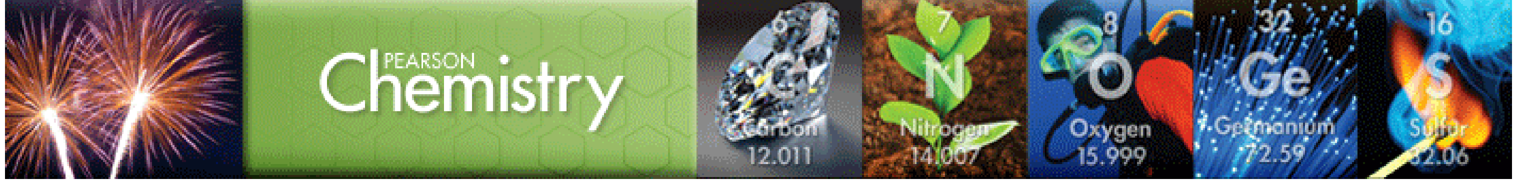

كلية المعارف الجامعة
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Lecture 3

First Year
Medical Laboratory Techniques Department
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2023-2024

Dilution of Solutions

In a dilution, a solvent, usually water, is added to a solution, which increases its volume and decreases the concentration of the solution while The mass of solute in the solution remains the same. For example. making orange juice from concentrate is an example of a dilution.

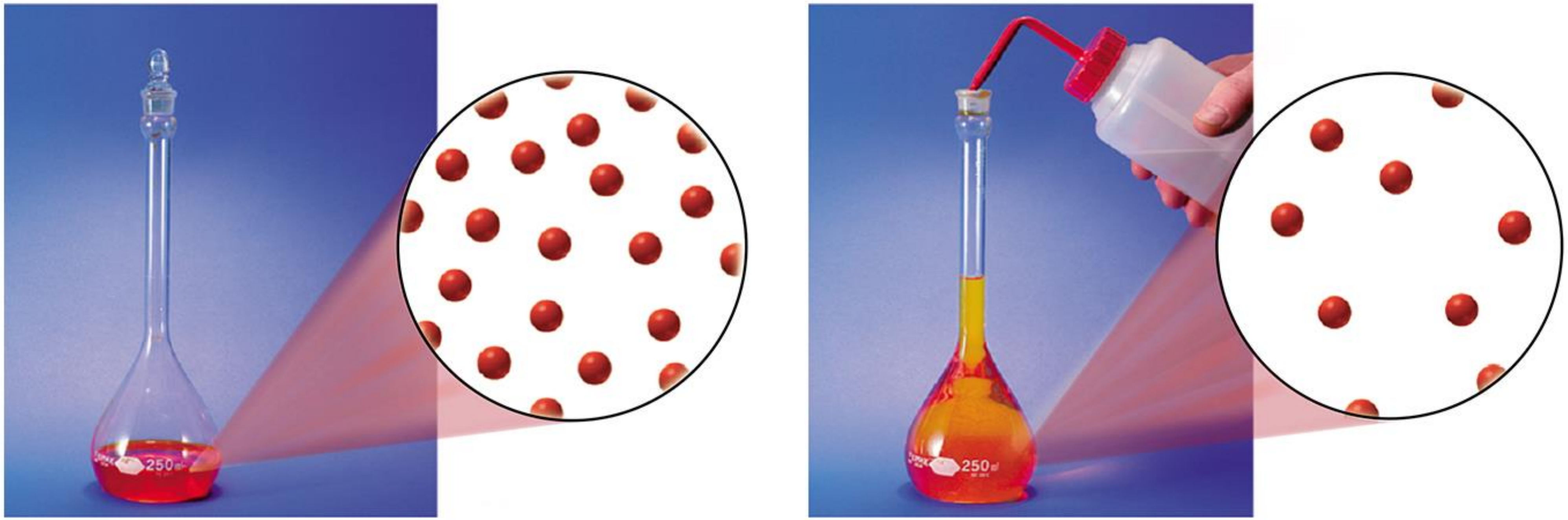
(a) A volume (V_s) containing the desired moles of solute (M_s) is measured from a stock solution of known concentration.

(b) The measured volume of stock solution is transferred to a second volumetric flask.

(c) The measured volume in the second flask is then diluted with solvent up to the volumetric mark $[(V_s)(M_s) = (V_d)(M_d)]$.

Dilution of a Solution

When water is added to a concentrated solution, there is no change in the number of particles. The solute particles spread out as the volume of the solution increases.



Dilutions and Molarity

- Use this formula to make a more dilute solution from a concentrated solution

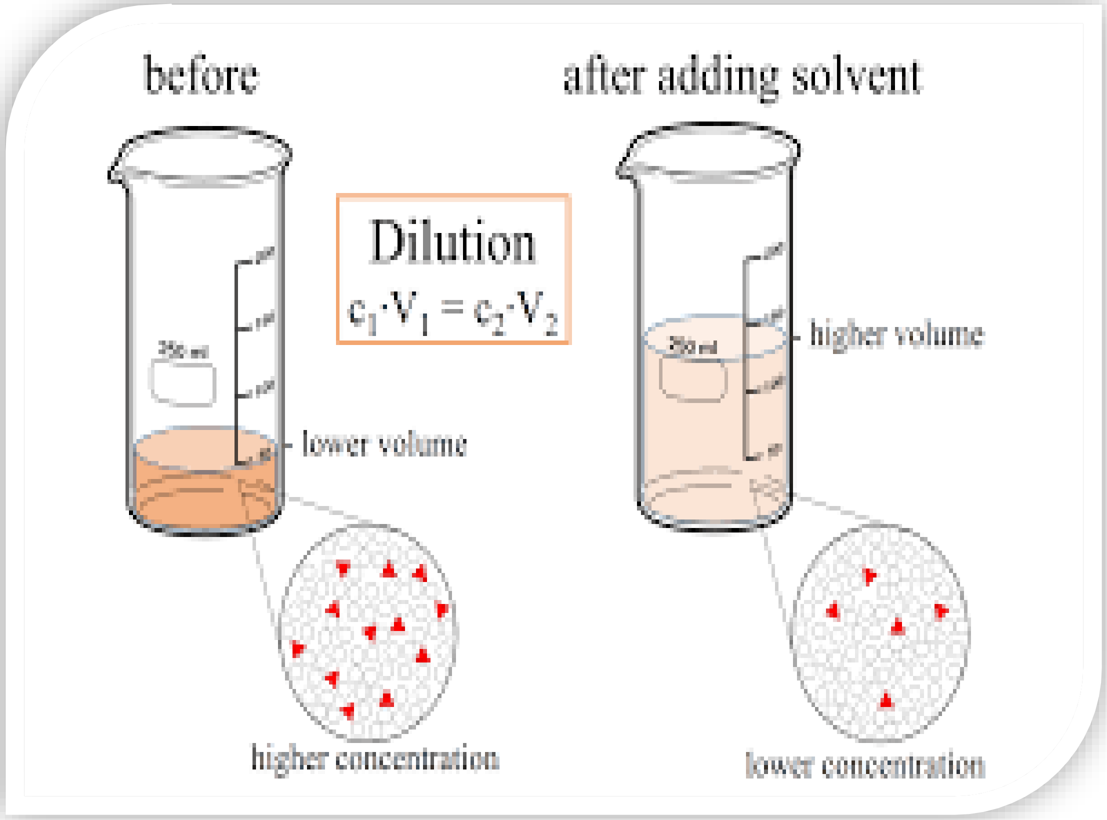
$\text{Molarity}_1 \text{ x Volume}_1 = \text{Molarity}_2 \text{ x Volume}_2$

(Concentrated) (Dilute)

(Before) = (After)

$M_1V_1 = M_2V_2$ or $C_1V_1 = C_2V_2$

initial diluted



Example 1

How many liters of 2.5 M HCl are required to make 1.5 L of 1.0 M HCl?

The solution:

Applying the law of the dilution:

$$M_1 \times V_1 = M_2 \times V_2$$

$M_1 = 2.5$ M

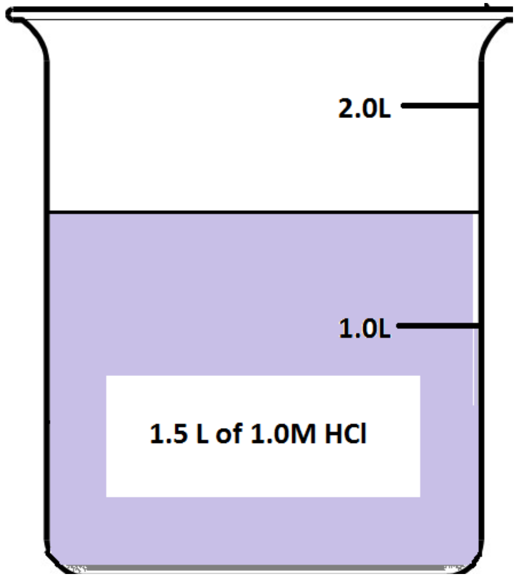
$V_1 = ?$

$M_2 = 1.0$ M

$V_2 = 1.5$ L

$$2.5\ \cancel{\mathrm{mol}}/\cancel{\mathrm{L}} \times V1 = 1.0\ \cancel{\mathrm{mol}}/\cancel{\mathrm{L}} \times 1.5\mathrm{L}$$

$$V1 = \frac{1.5\ L \times 1.0}{2.5} = 0.6\ \mathrm{L}$$



Example 2

250.0 mL of a 0.500 M HCl solution needs to be made from concentrated HCl. What volume of the concentrated solution is needed if its molarity is 12.0 M?

M_1 = 12.0M

V_1 = ?

M_2 = 0.500M

V_2 = 250.0mL

How much water would you add to make the final solution?

$$M_1V_1 = M_2V_2$$

12 mol/L x V= 0.5 mol/L x 250 ml

V1= 10.4 ml

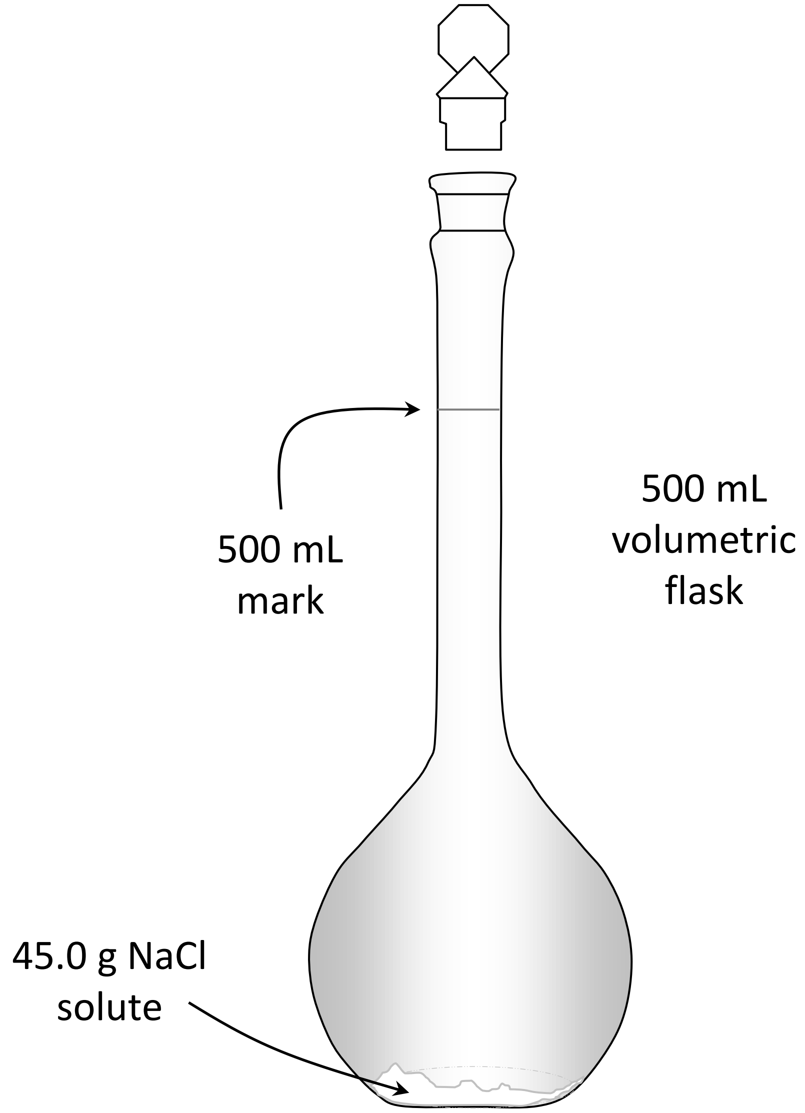
Example 3

How to prepare 500 mL of 1.54 *M* NaCl solution?

$$M = \frac{W}{MWt} x \frac{1000}{V_{ml}}$$

$$W = \frac{M \times MWt \times V_m}{1000} = \frac{1.54 \times 58.5t \times 500}{1000}$$

W of NaCl = 45.0 g.



Example 4

Initial Solution →

What volume of water must be added to 150.0 mL of 0.200 M NaCl solution to change its concentration to 0.0250 M ?

$$M_1V_1 = M_2V_2$$

$$(0.200\ \text{M})(150.0\ \text{mL}) = (0.0250\ \text{M})(V_2)$$

$$V_2 = 1200\ \text{mL}$$

$$\text{Volume of } H_2O \text{ added} = 1200\ \text{mL}$$

$$- 150.0\ \text{mL}$$

$$= 1050\ \text{mL}$$

Example 5

What volume of **0.200 M $CuSO_4$** solution must be diluted to **500.0 mL** to produce a **0.150 M** solution?

$$M_1V_1 = M_2V_2$$

$$(0.200\ \text{M})(V_1) = (0.150\ \text{M})(500.0\ \text{mL})$$

$$V_1 = 375\ \text{mL}$$

Example 6

1.53 g of NaCl is dissolved in **100.0 mL** of water. Calculate the molarity. The solution is concentrated by evaporating off **60.0 mL** of water. Calculate the **new molarity**. To your knowledge that the molar mass of NaCl= 58.5 g/mol.

Solution:

$$M = \frac{m}{M} \text{x} \frac{1}{V\ l}$$

$$M = \frac{1.53\ g}{58.5\ g/mol} \text{x} \frac{1}{0.1\ L} = 0.262\ \text{M}$$

100.0 mL – 60.0 mL

$$\mathbf{M_1V_1 = M_2V_2}$$

$$\mathbf{(0.26153\ M)(100.0\ mL) = M_2(40.0\ mL)}$$

The molarity of NaCl → $\mathbf{M_2 = 0.654\ M}$

Example 7

How much water should be mixed with 5000ml of 85% alcohol to make 50% (v/v) solution ?

Solution:

- $C_1V_1 = C_2V_2$

5000ml x 85% = 50% x V_2

V_2 = 8500ml

8500 – 5000 = 3500ml of H_2O.

- .

Note

Standard solution is a solution of known concentration (normality, molarity and molality) or its concentration is exactly measured.

Standardization is determination of the molarity or normality of the solution

Example 8

If 500ml of 15% v/v solution are diluted to 1500ml. What will be the percentage strength ?

Answer:

C1V1 = C2V2

15% x 500ml =C2 x 1500ml

C2=5%

Example 9

How many mls of a 1:5000 (w/v) solution of potassium permanganate can be made from 50ml of a 5% solution ?

1:5000 = 0.02%

C1V1=C2V2

50ml x 0.5% = 0.02% x V2

V2= 1250ml

The History of Acids and Bases

In the early days of chemistry chemists were organizing physical and chemical properties of substances. They discovered that many substances could be placed in two different property categories:

Substance A

1. Sour taste
2. Reacts with carbonates to make CO_2
3. Reacts with metals to produce H_2
4. Turns blue litmus pink
5. Reacts with B substances to make salt and water

Substance B

1. Bitter taste
2. Reacts with fats to make soaps
3. Do not react with metals
4. Turns red litmus blue
5. Reacts with A substances make salt and water

Arrhenius Theory

The Swedish chemist Svante Arrhenius proposed the first definition of acids and bases. (Substances A and B became known as acids and bases)

Acids **are substances that dissociate in water to** produce H^+ ions **and** Bases **are substances that dissociate in water to** produce OH^- ions”

$NaOH\ (aq) \longrightarrow Na^+\ (aq) + OH^-\ (aq)$ **Base**

$HCl\ (aq) \longrightarrow H^+\ (aq) + Cl^-\ (aq)$ **Acid**

What if the acids and based are not dissolved in water

- The Arrhenius definition for acids and bases only refers to compounds **dissolved in water**.
- Does this mean that acids and bases cannot exist out of water ?
- Not quite, that's where the Bronsted-Lowry definition comes in.

Bronsted

Lowry

Bronsted-Lowry Theory

Johannes Brønsted and Thomas Lowry revised Arrhenius's acid-base theory to include other solvents besides water. They defined acids and bases as follows:

An acid is a hydrogen containing species that donates a proton.

A base is any substance that accepts a proton"

$$HCl\ (aq)\ +\ H_2O\ (l) \longrightarrow Cl^-\ (aq)\ +\ H_3O^+\ (aq)$$

Quiz Question: In the above example,
what is the Brønsted acid?
what is the Brønsted base?

In reality, the reaction of HCl with H_2O is an equilibrium and occurs in both directions, although in this case the equilibrium lies far to the right.

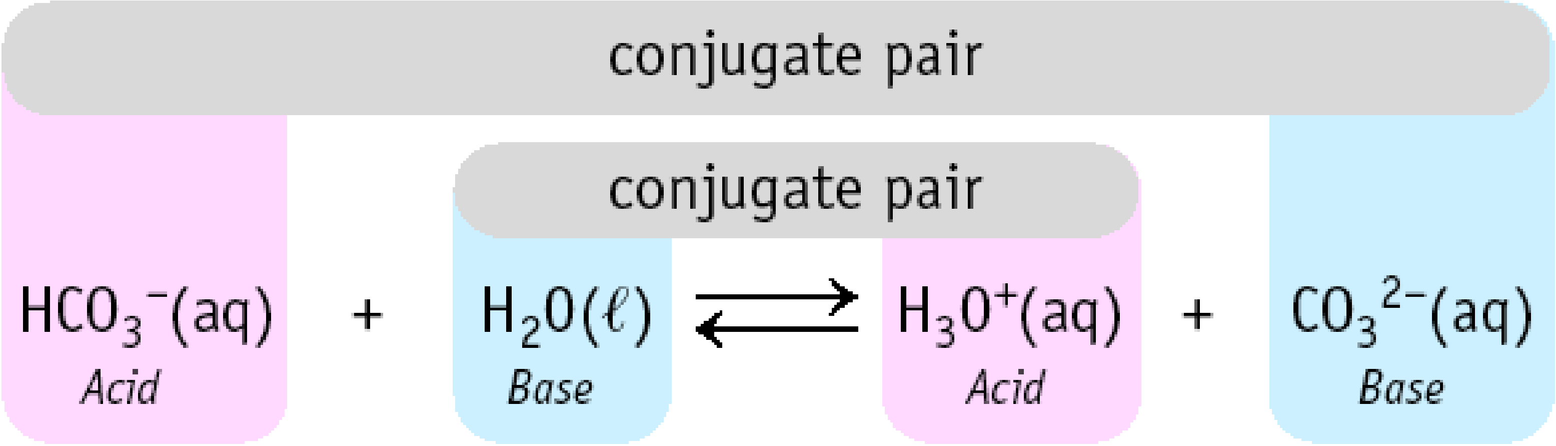
$$HCl\ (aq) + H_2O\ (l) \longrightarrow Cl^-\ (aq) + H_3O^+\ (aq)$$

- For the **reverse reaction** Cl^- behaves as a **Brønsted base** and H_3O^+ behaves as a **Brønsted acid**. The Cl^- is called the **conjugate base** of HCl.

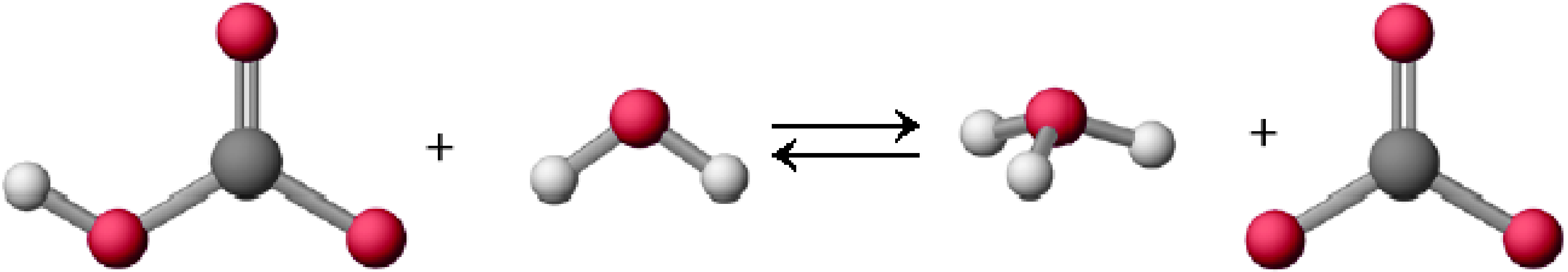
- **Brønsted** acids and bases always exist as **conjugate acid-base pairs**. Their formulas differ by only one proton.

- A conjugate acid is any hydrogen containing material (molecule or ion) that can release a proton or hydrogen ion to any other substances.
- A conjugate base is any substance (molecules or ions) that can accept a proton to any other substances to form the conjugate acid base pair.

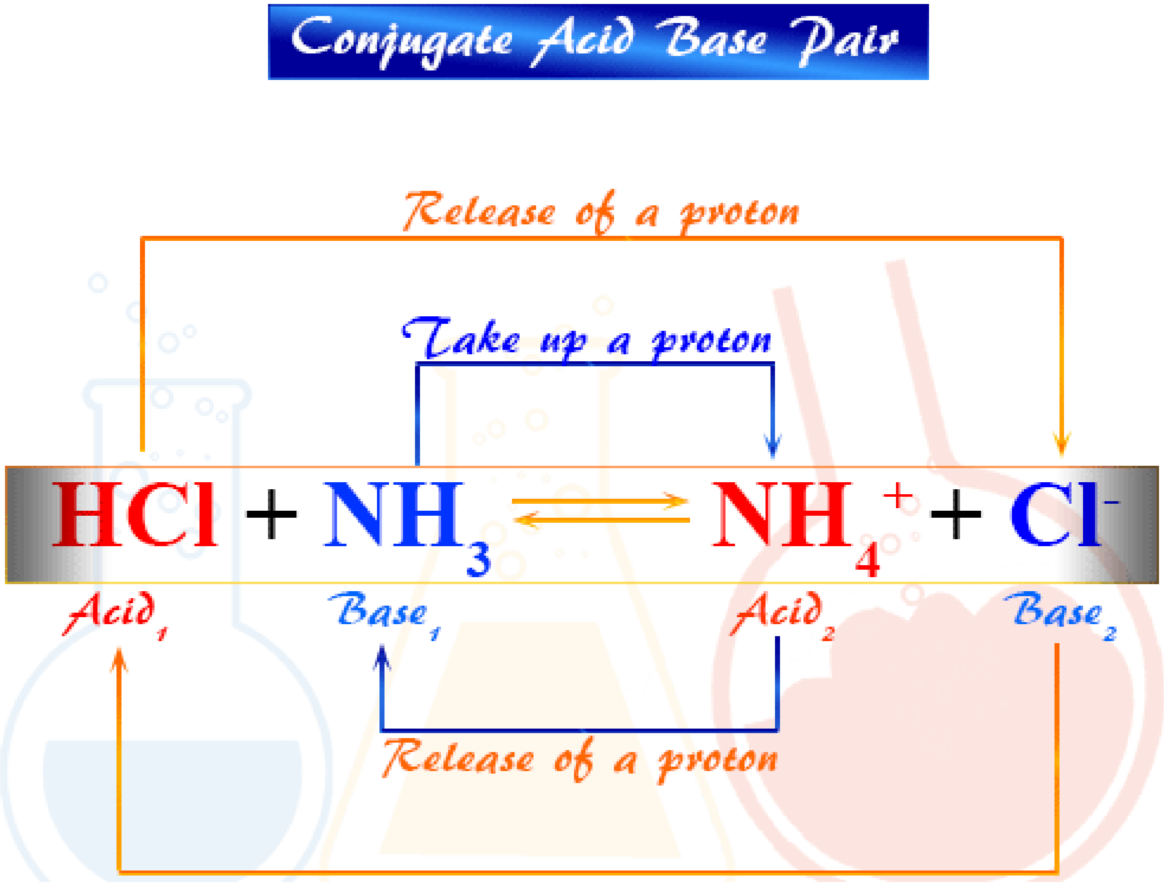
Conjugate Pairs



$$HCO_3^-(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq)$$



Conjugate acid-base theory can not explain the conjugate acid base phenomena in terms of electronic structure by the formation of the coordinate covalent bond between the vacant orbital and the orbital which contains lone pair of electrons. This can be explained by the Lewis acid base theory.



Quiz: Label the acid, base, conjugate acid, and conjugate base in each reaction:

$HCl + OH^- \rightarrow Cl^- + H_2O$

$H_2O + H_2SO_4 \rightarrow HSO_4^- + H_3O^+$

A **single arrow** is used to represent the ionization of strong acids.

$$HCl_{(g)} \longrightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$$

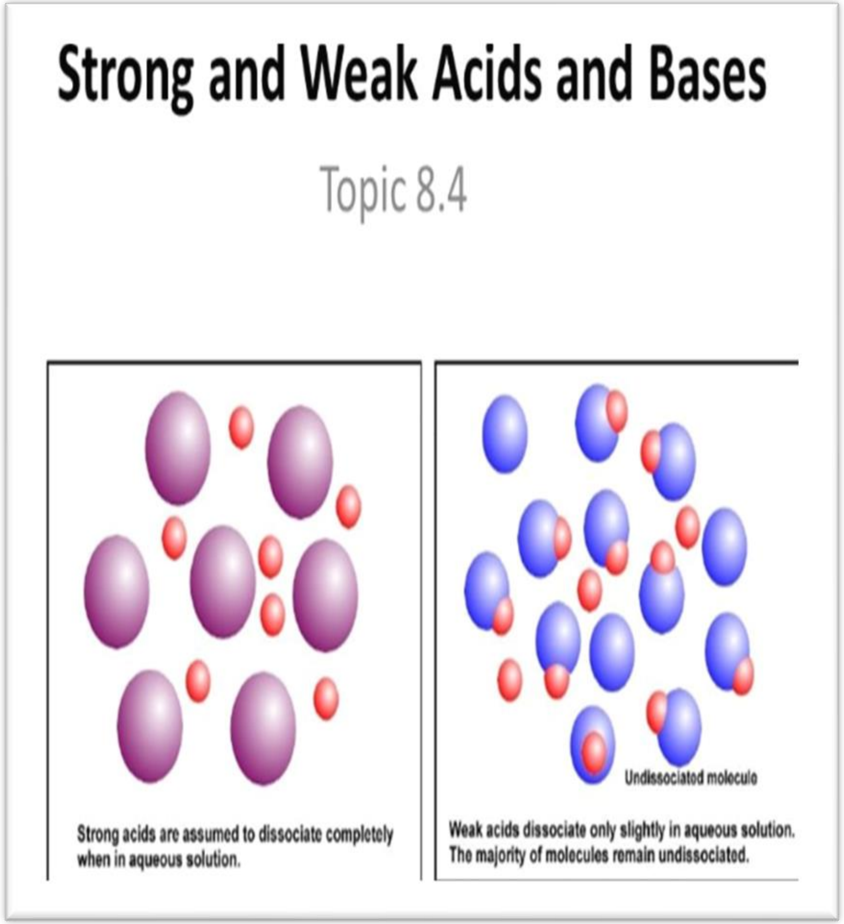
double arrows are used to represent ionization of weak acids because an equilibrium is created.

$$HF_{(g)} \rightleftarrows H^{+}_{(aq)} + F^{-}_{(aq)}$$

Note: Strong acids ionize 100% and weak ones do not!

Common Strong Acids

- These are the strong acids. What makes them 'strong' is that they completely dissociate into their ions when they are mixed with water.
- As the strong acids become more concentrated, they may be unable to fully dissociate.
- The rule of thumb is that a strong acid is 100% dissociated in solutions of 1.0 M or less.



Strong and Weak Acids and Bases

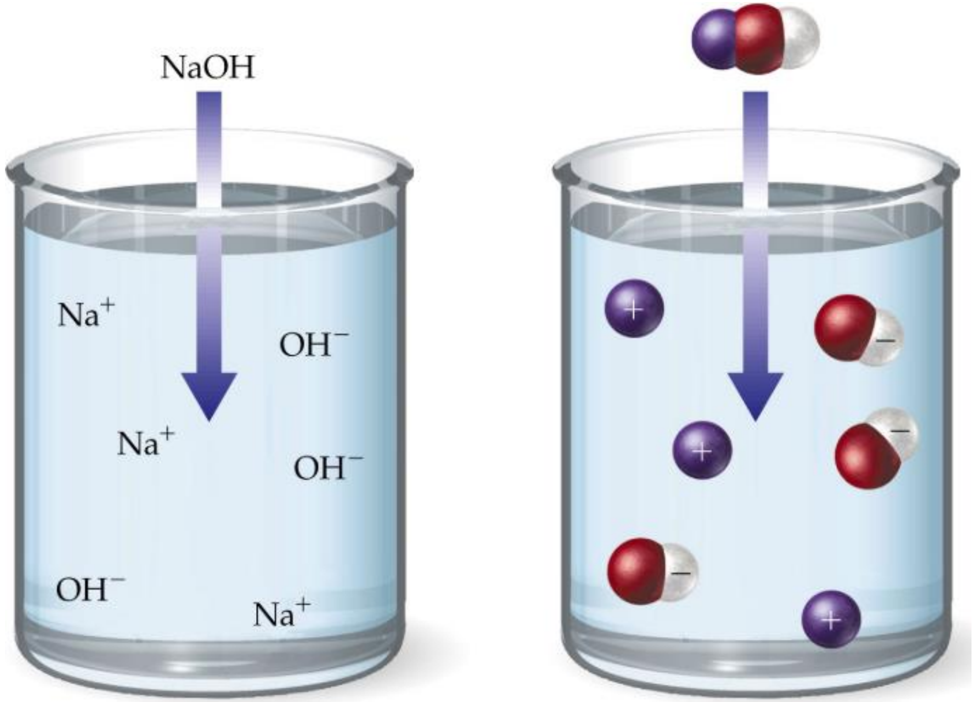
Topic 8.4

Strong acids are assumed to dissociate completely when in aqueous solution.

Weak acids dissociate only slightly in aqueous solution. The majority of molecules remain undissociated.

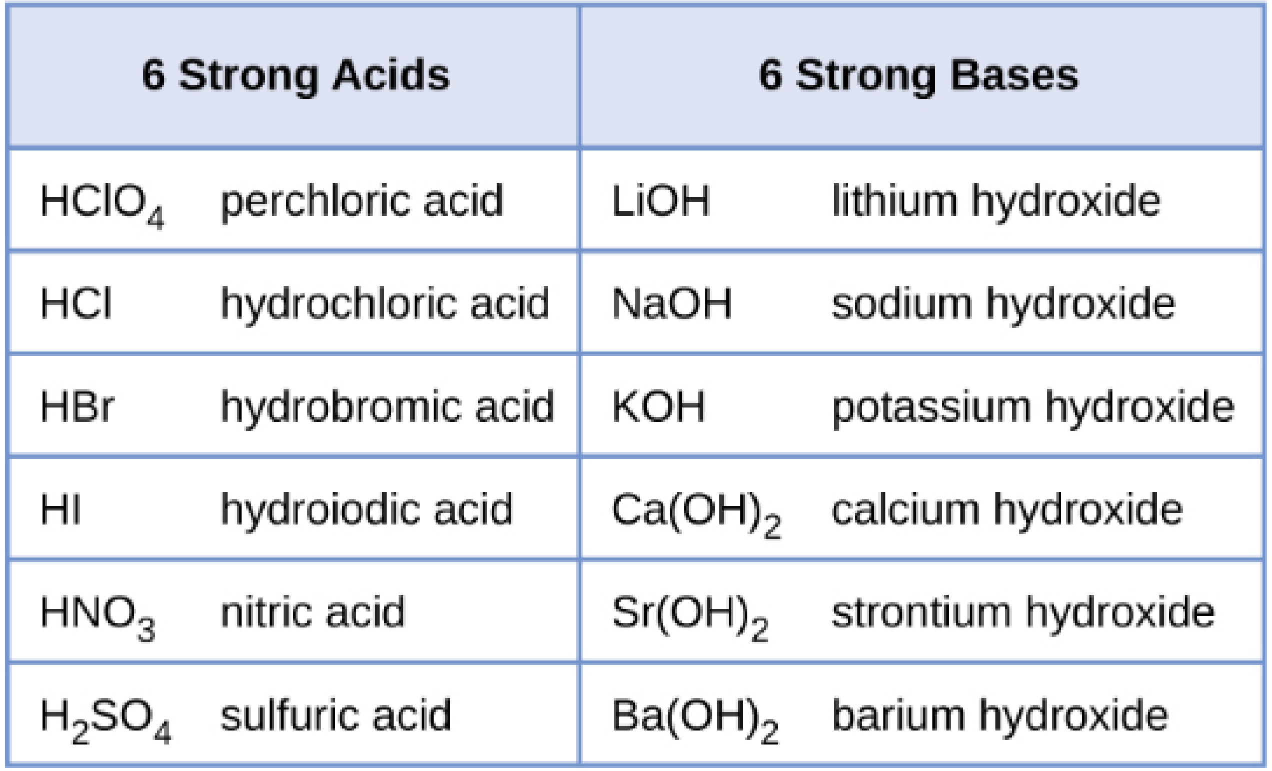
Common Strong Bases

- Strong bases are bases which completely dissociate in water into the cation and OH^- (hydroxide ion).



- The hydroxides of the Group I and Group II metals usually are considered to be strong bases.

Common Strong Acids and Bases



6 Strong Acids		6 Strong Bases	
$HClO_4$	perchloric acid	LiOH	lithium hydroxide
HCl	hydrochloric acid	NaOH	sodium hydroxide
HBr	hydrobromic acid	KOH	potassium hydroxide
HI	hydroiodic acid	$Ca(OH)_2$	calcium hydroxide
HNO_3	nitric acid	$Sr(OH)_2$	strontium hydroxide
H_2SO_4	sulfuric acid	$Ba(OH)_2$	barium hydroxide

Autoionization of Water

In **pure water** (no solute) water molecules behave as both an **acid** and **base**!!

It is called amphoteric meaning it will act as either an acid or a base depending on the situation.

$$H_2O\ (l) + H_2O\ (l) \longrightarrow H_3O^+\ (aq) + OH^-\ (aq)$$

This is called the **autoionization of** water. Although the equilibrium lies far to the **left** it is very important to take into consideration, especially for **living systems**.

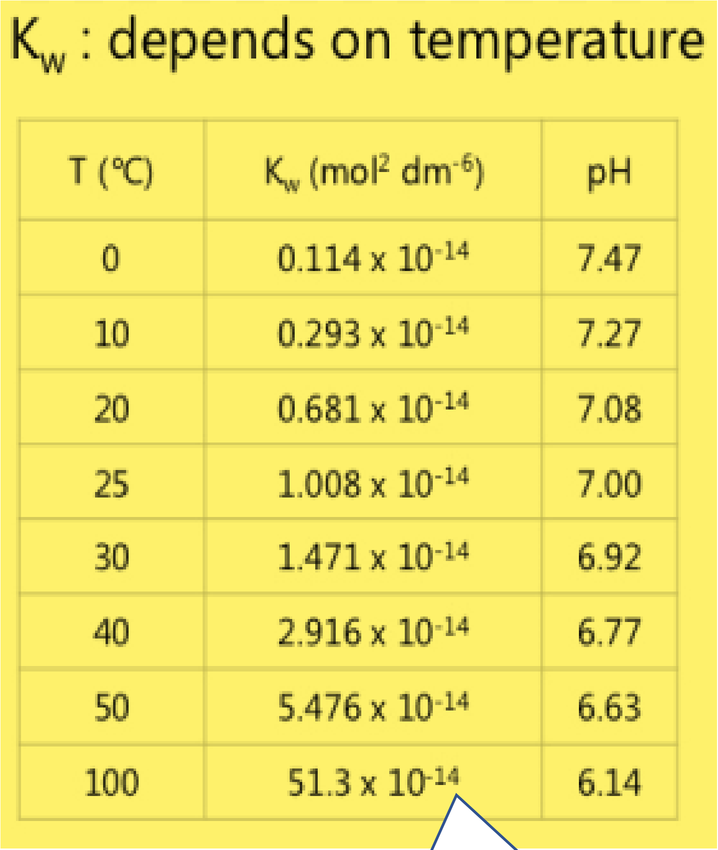
For pure water [OH-] = [H+]

K_w is called **ionization constant** of water and is very small. As with all Kw values, it is temperature dependent.

$K_w = 1.0 \times 10^{-14}$ @ 25°C

$$K_w = [H^+][OH^-]$$

$$K_w = (1 \times 10^{-7})(1 \times 10^{-7})$$



K_w : depends on temperature

T (°C)	K_w (mol² dm⁻⁶)	pH
0	0.114×10^{-14}	7.47
10	0.293×10^{-14}	7.27
20	0.681×10^{-14}	7.08
25	1.008×10^{-14}	7.00
30	1.471×10^{-14}	6.92
40	2.916×10^{-14}	6.77
50	5.476×10^{-14}	6.63
100	51.3×10^{-14}	6.14

This only means that the neutral value for pH is getting lower, it does not mean that the solution is becoming more acidic as the temperature increase.

We define an **aqueous solution** as being

- **neutral** when the $[H^+] = [OH^-]$
- **acidic** when $[H^+] > [OH^-]$
- **basic** when $[H^+] < [OH^-]$

CHECK THIS OUT !

$[H^+] = 0.0000001 = 10^{-7}$

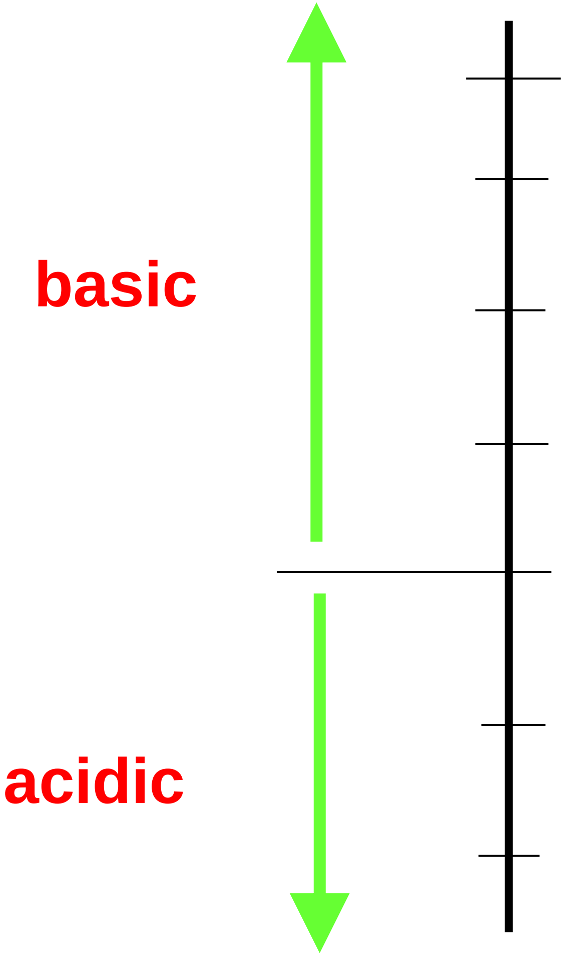
How can this be abbreviated further ?

By just describing the power called the **POWER OF H**

pH = 7

A pH Number line

The pH scale is a way of expressing the strength of acids and bases.



basic

acidic

pH = 14	$[H^+] = 10^{-14}$	$[OH^-] = 10^{0}$	$[H^+] < [OH^-]$	**basic**
pH = 12	$[H^+] = 10^{-12}$	$[OH^-] = 10^{-2}$	$[H^+] < [OH^-]$	**basic**
pH = 7	$[H^+] = 10^{-7}$	$[OH^-] = 10^{-7}$	$[H^+] = [OH^-]$	**neutral**
pH = 2	$[H^+] = 10^{-2}$	$[OH^-] = 10^{-12}$	$[H^+] > [OH^-]$	**acidic**

pH Calculations

To calculate pH or pOH

$pH = -\log [H^+]$, or $pOH = -\log [OH^-]$

$pH + pOH = 14$ for water solutions

Find the pH of these:

1) 0.15 M solution of Hydrochloric acid
2) 3.00×10^{-7} M solution of Nitric acid

Calculations for Molarity Concentrations from the PH

This is the
calculator
button

$M= 2^{nd} \log = 10^{-pH}$ $M= 2^{nd} \log = 10^{-pOH}$

The number of decimal places in the log answer is equal to the number of sig figs in the original Molar [ ]

Question: Find the Molarity:

A solution has a pH of 8.5. What is the Molarity of hydrogen ions in the solution?

Note: to be able to determine all the values associated with acids and base, you only need **one piece of information**.

$$\begin{tikzcd}[column sep=huge, row sep=huge]
H^+ \arrow[r, shift left, "{[OH^-] = \frac{1 \times 10^{-14}}{[H^+]}}"] \arrow[d, shift left, "{pH = -\log[H^+]}"] & OH^- \arrow[l, shift left, "{[H^+] = \frac{1 \times 10^{-14}}{[OH^-]}}"] \arrow[d, shift left, "{pOH = -\log[OH^-]}"] \\
pH \arrow[u, shift left, "{[H^+] = 10^{-pH}}"] \arrow[r, shift left, "{pOH = 14 - pH}"] & pOH \arrow[l, shift left, "{pH = 14 - pOH}"] \arrow[u, shift left, "{[OH^-] = 10^{-pOH}}"]
\end{tikzcd}$$

Buffered solutions – resists a change in its pH even when a strong acid or base is added to it

- A solution is buffered in the presence of a weak acid and its conjugate base

Example: Calculate the pH or pOH

a. $[H^+] = 1.0 \times 10^{-9}$ M

b. $[OH^-] = 1.0 \times 10^{-6}$ M

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

$pH = -\log (1.0 \times 10^{-9}\ M)$

$$\boxed{pH = 9.00}$$

$pOH = -\log [OH^-]$

$pOH = -\log (1.0 \times 10^{-6}\ M)$

$$\boxed{pOH = 6.00}$$

Example – Calculate the pH and pOH if the concentration of OH^{-1} is 1.0×10^{-3} M

$$pOH = -\log [OH^-]$$

$$pOH = -\log (1.0 \times 10^{-3}\ M)$$

$$pOH = 3.00$$

$$K_w = [H^+][OH^-]$$

$$1 \times 10^{-14} = [H^+][1.0 \times 10^{-3}\ M]$$

$$[H^+] = 1.0 \times 10^{-11}\ M$$

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

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

$$\boxed{pH = 11.00}$$

Example – Calculate the pH and pOH if the concentration of

a. $[H^+] = 1.0 \text{ x } 10^{-9}$ M

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

$\text{pH} = -\log (1.0 \text{ x } 10^{-9} \text{ M})$

$\boxed{\text{pH} = 9.00}$

b. $[OH^-] = 1.0 \text{ x } 10^{-6}$ M

$\text{pOH} = -\log [OH^-]$

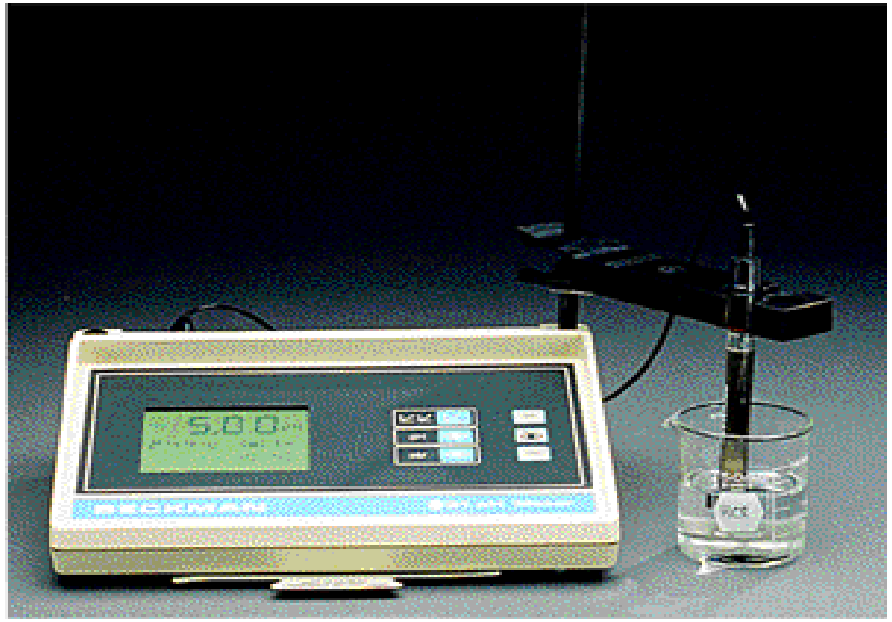
$\text{pOH} = -\log (1.0 \text{ x } 10^{-6} \text{ M})$

$\boxed{\text{pOH} = 6.00}$

Methods for measuring the PH of an aqueous solutions.

Indicators, including: (A) Litmus paper, are used for less accurate measurements; an indicator is one color in its acid form and another color in its basic form.

pH meters are used for accurate measurement of pH; electrodes indicate small changes in voltage to detect the pH.



Section Quiz.

3. Diluting a solution does NOT change which of the following?

   A. concentration
   B. volume
   C. milliliters of solvent
   D. moles of solute

The End of Lecture