Please watch the lecture video in canvas

GENERAL SOLUTION CHEMISTRY

PROPERTIES
OF
SOLUTIONS

GENERAL PROPERTIES OF SOLUTIONS

1. A solution is composed of:

    the **solute**: the minor component (least number of moles)

    the **solvent**: the major component (largest number of moles)

2. **Soluble / Insoluble:** A soluble substance readily dissolves in the solvent. An insoluble substance will NOT dissolve readily in a solvent.

3. **Miscible / immiscible:** Two liquids are miscible in each other if they readily mix to form a uniform solution. Two immiscible liquids will always separate out into two distinct layers.

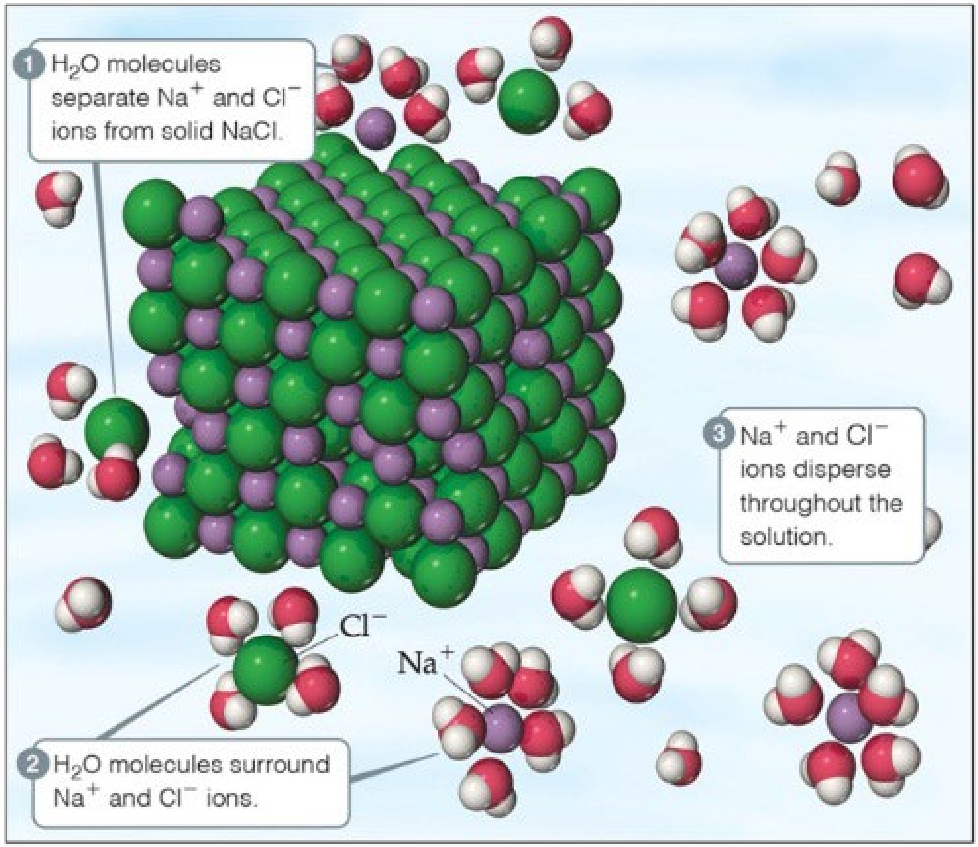
4. **Solubility** describes the amount of solute that will dissolve in a solvent. For example, 35.7 g of NaCl will dissolve in 100 mL of water at 0°C , no more.

GENERAL PROPERTIES OF SOLUTIONS

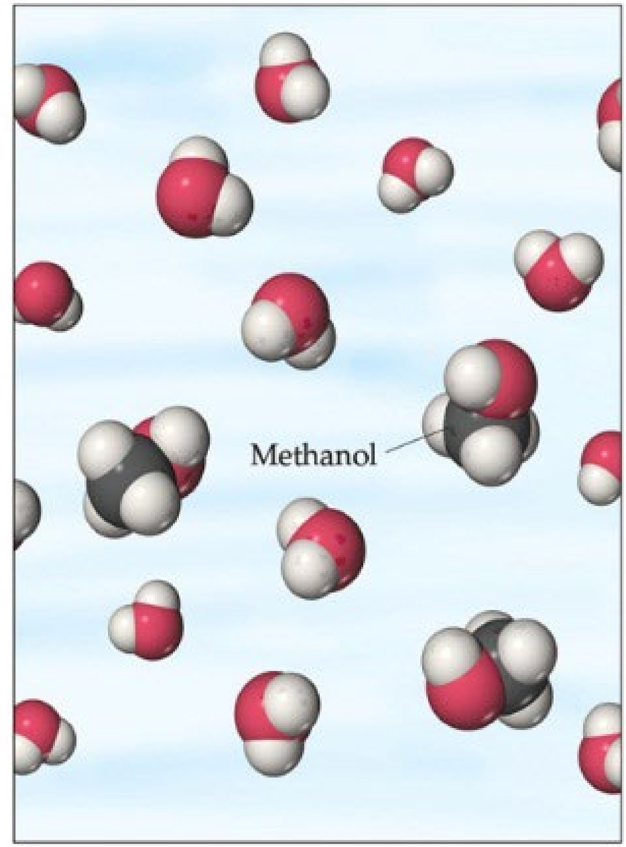
1. A solution is a homogeneous mixture of two or more components.

2. It has variable composition.
3. The dissolved solute is molecular or ionic in size.
4. A solution may be either colored or colorless but is generally transparent.
5. The solute remains uniformly distributed throughout the solution and will not settle out through time.
6. The solute can be separated from the solvent by physical methods.

Aqueous Solutions



(a) Ionic compounds like sodium chloride, NaCl, form ions when they dissolve.

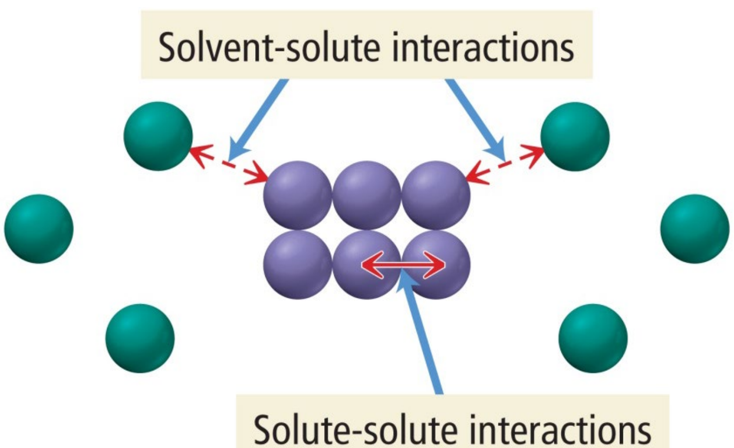


(b) Molecular substances like methanol, CH_3OH, dissolve without forming ions.

- Substances can dissolve in water by different ways:
  - Ionic compounds dissolve by **dissociation**, where water surrounds the separated ions.
  - Molecular compounds interact with water, but most do NOT dissociate.
  - Some molecular substances react with water when they dissolve.
    - All substances dissolve by **solvation**, surrounding of the solute by solvent.

What Happens When a Solute Dissolves?

- there are attractive forces between the solute particles holding them together
- there are also attractive forces between the solvent molecules
- when we mix the solute with the solvent, there are attractive forces between the solute particles and the solvent molecules
- if the attractions between solute and solvent are strong enough, the solute will dissolve



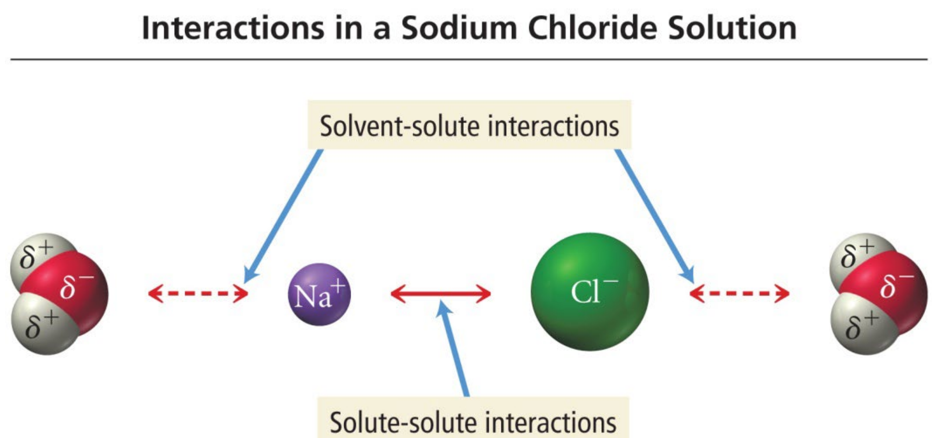
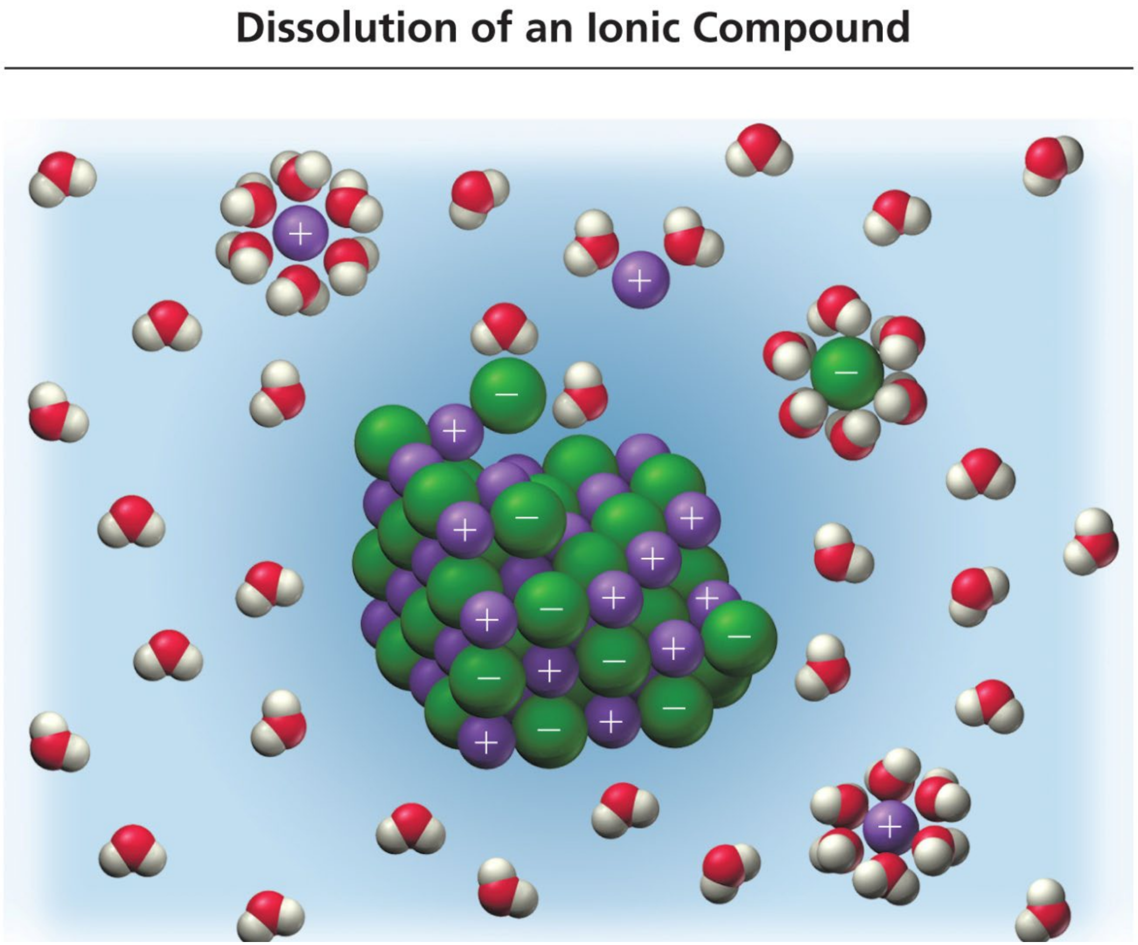
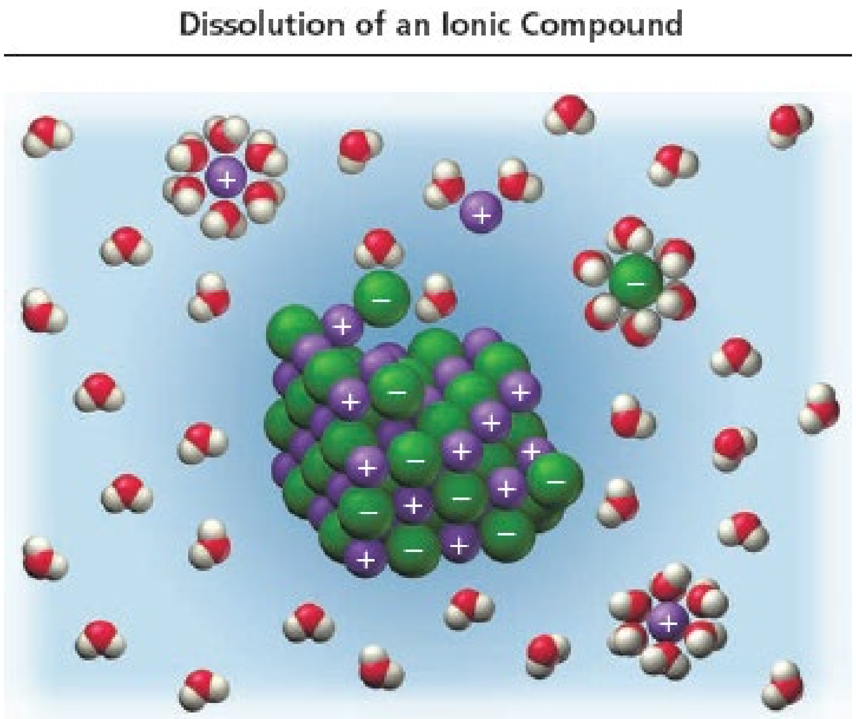


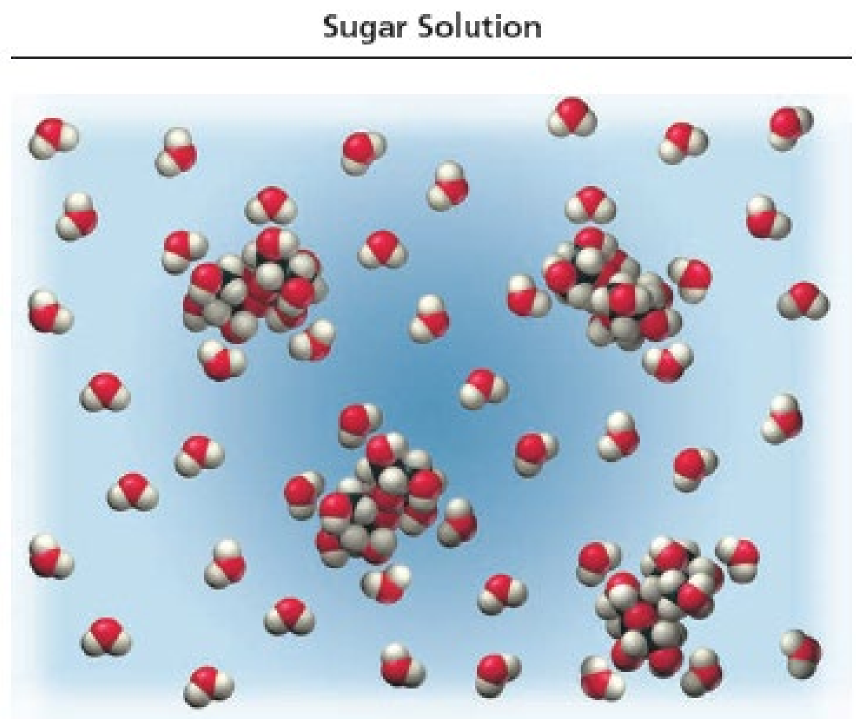
Table Salt Dissolving in Water



- Each ion is attracted to the surrounding water molecules and pulled off and away from the crystal
- When it enters the solution, the ion is surrounded by water molecules, insulating it from other ions
- The result is a solution with free moving charged particles able to conduct electricity

Salt vs. Sugar Dissolved in Water





ionic compounds dissociate into ions when they dissolve

molecular compounds do not dissociate when they dissolve

ELECTROLYTES

Electrolytes are species which conducts electricity when dissolved in water. Acids, Bases, and Salts are all electrolytes.

Salts and strong Acids or Bases form Strong Electrolytes. Salt and strong acids (and bases) are fully dissociated therefore all of the ions present are available to conduct electricity.

$$HCl_{(s)} + H_2O \rightarrow H_3O^+ + Cl^-$$

Weak Acids and Weak Bases for Weak Electrolytes.

Weak electrolytes are partially dissociated therefore not all species in solution are ions, some of the molecular form is present. Weak electrolytes have less ions avalible to conduct electricity.

$$NH_3 + H_2O \Leftrightarrow NH_4^+ + OH^-$$

Electrolytes

Pure water does not conduct electricity.

A **nonelectrolyte solution** does not conduct electricity.

An **electrolyte solution** conducts electricity.

Pure water, $H_2O(l)$

Sucrose solution, $C_{12}H_{22}O_{11}(aq)$

Sodium chloride solution, $NaCl(aq)$

- A **strong electrolyte** dissociates completely when dissolved in water.
- A **weak electrolyte** only dissociates partially when dissolved in water.
- A **nonelectrolyte** does NOT dissociate in water.

Solubility of Ionic Compounds

- Not all ionic compounds dissolve in water.
- A list of **solubility rules** is used to decide what combination of ions will dissolve.

Table 4.1 Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds		**Important Exceptions**
Compounds containing	NO_3^-	None
	CH_3COO^-	Ag^+ Slightly soluble (see chem 102)
	Cl^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	Br^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	I^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
	SO_4^{2-}	Compounds of Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Pb^{2+}

Insoluble Ionic Compounds		**Important Exceptions**
Compounds containing	S^{2-}	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}
	CO_3^{2-}	Compounds of NH_4^+, and the alkali metal cations
	PO_4^{3-}	Compounds of NH_4^+, and the alkali metal cations
	OH^-	Compounds of NH_4^+, the alkali metal cations, Ca^{2+}, Sr^{2+}, and Ba^{2+}

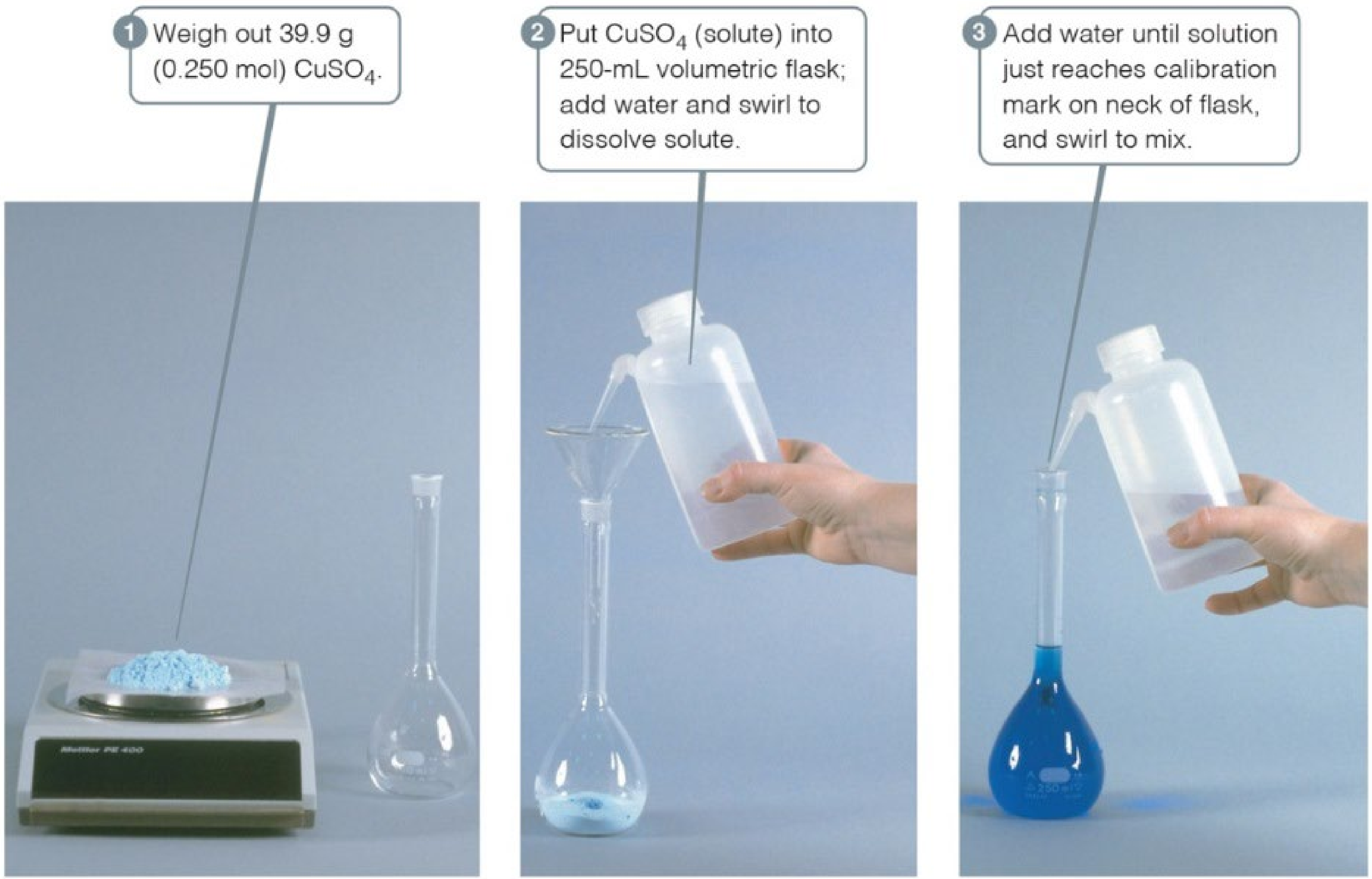
GENERAL SOLUTION CHEMISTRY

MOLARITY

Watch the lecture video on Molarity basics! M = n/V

Mixing a Solution

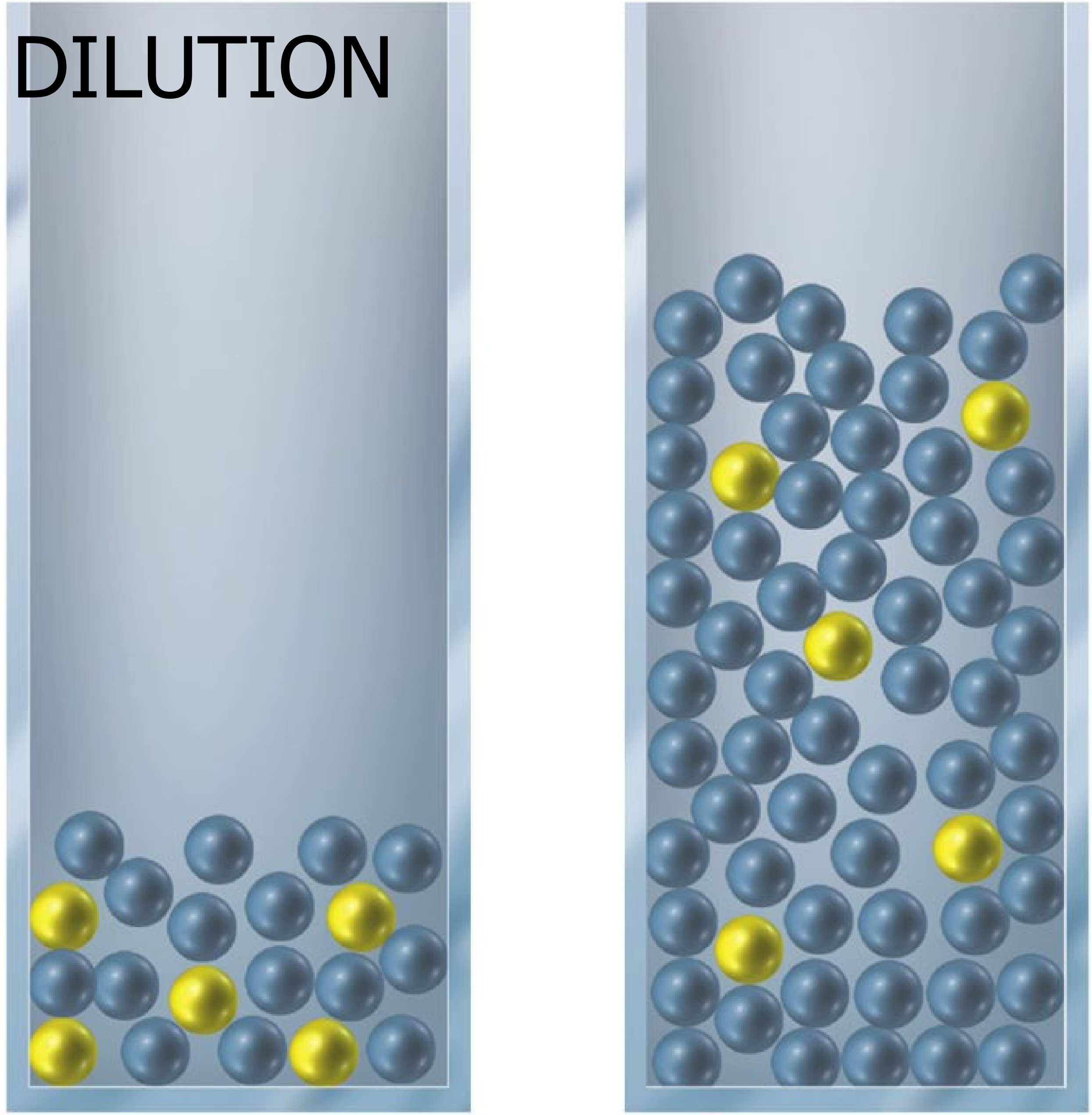
- To create a solution of a known molarity, weigh out a known mass (and, therefore, number of moles) of the solute.
- Then add solute to a volumetric flask, and add solvent to the line on the neck of the flask.



MOLARITY & DILUTION

$$M_1V_1 = M_2V_2$$

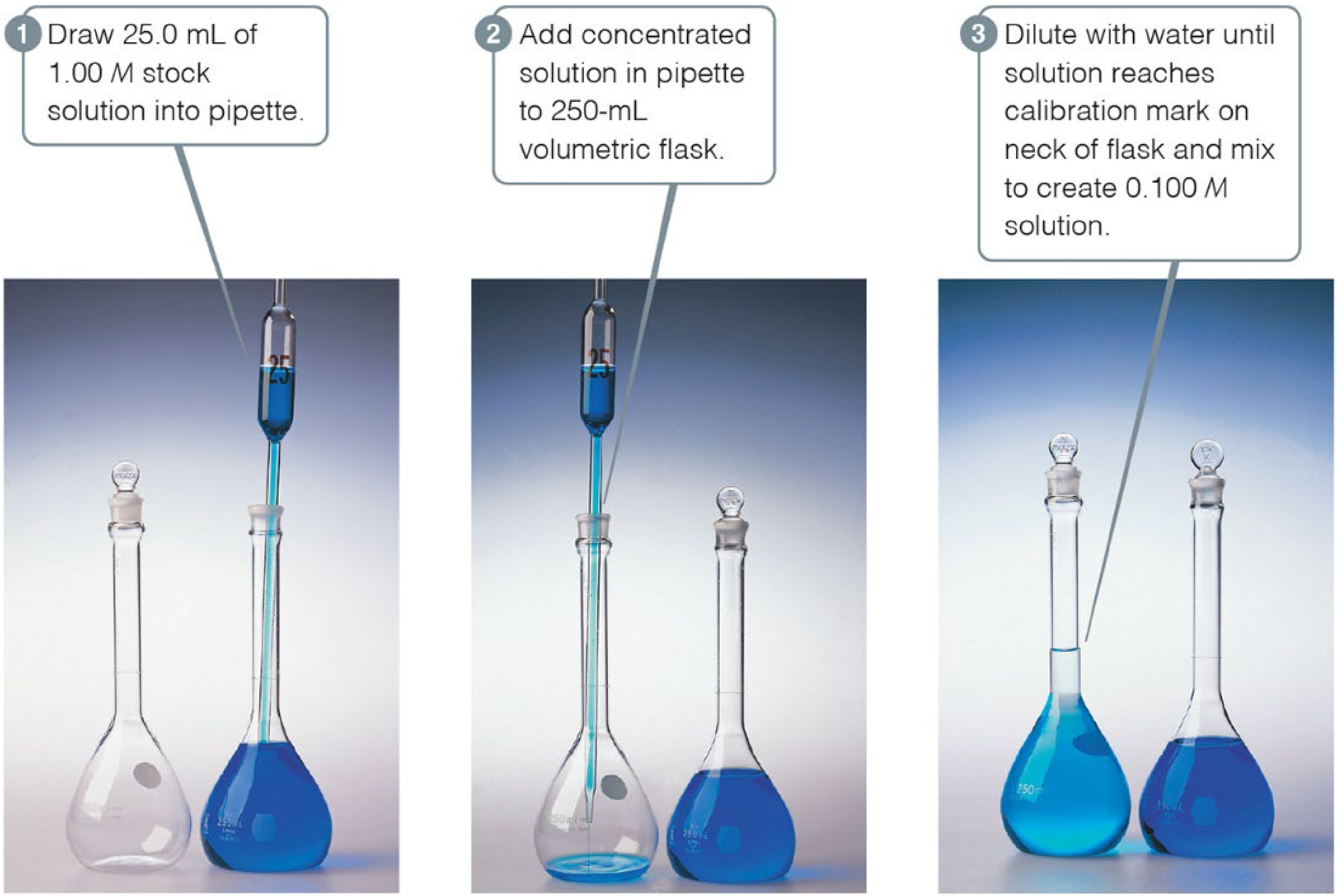
The act of diluting a solution is to simply add more water (or solvent) thus leaving the amount of solute unchanged.

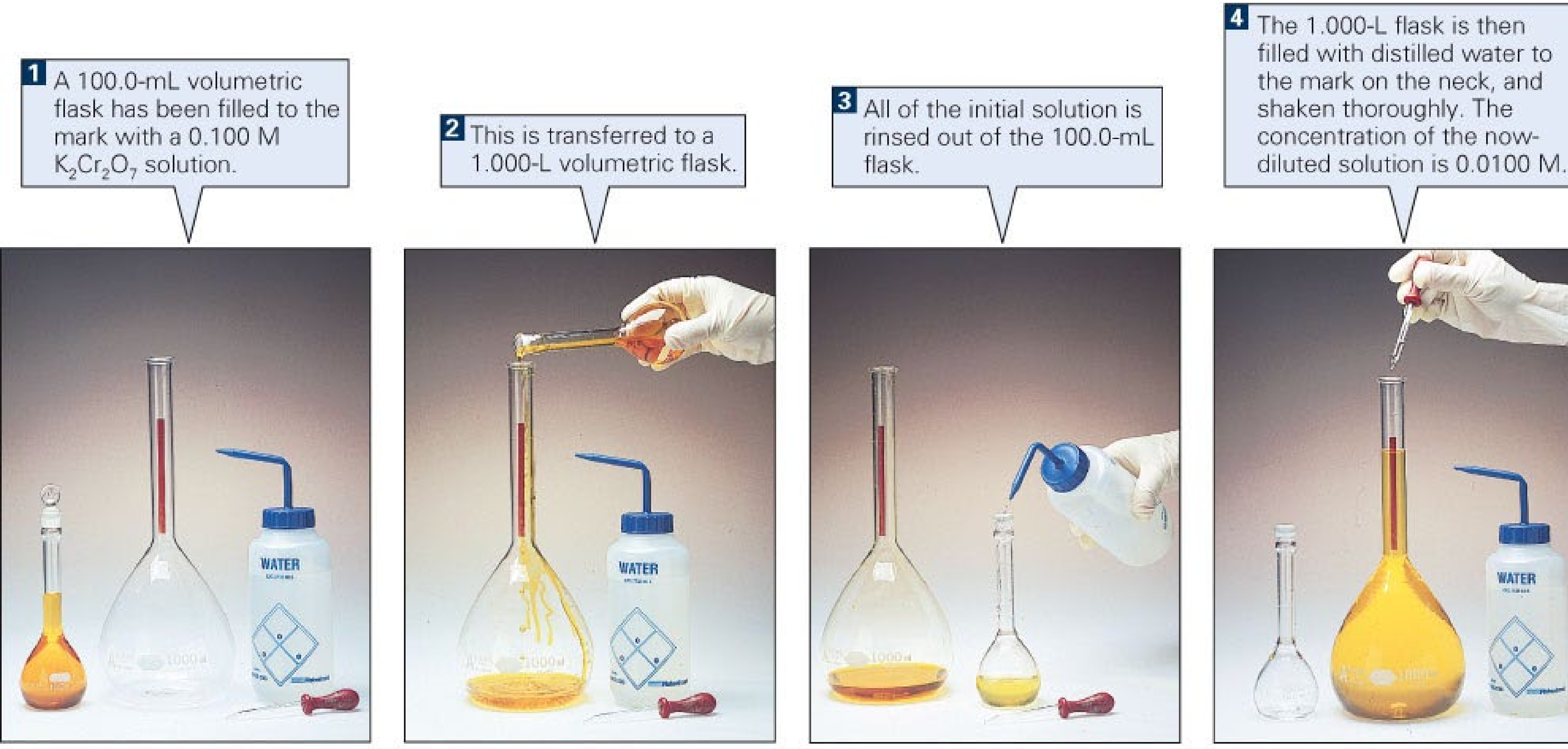


(a) (b)

Dilution

A solution can be **diluted** by adding ONLY solvent. The concentration is LOWER, but the MOLES don't change.





Example: Calculate the concentration of a $K_2Cr_2O_7$(aq) made by diluting 20.0 mL of 2.55 M with 235.0 mL of water. **0.20M**

Describe in detail how you would prepare 0.500 liters of 0.456 M Li_3PO_4 solution from (a) 6.00 M Li_3PO_4 and (b) from pure solid Li_3PO_4

(a) $M_1V_1 = M_2V_2$

$(0.456M)(0.5L) = (6.00M)\ V_2$

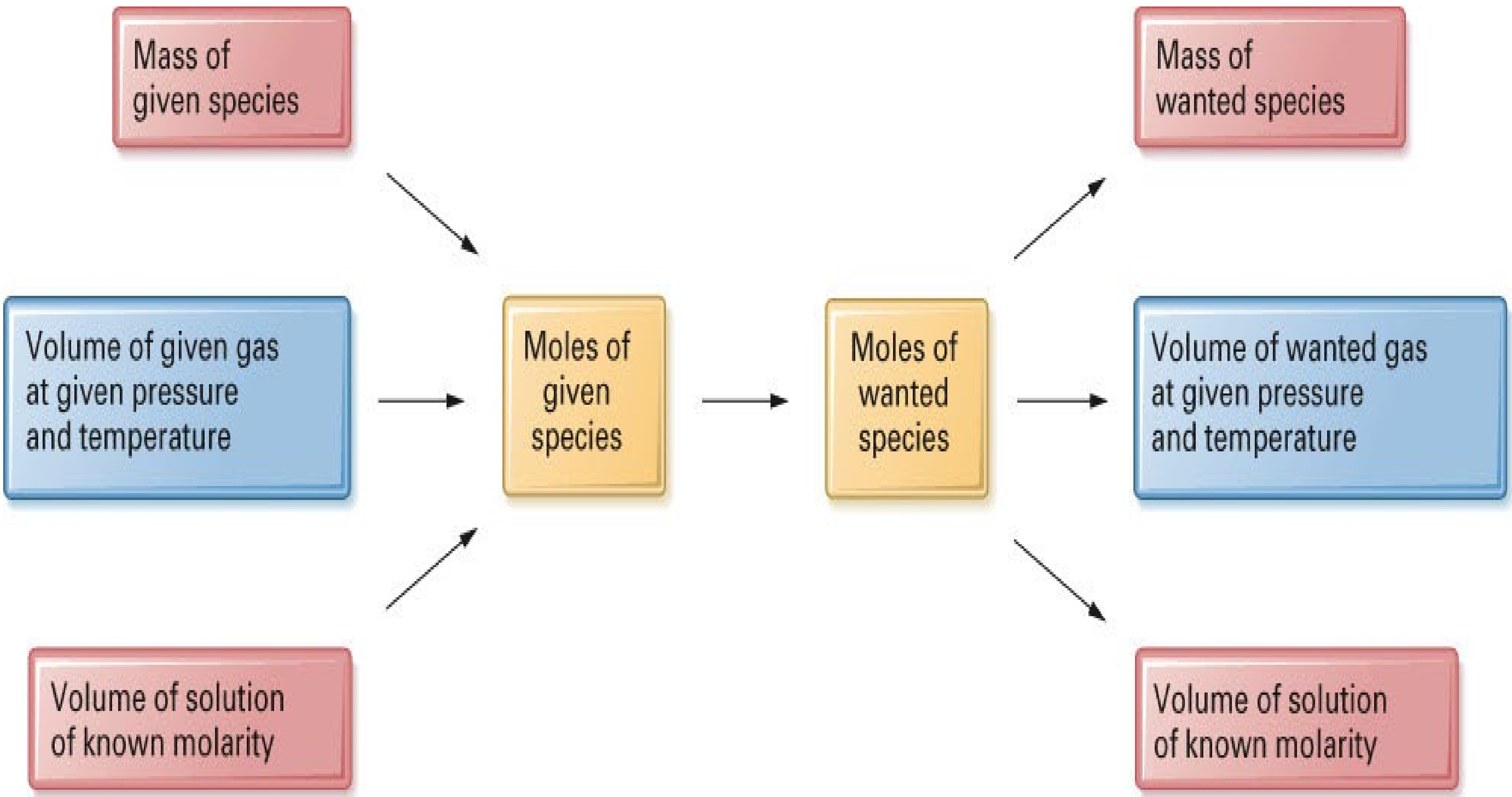
$V_2 = 0.038L = 38.0\ mL$

Add 38.0 mL of 6.00M Li_3PO_4 to a flask. Add enough DI water to bring the volume to 500 mL (0.500L)

(b) $$0.500\cancel{L}\left(\frac{0.456\ \cancel{mole}}{1\ \cancel{L}}\right)\left(\frac{115.79\ g\ Li_3PO_4}{1\ \cancel{mole}}\right) = 26.4\ g\ Li_3PO_4$$

Dissolve 26.4 grams of solid Li_3PO_4 in DI water, then add enough DI water to bring the volume to 500 mL.

Solution Stoichiometry



MOLARITY & Stoichiometry

How much calcium carbonate will be precipitated by adding 25.0 mL calcium chloride to 25.0 mL of 0.56 M potassium carbonate?

$$CaCl_2 + K_2CO_3 \rightarrow CaCO_3 + 2\ KCl$$

V = 25.0 mL | **M = 0.56 mol/L** | **m=?**

V = 25.0 mL

First convert volume of A to moles of A:

0.025 L K_2CO_3 (0.56 mol/L) = 0.014 moles of K_2CO_3

Now convert moles of A to moles of B:

0.014 mol K_2CO_3(1 mol $CaCO_3$/1 mol K_2CO_3) =0.014 mol $CaCO_3$

Next convert moles of B to grams of B:

0.014 mol $CaCO_3$ (100 g/mol) = 1.40 g of $CaCO_3$

MOLARITY & Stoichiometry

What would be the molarity of the potassium chloride solution from the last problem?

$$CaCl_2 + K_2CO_3 \rightarrow CaCO_3 + 2\ KCl$$

V = 25.0 mL   M = 0.56 mol/L   M = ?

V = 25.0 mL

First convert volume of A to moles of A:

0.025 L K_2CO_3 (0.56 mol/L) = 0.014 moles of K_2CO_3

Now convert moles of A to moles of B:

0.014 mol K_2CO_3(2 mol KCl /1 mol K_2CO_3) =0.028 mol KCl

Next convert moles of B to molarity of B:

0.028 mol KCl / 0.050 L = 0.56 M of KCl

MOLARITY & Stoichiometry

2. If 30.0 g of solid zinc are treated with 3.13 L 0.200 M HCl,

a) how many grams of hydrogen gas will theoretically be formed?

b) What is the concentration of the remaining solution?

Zn	+	$2HCl$	$\rightarrow$	$ZnCl_2$	+	H_2
30.0 g		3.13L		M= ?		m=?
$65.38\ ^{g}/_{moL}$		0.200M				$2\ ^{g}/_{moL}$

a) $30.0\text{ g Zn}\left(^{1\text{mol Zn}}/_{65.38\text{g Zn}}\right)\left(^{1\text{mol H}_2}/_{1\text{mol Zn}}\right)\left(^{2\text{g}}/_{1\text{mol}}\right) = \cancel{0.918\text{ g H}_2}$

$3.13\text{ L}\left(^{0.200\text{ mol HCl}}/_{1\text{ L}}\right)\left(^{1\text{mol H}_2}/_{2\text{mol HCl}}\right)\left(^{2\text{g}}/_{1\text{mol}}\right) = 0.626\text{ g H}_2$

b) $30.0\text{ g Zn}\left(^{1\text{mol Zn}}/_{65.38\text{g Zn}}\right)\left(^{1\text{mol ZnCl}_2}/_{1\text{mol Zn}}\right) = \cancel{0.458}\text{ mol}$

$3.13\text{ L}\left(^{0.200\text{ mol HCl}}/_{1\text{ L}}\right)\left(^{1\text{mol ZnCl}_2}/_{2\text{mol HCl}}\right) = 0.313\text{ mol}$

$M = n/V$ = moles of solute / total volume solution

$M = 0.313\text{ moles}/3.13\text{L} = 0.100\text{M ZnCl}_2$ solution

TITRATION

Titration of a strong acid with a strong base

ENDPOINT = POINT OF NEUTRALIZATION = EQUIVALENCE POINT

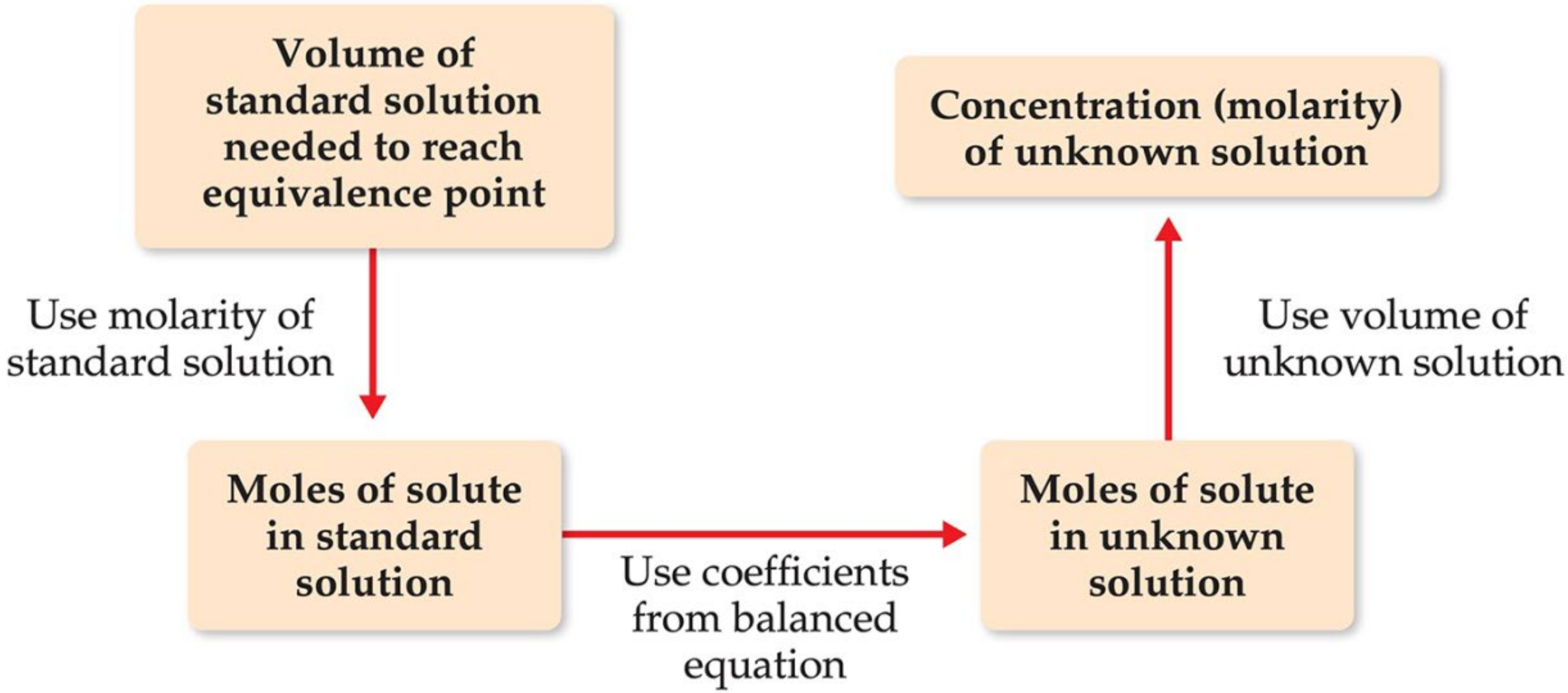
At the end point for the titration of a strong acid with a strong base, the moles of acid (H^+) equals the moles of base (OH^-) to produce the neutral species water (H_2O). If the mole ratio in the balanced chemical equation is NOT 1:1 then you must rely on the mole relationship and handle the problem like any other stoichiometry problem.

MOLES OF ACID = MOLES OF BASE

$$n_{acid} = n_{base}$$

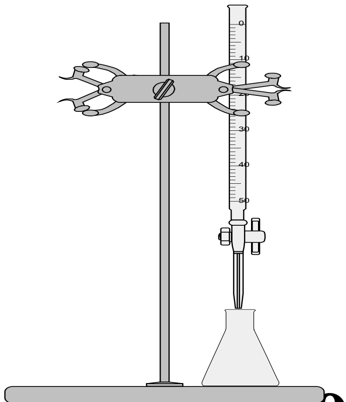
Remember: M = n/V

Titration



- A solution of known concentration, called a **standard solution**, is used to determine the unknown concentration of another solution.
- The reaction is complete at the **equivalence point**, which is based on the seen **end point** (color change).

Acid-Base Titrations



Acid-base titrations are an example of volumetric analysis, a technique in which one solution is used to analyze another. The solution used to carry out the analysis is called the titrant and is delivered from a device called a buret, which measures the volume accurately. The point in the titration at which enough titrant has been added to react exactly with the substance being determined is called the equivalence point (or stoichiometric point). This point is often marked by the change in color of a chemical called an indicator. The titration set-up is illustrated in the schematic shown above.

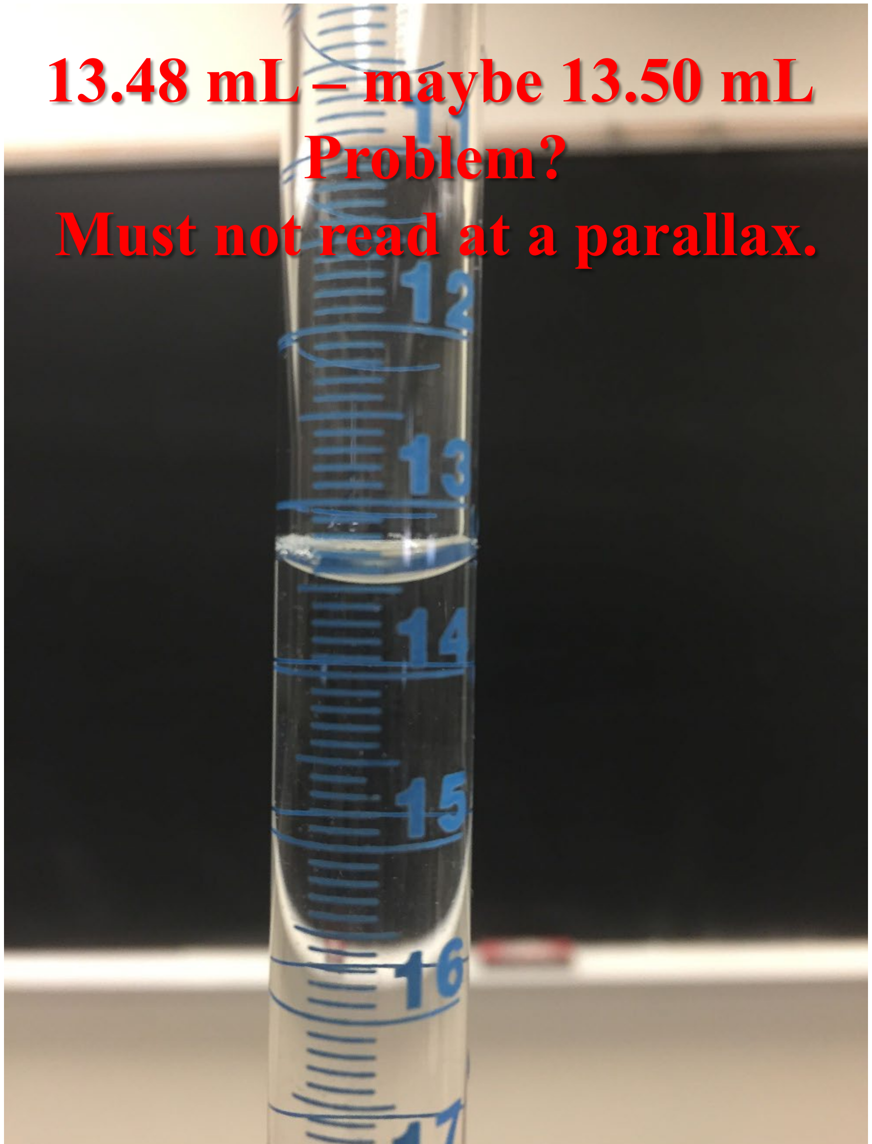
Acid-Base Titrations

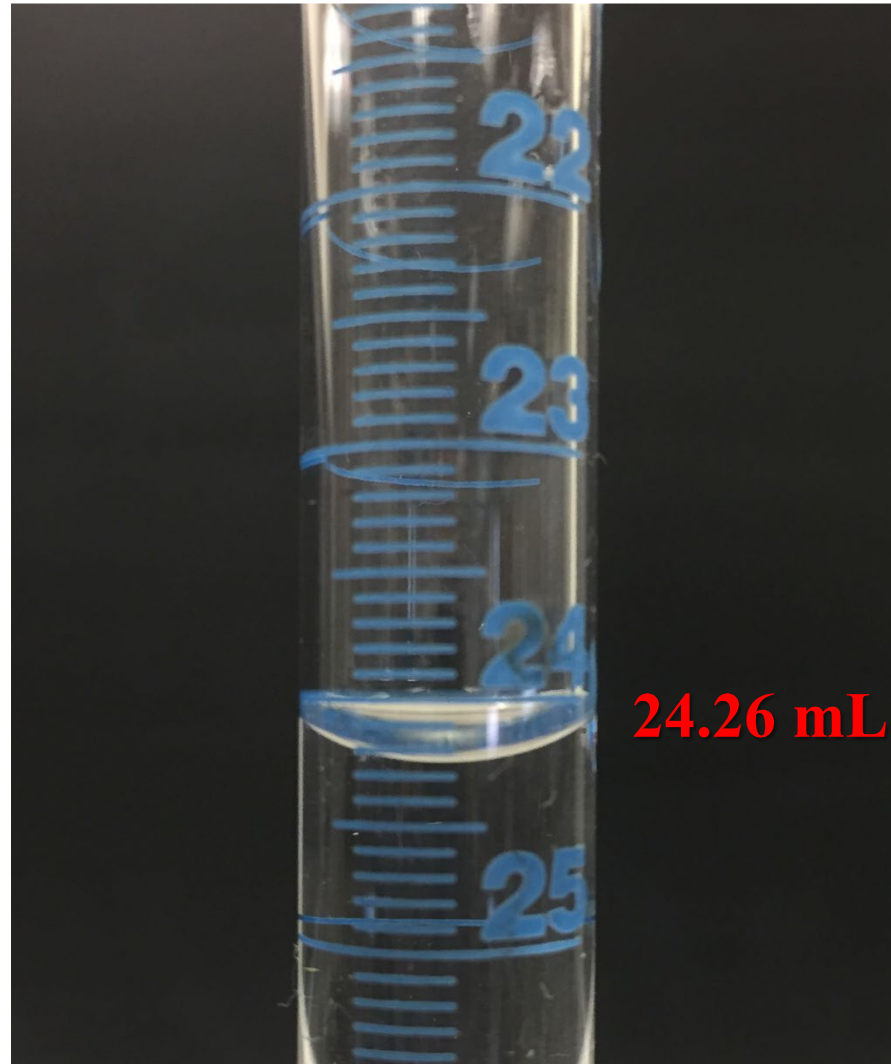
The following requirements must be met in order for a titration to be successful:

1. **The concentration of the titrant must be known (called the standard solution).**
2. **The exact reaction between the titrant and reacted substance must be known.**
5. **The volume of titrant required to reach the equivalence point must be known (measured) as accurately as possible.**

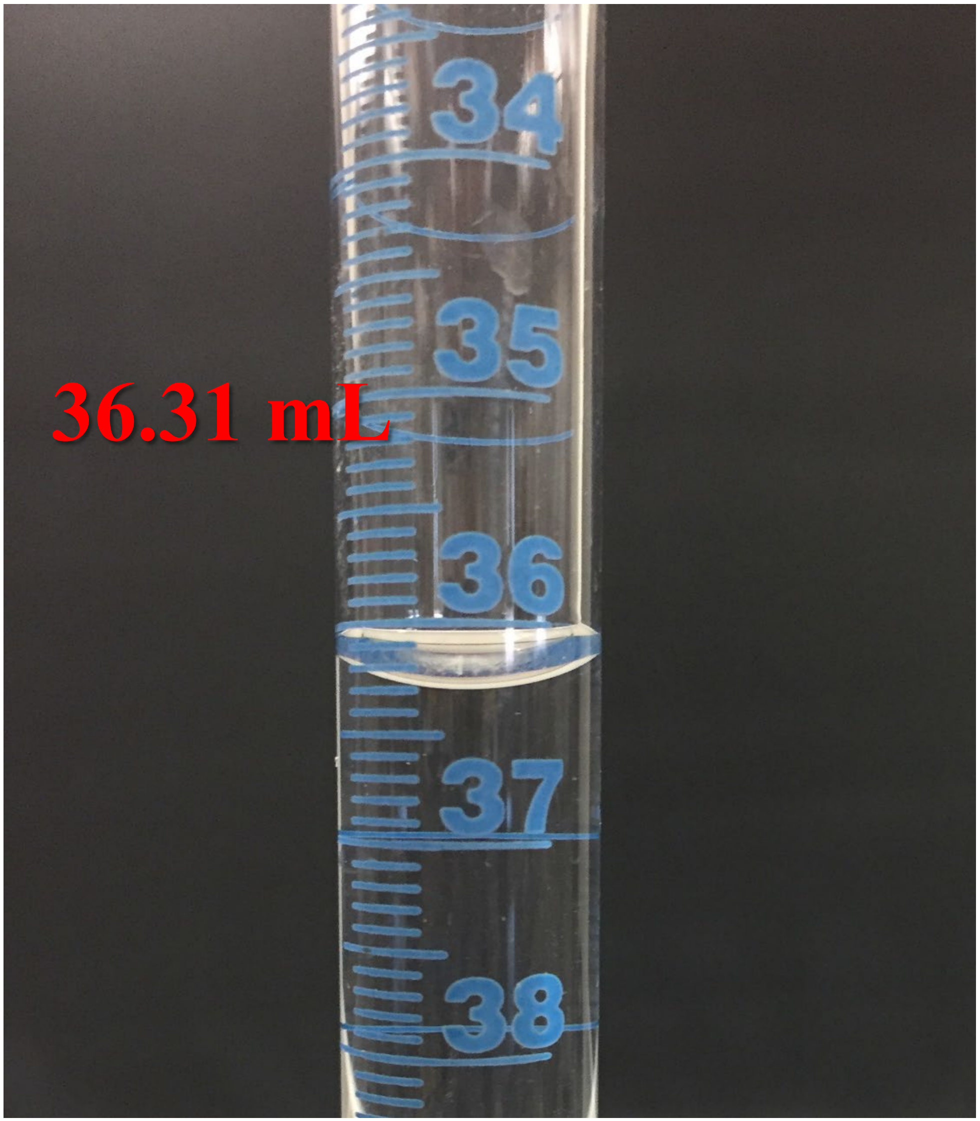
Reading a burette

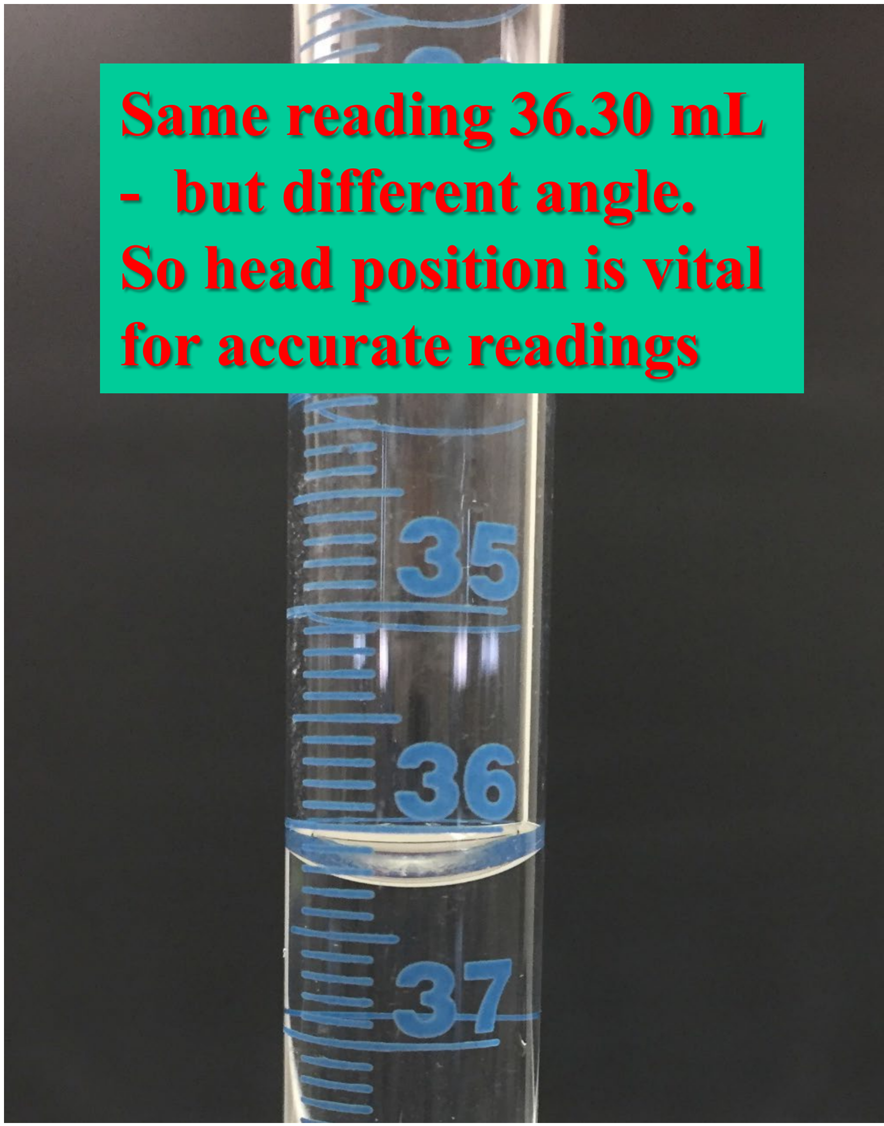
What value would you read for these burettes?



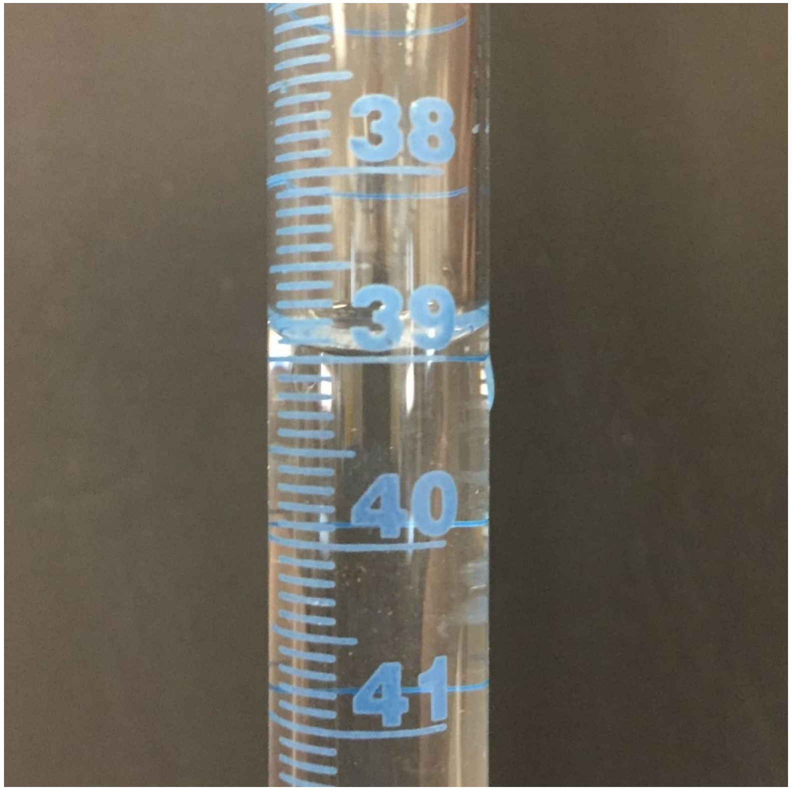


Reading a burette - 2

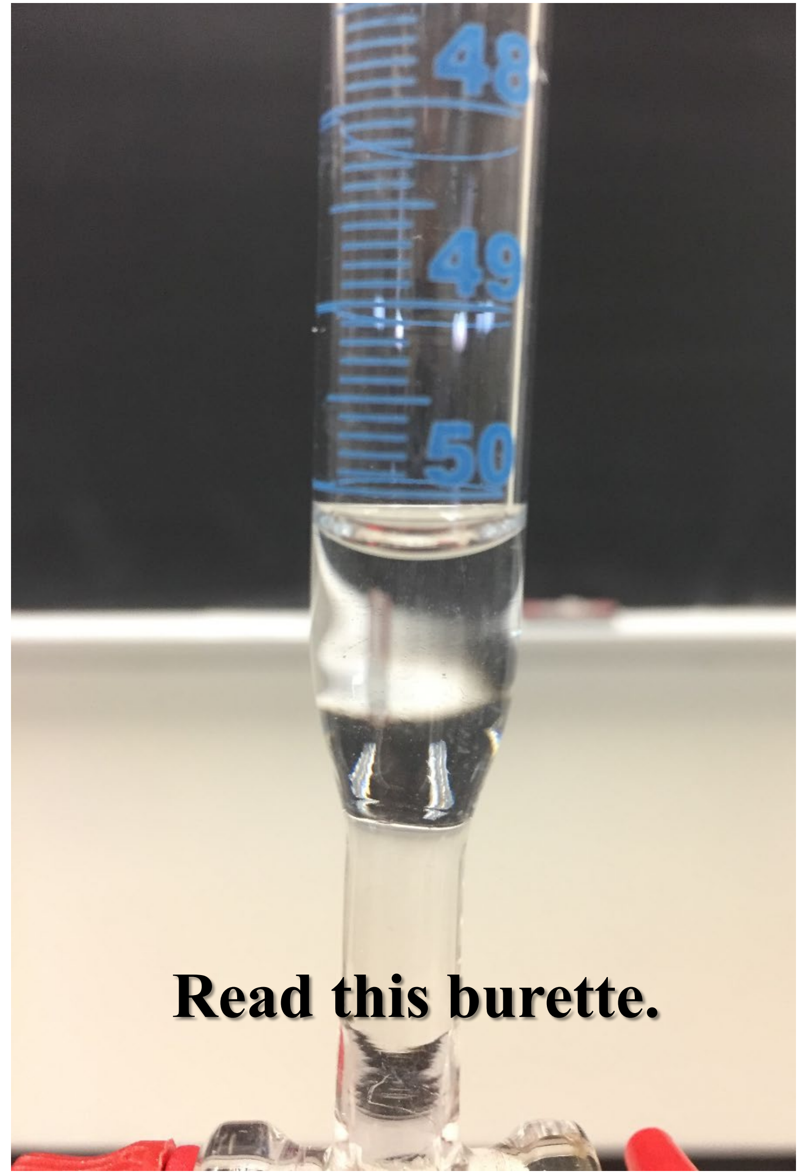




Reading a burette - QUIZ



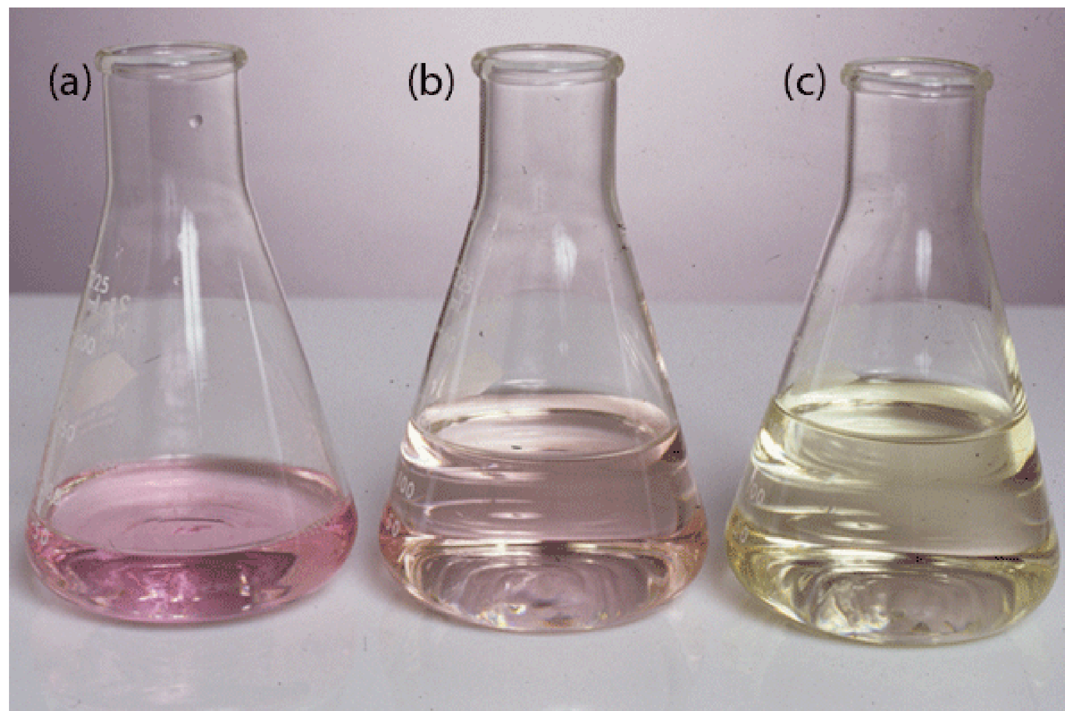
a) How would you read this?



Acid-Base Titrations

The following requirements must be met in order for a titration to be successful:

3. **The equivalence point must be known. An indicator that changes color at, or very near, the equivalence point is often used.**
4. **The point at which the indicator changes color is called the end point. The goal is to choose an indicator whose end point coincides with the equivalence point.**



NOTE: *Equivalence Point ≠ End Point!* WHY???

Acid-Base Titrations

When a substance being analyzed contains an acid, the amount of acid present is usually determined by titration with a standard solution containing hydroxide ions. The pH at certain points in the titration can be taken using different indicators, or alternatively, a pH meter can be used to give a readout of the exact pH.

$$pH = - \log [H_3O^+]$$

pH > 7 is referred to as a base

pH < 7 is referred to as an acid

What is the pH of a 0.00122 M HCl solution? pH = 2.914

If the hydronium ion concentration is 3.256×10^{-5}, what is its pH?
pH = 4.4873

Acid – Base Titration

1. **What is the molarity of an NaOH solution if 48.0 mL is needed to neutralize 35.0 mL of 0.144 M H_2SO_4?**

$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$$

M=?	**0.144M**
48.0 mL	**35.0 mL**

Step 2: convert volume of acid to moles using its molarity then use the mole ratio to convert from mole of acid to moles of base:

$$0.0350\ L\ H_2SO_4\ \left(^{0.144\ mol\ H_2SO_4}/_{L}\right)\left(^{2\ mol\ NaOH}/_{1 mol\ H_2SO_4}\right) =$$
$$= 0.01008\ mol\ NaOH$$

Step 3: use the definition of molarity to find the concentration:

$$M = n/V = 0.01008\ mol / 0.048\ L = 0.210\ M\ NaOH$$

Acid – Base Titration

2. **What is the concentration of 25.00 mL phosphoric acid if it took 32.60 mL of 0.01770 M calcium hydroxide.**

$$3\ Ca(OH)_2 \quad + \quad 2H_3PO_4 \rightarrow \quad Ca_3(PO_4)_2 \quad + \quad 6\ H_2O$$

0.0177 M	**M =?**
32.6 mL	**25.0 mL**

$$0.0326\ L\ Ca(OH)_2\ (^{0.177\ mol\ Ca(OH)2}/L)\ (^{2mol\ H3PO4}/_{3mol\ Ca(OH)_2}) =$$

$$= 0.003847 mol\ H_3PO_4$$

$$M = n/V = 0.003847\ mol / 0.025\ L = \color{red}{0.1539\ M\ H_3PO_4}$$

HOME PRACTICE PROBLEMS

2.4 M 1. What is the concentration of 250.0 mL of 0.60 moles of HCl?

3.00 L 2. What volume of 0.7690 M LiOH will contain 55.3 g of LiOH?

1.80 L 3. How many liters of water must be added to 100.0 mL of 4.50 M HBr to make a solution that is 0.250 M HCl?

39.6 g 4. How many grams of barium sulfate that will precipitate when 500.0 mL of 0.340 M $BaCl_2$ and 300.0 mL of 1.70 M Na_2SO_4 are mixed?

5. How would you prepare 850.0 mL of a 0.020 M ferric chloride solution if you start with crystals of $FeCl_3 \cdot 6H_2O$?

Weigh out 4.59 g of the hydrated salt to a graduated cylinder with 800.0 mL then add enough DI water to make 850 mL exactly.