3: Concentration, Solutions, pH

I. Components of a Solution

II. Concentration: Units and Calculations

III. Preparing Solutions

   A. Measuring and Transferring Solute

   B. Using Stock Solutions

IV. Measuring pH

   A. pH meters

V. Desiccator Use

Solutions

<u>Solution</u>: Homogeneous mixture

<u>Solute</u>: Minor component in solution

<u>Solvent</u>: Major component in solution

Example: If a teaspoon of sugar is dissolved in a cup of water, sugar is the *solute*, water is the *solvent*, and the entire mixture is the *solution*.

Concentration

Concentration: The amount of *solute* in the *solution*

"Amount" may be measured in different ways

Solute	Solution
Number (moles)	
Mass	Mass
Volume	Volume

Concentration Units

Solute: Number / Solution: Volume	$\frac{\text{Moles}}{\text{Liter}}$		= Molarity (M)
Solute: Mass / Solution: Mass	$\frac{\text{g}}{\text{g}}$	×100	= % by weight
Solute: Volume / Solution: Volume	$\frac{\text{L}}{\text{L}}$	×100	= % by volume
Solute: Mass / Solution: Volume	$\frac{\text{grams}}{\text{Liter}}$		= g/L

Note: Density is also mass per volume (g/mL), but of the whole thing

Concentration Units: "Parts Per…"

Parts per million means: $\frac{\text{One part solute}}{\text{One million parts solution}}$

Parts per billion means: $\frac{\text{One part solute}}{\text{One billion parts solution}}$

Solute / Solution	Mass / Mass	mg / Liter	= ppm
Solute: / Solution:	Mass / Mass	μg / Liter	= ppb

Note: 1 L of water is 1 kg, which is one million mg

Molality: A Weird One

Solute: / Solvent:	Number / Mass	Moles / kg	Molality (*m*)

Denominator is "Solvent"!

- Easy to calculate
- Difficult to use
- Will not see in this class

Examples

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$$\text{Molarity (M)} = \frac{\text{Moles}_{\text{solute}}}{\text{Liter}_{\text{solution}}}$$

$$\text{moles} = 0.500\text{ g} \times \frac{1\text{ mole}}{58.44\text{ g}} = 0.00855\text{ mol}$$

$$\text{Liter} = 25.0\text{ mL} \times \frac{1\text{ L}}{1000\text{ mL}} = 0.0250\text{ L}$$

$$\text{Molarity (M)} = \frac{\text{Moles}}{\text{Liter}} = \frac{0.00855\text{ mol}}{0.0250\text{ L}} = 0.342\text{ M}$$

Examples

II. 0.010 g $Pb(NO_3)_2$ (MW 331.2 g) is dissolved in 2.0 L water. What is the ppm of Pb^{2+} (AW 207.2g)?

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$$\text{g}_{\text{solute}} = 0.010\text{ g} \times \frac{207.2\text{g}_{\text{Pb}}}{331.2\text{ g}_{\text{Pb(NO3)2}}} = 0.0062\text{ g}_{\text{Pb}}$$

$$\text{mg}_{\text{solute}} = 0.0062\text{ g} \times \frac{1000\text{ mg}}{1\text{g}} = 6.2\text{ mg}_{\text{Pb}}$$

$$\text{ppm} = \frac{\text{mg}_{\text{solute}}}{\text{Liter}_{\text{solution}}} = \frac{6.2\text{ mg}_{\text{Pb}}}{2.0\text{ L}} = 3.1\text{ ppm}$$

Examples

III. A 355 mL (12 oz) can of Coca-Cola contains 42.1 g of fructose and has a density of 1.15 g/mL. What is the weight percent of fructose?

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$$\%wt = \frac{g_{solute}}{g_{solution}}$$

$$g_{solution} = 355\text{ mL} \times \frac{1.15\text{g}}{1\text{ mL}} = 408\text{ g}$$

$$\%wt = \frac{g_{solute}}{g_{solution}} = \frac{42.1\text{ g}}{408\text{ g}} = 10.3\%$$

Preparing Solutions

<u>Steps</u>

1. Measure out Solute
2. Transfer to flask designed "To Contain"
3. Dilute with Solvent to final volume

The Solute

I. Usually most accurate to <u>weigh solute</u>, but...
   A. Liquids may evaporate while being weighed
   B. Some solutes are deliquescent

II. Volumetric transfer pipettes good for liquids if little less accuracy is OK

Transferring Solute

I. "Quantitatively Transfer" the solute
   A. Every solute molecule ends up in the solution flask
   B. Rinse weighing vessel with solvent 3 times
   C. May pre-dissolve in weighing vessel with solvent
      1. Helpful for static electricity and evaporating liquids

II. Heat needed to dissolve solute?
   A. Gently heating Pyrex vol. flasks OK
   B. Better to pre-dissolve by <u>heating in a beaker</u>

Adding Solvent

I. Add small amount of solvent before solute
   A. Minimizes evaporation of liquid solute
   B. Aids dissolution of solute

II. After adding solute, fill the flask half way, stopper and swirl to dissolve

III. When completely dissolved, fill flask to the mark
   A. USE A DROPPER FOR LAST mL
   B. <u>NEVER</u> GO OVER THE MARK!!

IV. Stopper flask; gently invert several times

Questions when Preparing Solutions

I. What will the solution be used for?

II. When will the solution be used?

III. What accuracy is needed for concentration?

IV. To what extent is purity / contamination an issue?

V. What contaminates are most critical?

VI. What volume is needed?

VII. Will the solution change over time and how?

Diluting Stock Solutions

I. Accurate and convenient way to make lower concentration solutions

II. Concentrated stock solutions have less tendency to change over time

III. Accuracy 1% or better with volumetric pipettes

IV. Accuracy 2-3 % with graduated cylinders

http://www.chemistry.umu.se/digitalAssets/4/4612_science_chemistry.gif

Calculating Stock Solution Dilution

Moles removed from Stock = Moles in Diluted Solution

Volume × Concentration = Moles

(L) (mol/L) (mol)

$$V_s \times C_s = V_d \times C_d \qquad (s = \text{stock},\ d = \text{diluted})$$

$$V_s = V_d \times (C_d / C_s)$$

Examples

I. What volume of 1.25 M Stock solution is needed to make 25.0 mL of 0.0250M solution?

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$$V_s = V_d \times \frac{C_d}{C_s}$$

$$V_s = 25.0 \text{ mL} \times \frac{0.0250 \text{ M}}{1.25 \text{ M}} = 0.500 \text{ mL}$$

Examples

II. 1.00 mL of 5.0 g/L solution is diluted to 50.0 mL. What is the concentration in ppm?

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$$\text{ppm} = \frac{\text{mg}_{\text{solute}}}{\text{Liter}_{\text{solution}}}$$

$$\text{mg}_{\text{solute}} = 1.00\ \text{mL} \times \frac{5.0\ \text{g}_{\text{solute}}}{1\ \text{L}_{\text{solution}}} = 5.0\ \text{mg}$$

$$\text{L}_{\text{solution}} = 50.0\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}} = 0.0500\ \text{L}$$

$$\text{ppm} = \frac{\text{mg}_{\text{solute}}}{\text{Liter}_{\text{solution}}} = \frac{5.0\ \text{mg}}{0.0500\ \text{L}} = 100\ \text{ppm}$$

pH

$$\text{pH} = -\log\,[H^+]$$

Check: Take out your calculator and calculate:

I. The pH when $[H^+] = 1.25 \times 10^{-4}$ M

II. $[H^+]$ when pH = 8.7

$$[H^+] = 10^{-\text{pH}}$$

pH

$$pH = -\log [H^+]$$

Check: Take out your calculator and calculate:

I. The pH when $[H^+] = 1.25 \times 10^{-4}$ M

II. $[H^+]$ when pH = 8.7

I. 3.9

II. 2.00×10^{-9} M

$$[H^+] = 10^{-pH}$$

pH Electrode

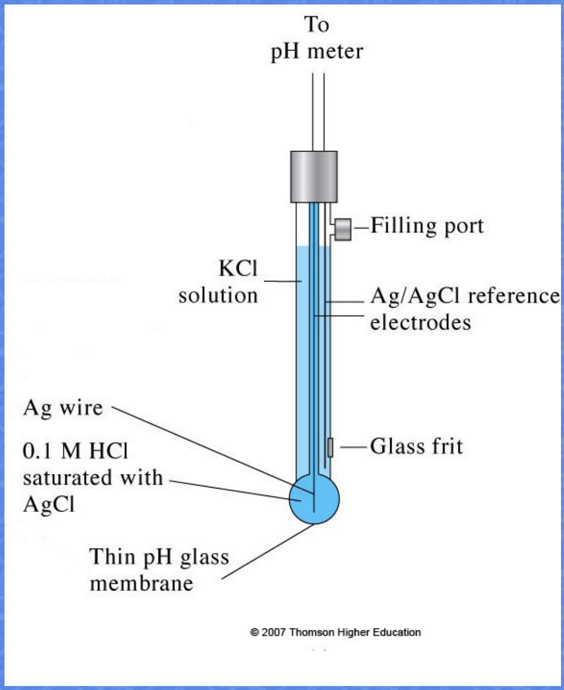
I. Two electrodes in one barrel

II. Outside one responds to $[H^+]$

III. Other is reference electrode: voltage does not change

IV. Both have own electrolyte solutions (must keep filled)

V. Sometimes electrodes are sealed and do not need filling



pH Electrodes

I. Tip has thin glass membrane (~0.1 mm)
   A. Easily broken!
II. Membrane must be hydrated to function
   A. Immerse electrode in water or buffer solution at least 30 minutes before use
III. Water and tiny H^+ ions move through membrane
   A. Charge difference creates a voltage
   B. pH meter converts voltage to pH
   C. Best to store pH probe in 7.0 buffer
   D. Requires calibration!

pH Meter Calibration

I. pH meter only as good as buffers used to calibrate it
II. Usually use commercially prepared buffers
III. Buffers accurate to ±0.01 pH units
IV. pH 4, 7, 10 used to calibrate meters
V. Choose pair based on expected measurements
VI. Re-calibrate often

Desiccators

I. Dry Storage of:
   A. Hygroscopic reagents
   B. Cooling glassware
II. Lid-base seal is greased
   A. Slide lid off—NEVER pull up!
   B. Do NOT set lid down (except back on base)

Today's lab: Preparing Solutions and Measuring pH

I. Acid Solution using Solid Solute
II. Base Solution by Diluting Solution of Known Concentration

<u>Note</u>: Next week, **NO pre-lab**

http://www.bbc.co.uk/bitesize/ks3/science/images/universal_indicator.jpg