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Solution Chemistry

Dealing with mixtures

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Solutions

A **solution** is a homogenous mixture consisting of a **solvent** and at least one **solute**.

The solvent is the most prevalent species.

The solute is the less prevalent species.

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Examples of Solutions

Saline (salt water) is a solution. The solvent is water, the solute is salt.

Wet salt is also a solution. The solvent is salt, the solute is water.

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160 proof Vodka

What is the solvent?

Alcohol – It is 80% alcohol.

What is the solute?

Water – It is 20% water.

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Aqueous Solutions

Aqueous solutions are specifically solutions where water is the solvent.

Aqueous solutions are a very common medium for performing chemical reactions.

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Advantages of Aqueous Solutions

1. Mixing – you can stir the solution.
2. Ability to dissipate heat (or cold) – the mass of the solvent allows it to absorb significant amounts of heat (or cold).
3. “Universal solvent” – water dissolves many different materials, especially ionic materials.

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Concentration

Because a solution is a mixture – there are different ratios of solvent/solute quantities possible.

For example, I could put 1 teaspoon of salt in a cup of water OR I could put 2 teaspoons of salt in a cup of water.

Both are saline solutions, but they have different amounts of salt.

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Concentration

Almost any unit of measure can be used to specify concentration. (teaspoon solute/cup solvent would work!)

There are certain common units of measuring solution concentration that are most frequently used. Understanding their UNITS! UNITS! UNITS! And being able to manipulate those UNITS! UNITS! UNITS! is crucial.

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Common Concentration Units

- % by mass - g solute/100 g solution
- Molarity (M) - moles solute/L solution
- Molality (m) - moles solute/kg solvent
- % by volume - mL solute/100 mL solution
- mass-volume % - g solute/100 mL solution
- Mole Fraction – moles solute/total moles of all species
- Mole % - moles solute/100 moles of all species

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Converting units

What is the molarity of a 10% by mass aqueous NaCl solution?

UNITS! UNITS! UNITS!

$$\frac{10\text{ g NaCl}}{100\text{ g NaCl solution}} = \frac{\text{moles NaCl}}{\text{L solution}}$$

To convert g NaCl to moles, you need to know…
Molar mass of NaCl

To convert g solution to L solution, you need to know…
Density of the solution

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The Density

We ALWAYS know the molar mass of any substance.

But what about the density?

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The Density

You don't always know the density.

Density depends on concentration.

Sometimes you know the density.
Sometimes you can figure out the density.
Sometimes you just have to ASSUME the density.

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The Density

If you don't know anything except what was given:

What is the molarity of a 10% by mass aqueous NaCl solution?

What would you do?

Assume the density is that of pure water (1.00 g/mL at 25°C)

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The Density

Suppose I had further information:

What is the molarity of a 10% by mass aqueous NaCl solution? (Density of 5% NaCl solution = 1.05 g/mL, Density of 20% NaCl solution = 1.13 g/mL)

Now what would you do?

I can either ASSUME that 5% is "close enough" to 10%.
OR I can "interpolate" the density between 5% and 20%.

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Linear Interpolation

Do you know what a "linear interpolation is"?

I assume that there is a linear (straight-line) dependence of the density on the concentration. (This is not, by the way, true but it is an OK assumption if the range is narrow enough.)

Then I draw a straight line between the two points I know and find the interpolated concentration at my concentration of interest.

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My Problem

What is the molarity of a 10% by mass aqueous NaCl solution? (Density of 5% NaCl solution = 1.05 g/mL, Density of 20% NaCl solution = 1.13 g/mL)

I find the slope of the line: $\frac{\Delta \text{ Density}}{\Delta \text{ \% NaCl}}$

$\frac{1.13 \text{ g/mL} - 1.05 \text{ g/mL}}{20\% - 5\%} = 5.33\times10^{-3} \frac{\text{g/mL}}{\%}$

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My Problem

What is the molarity of a 10% by mass aqueous NaCl solution? (Density of 5% NaCl solution = 1.05 g/mL, Density of 20% NaCl solution = 1.13 g/mL)

$5.33\times10^{-3} \frac{\text{g/mL}}{\%}$ means that every 1% change in concentration results in a 5.33×10^{-3} g/mL change in density

10%-5% = 5% change

$5.33\times10^{-3} \frac{\text{g/mL}}{\%} * 5\ \% = 0.0267$ g/mL change

1.05 g/mL + 0.0267 g/mL = 1.077 g/mL = 1.08 g/mL interpolated density

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Solving the problem

What is the molarity of a 10% by mass aqueous NaCl solution? (Density of 5% NaCl solution = 1.05 g/mL, Density of 20% NaCl solution = 1.13 g/mL)

$\frac{10 \text{ g NaCl}}{100 \text{ g solution}} * \frac{1 \text{ mol NaCl}}{58.45 \text{ g NaCl}} = \frac{0.171 \text{ mol NaCl}}{100 \text{ g solution}}$

$\frac{0.171 \text{ mol NaCl}}{100 \text{ g solution}} * \frac{1.08 \text{ g solution}}{1 \text{ mL solution}} * \frac{1000 \text{ mL}}{1 \text{ L}} = \frac{1.84 \text{ mol NaCl}}{\text{L solution}}$

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Converting units

Typically speaking, you can convert any of the concentration units into any of the others as long as you have the Molar Mass and the Density!

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Some Other Examples

I have an aqueous solution of carbon dioxide that is 1.1 % by mass. What is the Molarity of this solution?

$$\frac{1.1\text{ g }CO_2}{100\text{ g solut}} * \frac{1\text{ mol }CO_2}{44.01\text{ g }CO_2} * \frac{1.0\text{ g sol}}{1.0\text{ mL sol}} * \frac{1000\text{ mL}}{1\text{ L}} = 0.25\text{ M }CO_2$$

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Further Example

56.0 g of Fe_2O_3 was dissolved in water yielding a total solution volume of 2.65 L. What is the molarity of the resulting solution?

$$56.0\text{ g }Fe_2O_3 * \frac{1\text{ mol }Fe_2O_3}{159.69\text{ g }Fe_2O_3} = 0.351\text{ mol }Fe_2O_3$$

$$\frac{0.351\text{ mol }Fe_2O_3}{2.65\text{ L}} = 0.132\text{ M }Fe_2O_3$$

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What's it all about?

MOLES! MOLES! MOLES!

Specifically, doing reactions!

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An example

56.50 mL of a 2.15 M ammonium sulfate solution is mixed with 36.0 g of iron (III) chloride. If the reaction proceeds with a 65% yield, how much iron (III) sulfate would be acquired?

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Limiting Reagent Problem

What's the first thing you need?

A balanced equation!

$(NH_4)_2SO_4 + FeCl_3 \rightarrow Fe_2(SO_4)_3 + NH_4Cl$

How do you know this is the right products?

Charges! This is an example of a double replacement reaction. The cations get switched (or the anions, if you prefer).

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Limiting Reagent Problem

We still need to balance it!

$(NH_4)_2SO_4 + FeCl_3 \rightarrow Fe_2(SO_4)_3 + NH_4Cl$

$3\ (NH_4)_2SO_4 + 2\ FeCl_3 \rightarrow Fe_2(SO_4)_3 + 6\ NH_4Cl$

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Armed with Stoichiometry!

56.50 mL of a 2.15 M ammonium sulfate solution is mixed with 36.0 g of iron (III) chloride. If the reaction proceeds with a 65% yield, how much iron (III) sulfate would be acquired?

$3\ (NH_4)_2SO_4 + 2\ FeCl_3 \rightarrow Fe_2(SO_4)_3 + 6\ NH_4Cl$

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Armed with Stoichiometry!

56.50 mL of a 2.15 M ammonium sulfate solution is mixed with 36.0 g of iron (III) chloride. If the reaction proceeds with a 65% yield, how much iron (III) sulfate would be acquired?

$3\ (NH_4)_2SO_4 + 2\ FeCl_3 \rightarrow Fe_2(SO_4)_3 + 6\ NH_4Cl$

$$36.0\text{ g } FeCl_3 * \frac{1\text{ mol } FeCl_3}{162.21\text{ g } FeCl_3} * \frac{1\text{ mol } Fe_2(SO_4)_3}{2\text{ mol } FeCl_3} * \frac{399.87\text{ g } Fe_2(SO_4)_3}{1\text{ mol } Fe_2(SO_4)_3} = 44.37\text{ g } Fe_2(SO_4)_3$$

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Armed with Stoichiometry!

56.50 mL of a 2.15 M ammonium sulfate solution is mixed with 36.0 g of iron (III) chloride. If the reaction proceeds with a 65% yield, how much iron (III) sulfate would be acquired?

$3\ (NH_4)_2SO_4 + 2\ FeCl_3 \rightarrow Fe_2(SO_4)_3 + 6\ NH_4Cl$

2.15 M $(NH_4)_2SO_4$ =

$$\frac{2.15\ \text{mol}\ (NH_4)_2SO_4}{1\ \text{L solution}} * \frac{1\ \text{L solut}}{1000\ \text{mL solut}} * 56.50\ \text{mL solut} = 0.121\ \text{mol}\ (NH_4)_2SO_4$$

$$0.121\ \text{mol}\ (NH_4)_2SO_4 * \frac{1\ \text{mol}\ Fe_2(SO_4)_3}{3\ \text{mol}\ (NH_4)_2SO_4} * \frac{399.87\ \text{g}\ Fe_2(SO_4)_3}{1\ \text{mol}\ Fe_2(SO_4)_3} = 16.13\ \text{g}\ Fe_2(SO_4)_3$$

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Armed with Stoichiometry!

56.50 mL of a 2.15 M ammonium sulfate solution is mixed with 36.0 g of iron (III) chloride. If the reaction proceeds with a 65% yield, how much iron (III) sulfate would be acquired?

$3\ (NH_4)_2SO_4 + 2\ FeCl_3 \rightarrow Fe_2(SO_4)_3 + 6\ NH_4Cl$

Limiting Reagent is $(NH_4)_2SO_4$: 16.13 g $Fe_2(SO_4)_3$ theoretical

$$16.13\ \text{g}\ Fe_2(SO_4)_3\ \text{theoretical} * \frac{65\ \text{g actual}}{100\ \text{g theoretical}} = 10.48\ \text{g actual}\ Fe_2(SO_4)_3$$

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Clicker Question

I have 1 L of a solution that is 5.4% by mass sodium sulfate. If the density of 5% sodium sulfate is 1.085 g/mL, how much silver (I) chloride would I need to add to precipitate all of the sulfate?

A. 59 g
B. 257 g
C.118 g
D. 129 g
E. 25.4 g

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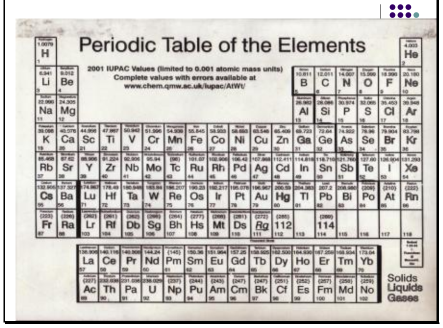

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$Na_2SO_4 + 2\ AgCl \rightarrow 2\ NaCl + Ag_2SO_4$

$1L * \frac{1000\ mL}{1L} * \frac{1.085\ g}{1\ mL} * \frac{5.4\ g\ Na_2SO_4}{100\ g\ solution} = 58.59\ g\ Na_2SO_4$

$58.59\ g\ Na_2SO_4 * \frac{1\ mol\ Na_2SO_4}{142\ g\ Na_2SO_4} = 0.4126\ mol\ Na_2SO_4$

$0.4126\ mol\ Na_2SO_4 * \frac{2\ mol\ AgCl}{1\ mol\ Na_2SO_4} = 0.825\ mol\ AgCl$

$0.825\ mol\ AgCl * \frac{143\ g\ AgCl}{1\ mol\ AgCl} = 118\ g\ AgCl$

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Question

When 50.00 mL of 0.125 M silver (I) nitrate is mixed with 50.00 mL of 0.250 M sodium sulfate a greyish solid forms. If I recover 0.813 g of solid, what is the yield of the reaction?