm{mL}$, are some units for this
o adding a little at a time until saturated; or evaporating a saturated solution and measuring the mass of the solid left behind, are ways to determine solubility
o Solubility Table defines soluble as $>0.1 \mathrm{~mol} / \mathrm{L} @ 250 \mathrm{C}$


## Concentration of Solutions

Concentration describes the amount of solute in a given amount of solutions.
a) \% solution $\quad \mathbf{X g} / 100 \mathrm{~g}$ of solution (W/W) $\mathbf{X ~ m L} / 100 \mathrm{~mL}$ of solution (V/V)
b) Molarity $\quad \rightarrow \quad$ \# of moles of solute litres of solution made

$$
\begin{array}{ll}
M=n / V & n=\# \text { mols, } V=\text { volume }(L) \\
n=M \times V & \\
V=n / M &
\end{array}
$$

c) Dilute versus Concentrated

- dilute solutions contain less solute per litre of solution than concentrated solutions

Ex. Dilute $\mathrm{HCl}: 0.1 \mathrm{M}$ or 0.5 M
Concentrated HCl : 12M

## Dilution Calculations

Often solutions are made by diluting a more concentrated (stock) solution

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\(\mathbf{n}=\mathbf{M} \mathbf{x} \mathbf{V}\)
Stock sol'n (conc.): \(n_{1}=M_{1} \times V_{1}\)
New sol'n (dilute): \(n_{2}=M_{2} \times V_{2}\)
\(\mathrm{n}_{1}=\mathrm{n}_{2}\)
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So...
$\mathbf{M}_{\mathbf{1}} \times \mathbf{V}_{\mathbf{1}}=\mathbf{M}_{\mathbf{2}} \times \mathbf{V}_{\mathbf{2}}$
(\# moles does not change, only M does!)

Ex. If you have 500.0 mL of 0.40 M solution, what volume of a 0.10 M solution can you make? How much water must be added?
\# moles =
Vol. solution to be made $=$ Vol. $\mathrm{H}_{2} \mathrm{O}$ to be added $=$

OR
$M_{1} \times V_{1}=M_{2} \times V_{2}$

$$
V_{2}=
$$

Ex.
What final concentration results when 150 mL of 0.36 M solution of $\mathrm{MgSO}_{4}$ is added to 750 mL of water?
moles $\mathrm{MgSO}_{4}=$
final vol. = molarity $=$

OR

Ex. What volume of $0.50 \mathrm{M} \mathrm{CuCl}_{2}$ stock solution is required to make 2.0 L of $0.20 \mathrm{M} \mathrm{CuCl}_{2}$ ?
mols $\mathrm{CuCl}_{2}$ needed $=$
vol. $=\mathrm{n} / \mathrm{M}=$

$$
=
$$

OR
$0.50 \mathrm{M}\left(\mathrm{V}_{1}\right)=(0.20 \mathrm{M})(2.0 \mathrm{~L})$
so $\mathrm{V}_{1}=$
vol. of $\mathrm{H}_{2} \mathrm{O}$ to be added $=$

## Aqueous solutions of ionic substances

Ionic substances dissolve in water to form ions
Ex. $\mathrm{NaCl}_{(\mathrm{s})} \rightarrow \mathrm{Na}^{+}{ }_{(\text {aq })}+\mathrm{Cl}^{-}{ }_{(\text {aq) }}$ salt ion ion

The concentration of ions can be different from the concentrations of the salts
Ex. $\mathrm{CaCl}_{2(\mathrm{~s})} \rightarrow \mathrm{Ca}^{2+}{ }_{(\mathrm{aq})}+2 \mathrm{Cl}^{-}{ }_{(\mathrm{aq})}$
$\left[\mathrm{CaCl}_{2}\right]=$
$\left[\mathrm{Ca}^{2+}\right]=$
[Cl'] =
This is because every mole of $\mathrm{CaCl}_{2}$ dissolves to give one mole of $\mathrm{Ca}^{2+}$ ions and two moles of $\mathrm{Cl}^{-}$ions.

## Examples:

Write a dissociation equation for each of the following and give the concentration of the ions:

Ex. $\quad 0.10 \mathrm{M} \mathrm{Al}\left(\mathrm{NO}_{3}\right)_{3}$

$$
\mathrm{Al}\left(\mathrm{NO}_{3}\right)_{3} \rightarrow \mathrm{Al}^{3+}+3 \mathrm{NO}_{3}^{-}
$$



Ex. $\quad 0.35 \mathrm{M} \mathrm{Cu}\left(\mathrm{CH}_{3} \mathrm{COO}\right)_{2}$

$$
\mathrm{Cu}\left(\mathrm{CH}_{3} \mathrm{COO}\right)_{2} \rightarrow \mathrm{Cu}^{2+}+2 \mathrm{CH}_{3} \mathrm{COO}^{-}
$$

$\left[\mathrm{Cu}^{2+}\right]=$
[ $\left.\mathrm{CH}_{3} \mathrm{COO}^{-}\right]=$
Ex. $\quad 0.25 \mathrm{M} \mathrm{Na}_{2} \mathrm{O}$

$$
\mathrm{Na}_{2} \mathrm{O} \rightarrow 2 \mathrm{Na}^{+}+\mathrm{O}^{2-}
$$

$\left[\mathrm{Na}^{+}\right]=$
$\left[\mathrm{O}^{2-}\right]=$

## Equation types

Dissociation Eq ${ }^{\text {n }}$
$\mathrm{Pb}\left(\mathrm{NO}_{3}\right)_{2(\mathrm{~s})} \rightarrow \mathrm{Pb}^{2+}{ }_{(\text {aq })}+2 \mathrm{NO}_{3}{ }^{-{ }^{-}(\mathrm{aq})}$
Formula Eq ${ }^{\text {n }}$
$\mathrm{Pb}\left(\mathrm{NO}_{3}\right)_{2(\mathrm{~s})}+2 \mathrm{KI}_{(\mathrm{aq})} \rightarrow 2 \mathrm{KNO}_{3(\mathrm{aq})}+\mathrm{PbI}_{2(\mathrm{~s})}$
Complete (Full) Ionic Eq ${ }^{\text {n }}$
$\mathrm{Pb}^{2+}{ }_{(\mathrm{aq})}+2 \mathrm{NO}_{3}^{-}{ }_{(\mathrm{aq})}+2 \mathrm{~K}^{+}{ }_{(\mathrm{aq})}+2 \mathrm{I}^{-} \rightarrow \mathrm{PbI}_{2(\mathrm{~s})}+2 \mathrm{~K}^{+}{ }_{(\mathrm{aq})}+2 \mathrm{NO}_{3}{ }^{-}(\mathrm{aq})$
Net Ionic Eqn:
Shows only the reaction species in a reaction in an aqueous environment (spectator ions are deleted)
$\mathrm{Pb}^{2+}{ }_{(\mathrm{aq})}+2 \mathrm{I}^{-} \rightarrow \mathrm{Pbl}_{2(\mathrm{~s})}$

## Qualitative Analysis

- Procedures by which a chemist decides what substances or ions are present in a sample
- Use the Solubility Table to select a reagent that gives a precipitate (ppt) with some substances but not with others
- If an anion is added to give a ppt, use an alkali metal salt of the anion (eg. $\mathrm{Na}^{+}$or $\mathrm{K}^{+}$) - they will not ppt with other anions
- If adding a cation, use the nitrate salt of the cation
- Generally ppt the "easier to ppt" ion first

Ex. Devise a procedure for separating a mixture of $\mathrm{Ba}^{2+}$ and $\mathrm{Pb}^{2+}$ by precipitating (ppt) them individually.

- $\mathrm{Pb}^{2+}$ ppts with $\mathrm{Cl}^{-}, \mathrm{Br}^{-}, \mathrm{I}^{-}$, and $\mathrm{S}^{2-}$ - but $\mathrm{Ba}^{2+}$ is soluble
- Add a reagent containing $\mathrm{Cl}^{-}, \mathrm{Br}^{-}, \mathrm{I}^{-}$, or $\mathrm{S}^{2-}$

Ex. $\mathrm{NaCl}, \mathrm{KI}, \mathrm{LiBr}$

- $\mathrm{Pb}^{2+}{ }_{(\mathrm{aq})}+2 \mathrm{Br}^{-}{ }_{(\mathrm{aq})} \rightarrow \mathrm{PbBr}_{2(\mathrm{~s})}$
- filter off ppt; only $\mathrm{Ba}^{2+}$ left
- add a reagent containing $\mathrm{SO}_{4}{ }^{2-}, \mathrm{CO}_{3}{ }^{2-}, \mathrm{PO}_{4}{ }^{3-}$, or $\mathrm{SO}_{3}{ }^{2-}$, ions to ppt $\mathrm{Ba}^{2+}$ Ex. $\mathrm{Na}_{3} \mathrm{PO}_{4}$
- $3 \mathrm{Ba}^{2+}{ }_{(\mathrm{aq})}+2 \mathrm{PO}_{4}{ }^{3-}{ }_{(\mathrm{aq})} \rightarrow \mathrm{Ba}_{3}\left(\mathrm{PO}_{4}\right)_{2(\mathrm{~s})}$

Ex. Devise a procedure to separate a mixture of $\mathrm{CO}_{3}{ }^{2-}$ and $\mathrm{Cl}^{-}$

## Titrations

An important analytical technique, titrations are used to determine the concentration of a substance in solution by comparing it to a solution of a known concentration; called a standard

The reaction between the unknown and the standard is allowed to proceed until mole ratios like those in the balanced chemical equation are reached.

This stoichiometric point (sometimes called the equivalence point or end point) is determined using some sort of chemical indicator

Often these reactions are used with acids/bases, for acid/base titrations the end point is reached when:

## Moles of $\mathrm{OH}^{-}$from the base= moles of $\mathrm{H}^{+}$from the acid

Titrations consist of several trials; the first is an estimate, usually done quickly to get an approximate endpoint, almost always the first trial will be over the endpoint (overshot) The average volume of the second and subsequent trials are used in calculations:

## Steps to follow

1) Write the balanced equation (as always - you can't go wrong with this)
2) Calculate the moles of standard used ( $n=M \times V$ )
3) Determine the moles of unknown using the mole ratio
4) Calculate the unknown concentration ( $M=n / V$ )

Ex. If 10.0 mL of HCl is titrated with 20.0 mL of 0.40 M NaOH , what is [ HCl ?

$$
\mathrm{NaOH}_{(\mathrm{aq})}+\mathrm{HCl}_{(\mathrm{aq)}} \rightarrow \mathrm{H}_{2} \mathrm{O}_{(\mathrm{l})}+\mathrm{NaCl}_{(\mathrm{aq})}
$$

Moles of standard (in this case NaOH ) $=$
Moles of unknown $(\mathrm{HCl})=$
$[\mathrm{HCl}]=$

- Some acids or bases produce two (or more) $\mathrm{H}^{+}$or $\mathrm{OH}^{-}$for each formula unit $\left(\mathrm{H}_{3} \mathrm{PO}_{4} \rightarrow \mathbf{3} \mathbf{H}^{+}+\mathrm{PO}_{4}{ }^{3-}\right)$

Ex. An average volume of 24.90 mL of 0.100 M NaOH was required to neutralize 15.00 mL of $\mathrm{H}_{2} \mathrm{SO}_{4}$. What is the concentration of the acid?

Moles of standard (in this case NaOH ) $=0.100 \mathrm{M} \times 0.02490 \mathrm{~L}=2.49 \times 10^{-3}$ Moles of unknown $(\mathrm{HCl})=$
$[\mathrm{HCl}]=$

