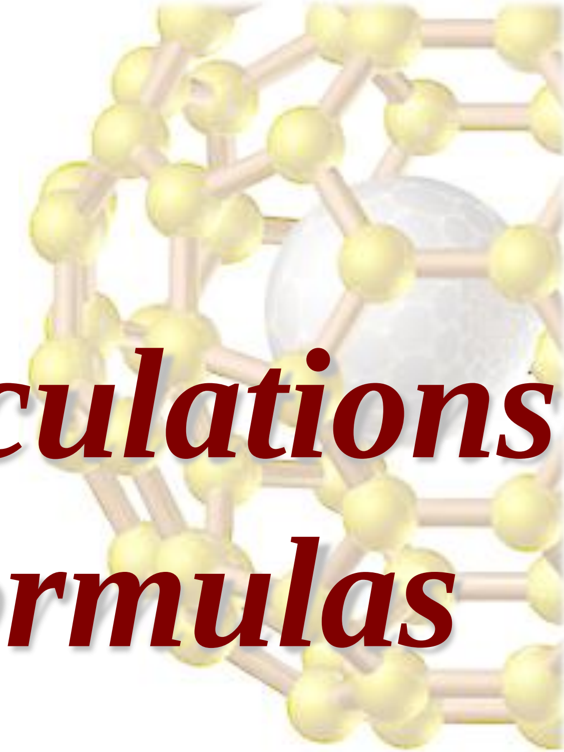
Stoichiometry: Calculations with Chemical Formulas and Equations



Chemical Equations

- **Lavoisier: mass is conserved in a chemical reaction.**
- **Chemical equations: descriptions of chemical reactions.**
- **Two parts to an equation: reactants and products:**

$$2H_2 + O_2 \rightarrow 2H_2O$$

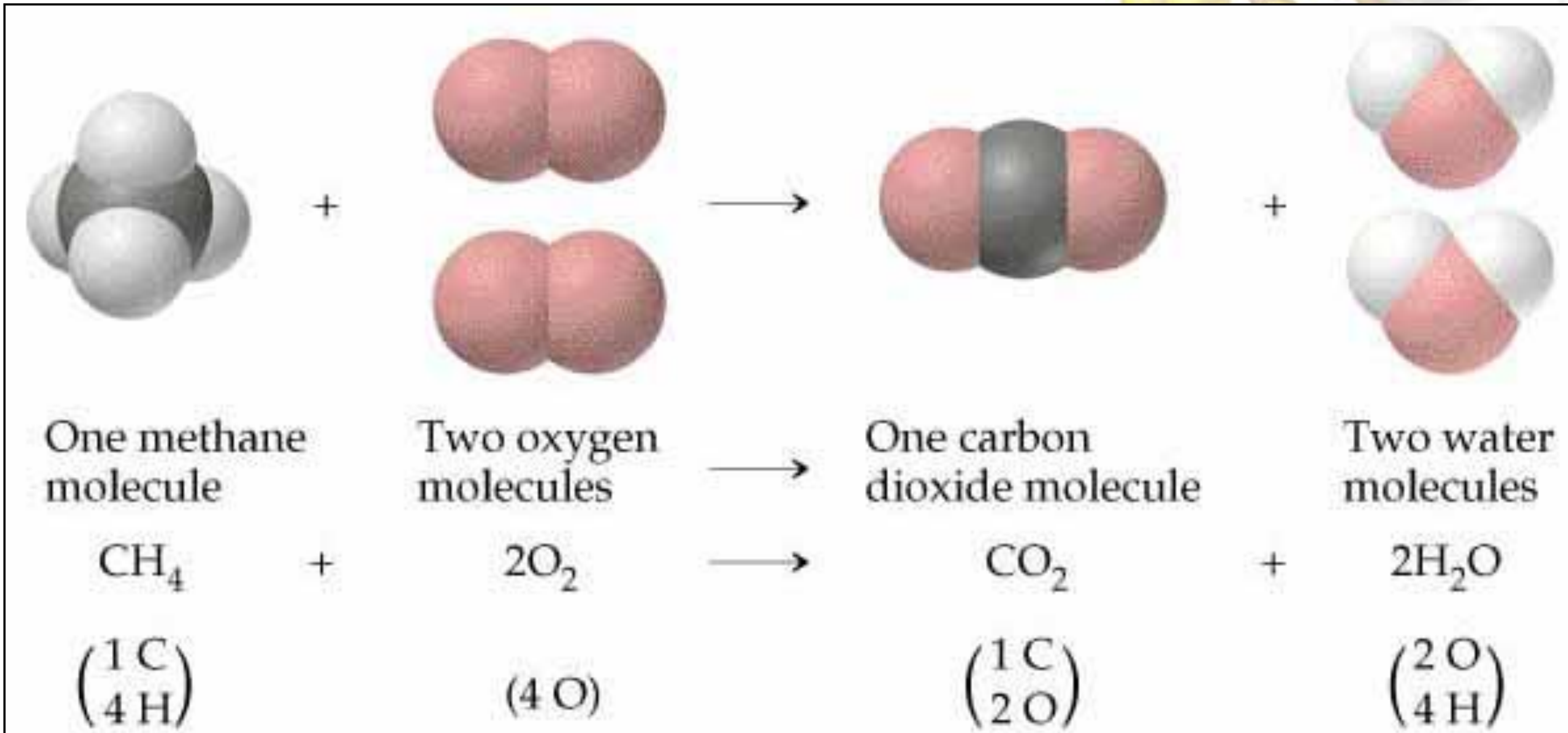
Chemical Equations

- **Stoichiometric coefficients: numbers in front of the chemical formulas; give ratio of reactants and products.**

Chemical symbol	Meaning		Composition
H_2O	One molecule of water:		Two H atoms and one O atom
$2H_2O$	Two molecules of water:		Four H atoms and two O atoms
H_2O_2	One molecule of hydrogen peroxide:		Two H atoms and two O atoms

Chemical Equations

- **Law of conservation of mass: matter cannot be lost in any chemical reactions.**



Patterns of Chemical Reactivity

Combination and Decomposition Reactions

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

There are fewer products than reactants; the Mg has combined with O_2 to form MgO.

$$2NaN_3(s) \rightarrow 2Na(s) + 3N_2(g)$$

(the reaction that occurs in an air bag)

There are more products than reactants; the sodium azide has decomposed into Na and nitrogen gas.

Patterns of Chemical Reactivity

Combination and Decomposition Reactions

Combination reactions: fewer reactants than products.

Decomposition reactions: more products than reactants.

TABLE 3.1 Combination and Decomposition Reactions

Combination Reactions	
$A + B \rightarrow C$ $C(s) + O_2(g) \rightarrow CO_2(g)$ $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$	Two reactants combine to form a single product. Many elements react with one another in this fashion to form compounds
Decomposition Reactions	
$C \rightarrow A + B$ $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$ $PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$	A single reactant breaks apart to form two or more substances. Many compounds behave in this fashion when heated.

The Mole

Molar Mass

Molar mass: sum of the molar masses of the atoms:

molar mass of N_2 = 2 x (molar mass of N).

Molar masses for elements are found on the periodic table.

Formula weights are numerically equal to the molar mass.

Quantitative Information from Balanced Equations

Balanced chemical equation gives number of molecules that react to form products.
Interpretation: ratio of number of moles of reactant required to give the ratio of number of moles of product.
These ratios are called stoichiometric ratios.

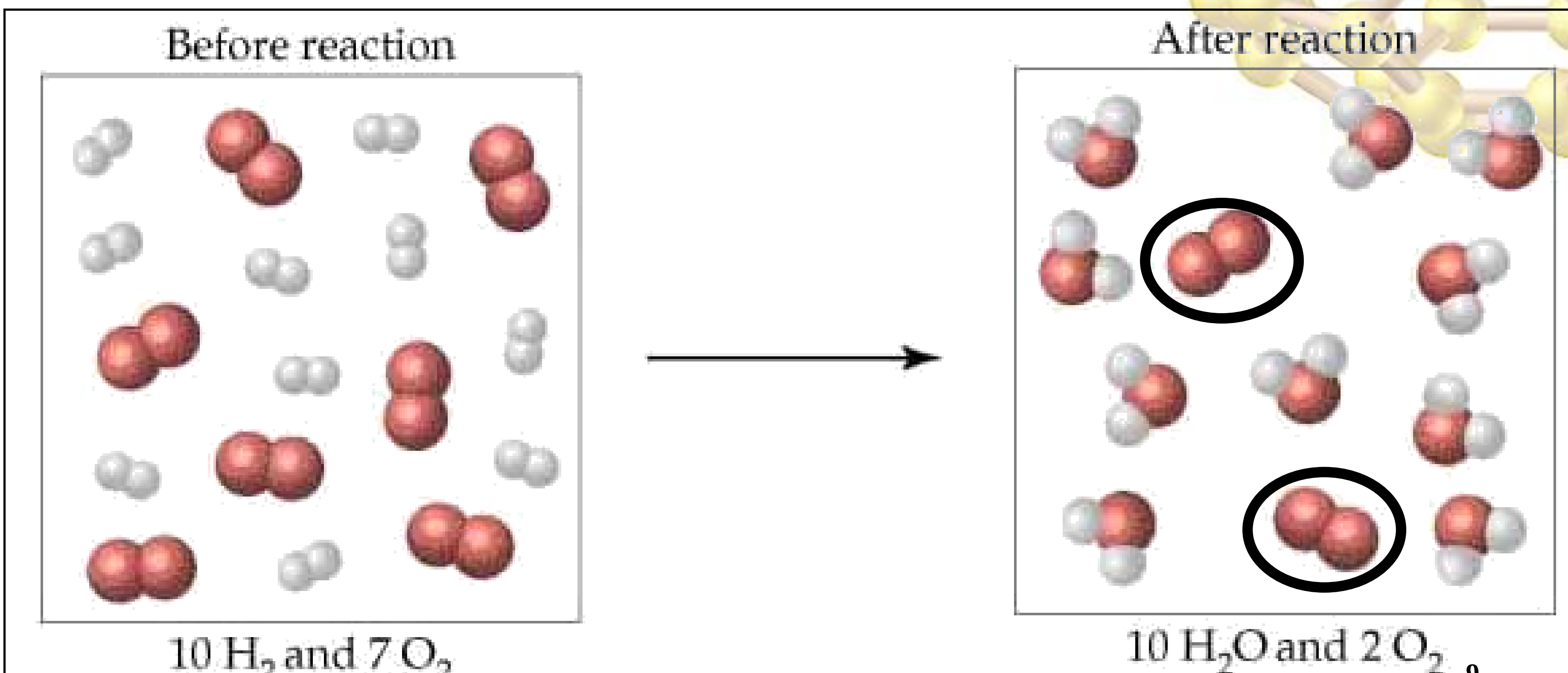
NB: Stoichiometric ratios are ideal proportions

Real ratios of reactants and products in the laboratory need to be measured (in grams and converted to moles).

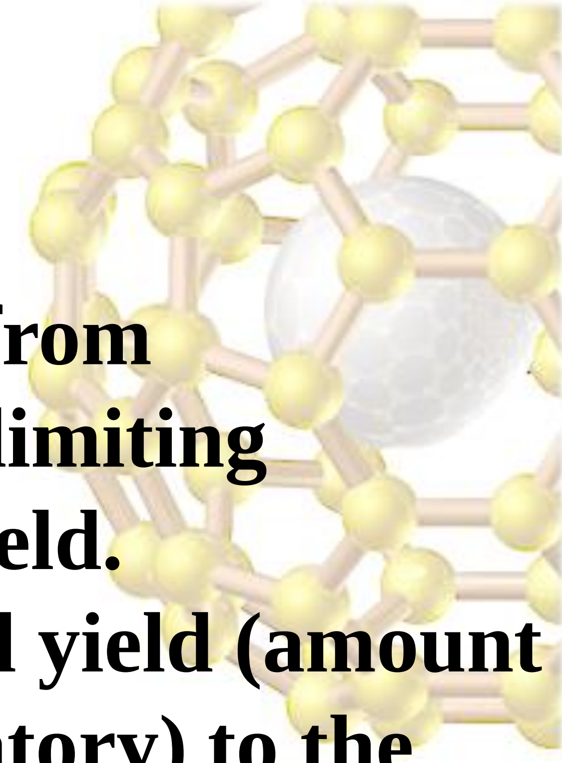
Limiting Reactants

If the reactants are not present in stoichiometric amounts, at end of reaction some reactants are still present (in excess).

Limiting Reactant: one reactant that is consumed.



Limiting Reactants



Theoretical Yields

The amount of product predicted from stoichiometry taking into account limiting reagents is called the theoretical yield.

The percent yield relates the actual yield (amount of material recovered in the laboratory) to the theoretical yield:

$$\%\text{Yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

Solution Stoichiometry and Chemical Analysis

There are two different types of units:

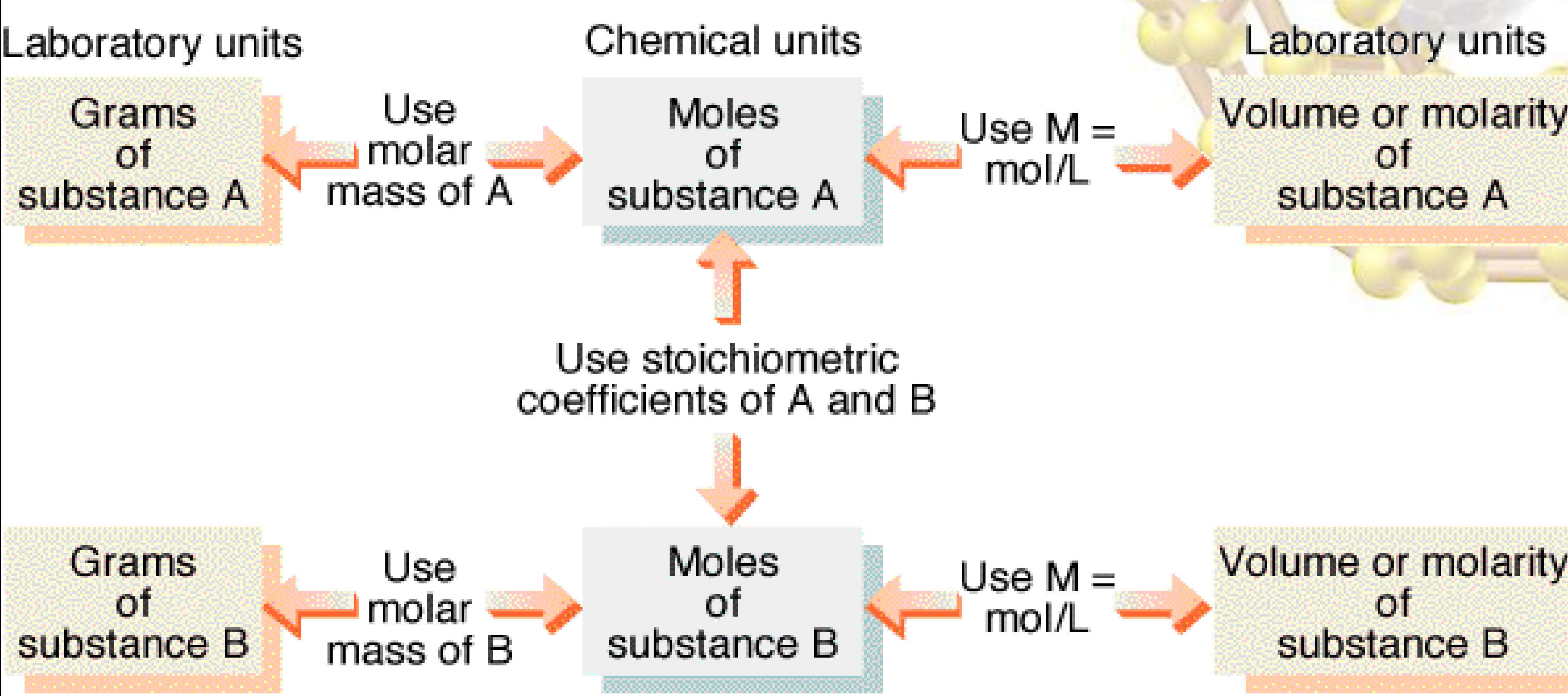
- **laboratory units (macroscopic units: measure in lab);**
- **chemical units (microscopic units: relate to moles).**

Always convert the laboratory units into chemical units first.

- **Grams are converted to moles using molar mass.**
- **Volume or molarity are converted into moles using *M* = mol/L.**

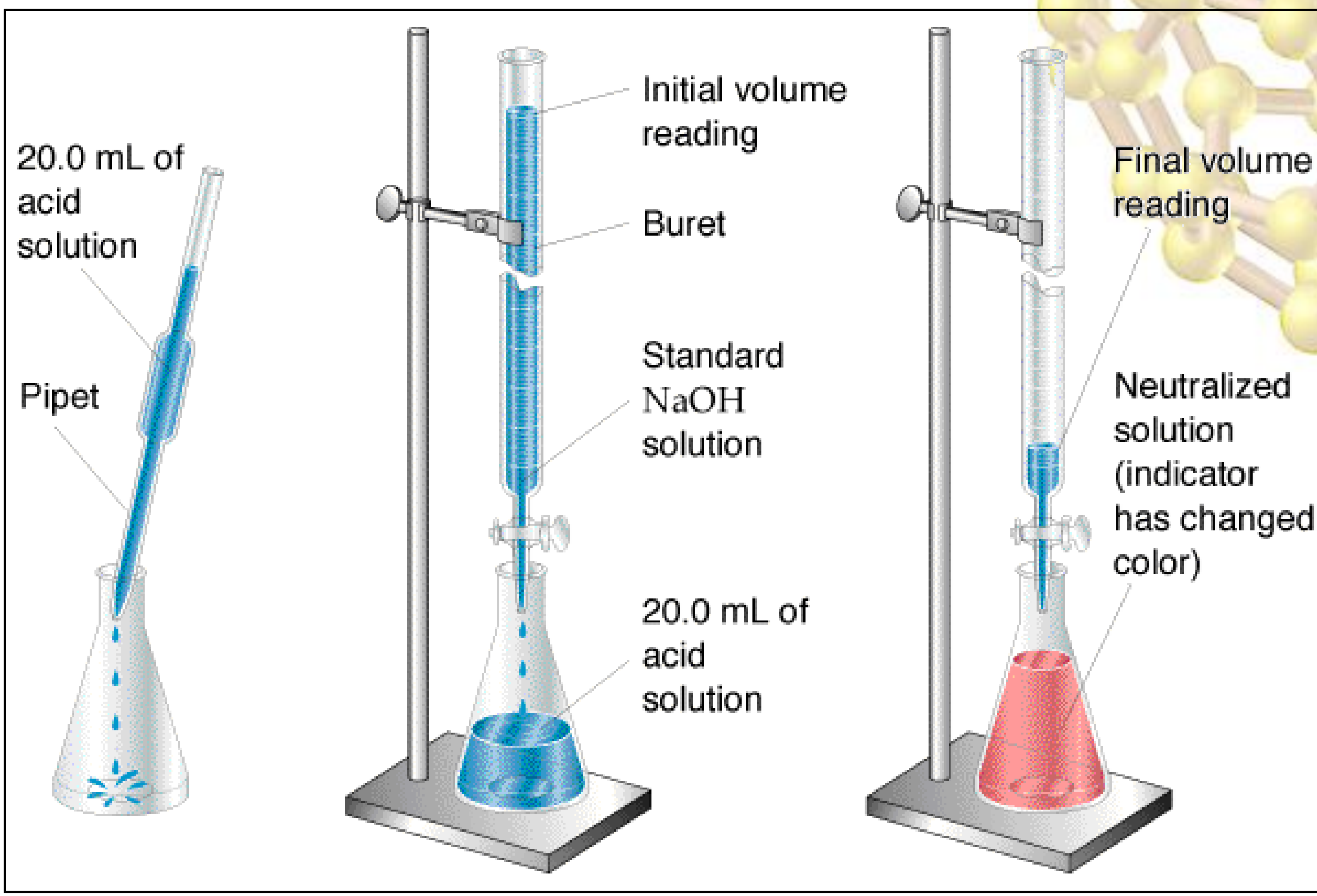
Use the stoichiometric coefficients to move between reactants and product.

Solution Stoichiometry and Chemical Analysis



Solution Stoichiometry and Chemical Analysis

Titrations



Solution Stoichiometry and Chemical Analysis

Titrations

Suppose we know the molarity of a NaOH solution and we want to find the molarity of an HCl solution.

We know: **Molarity of NaOH, Volume of HCl.**

What do we want?

Molarity of HCl.

What do we do?

Take a known volume of the HCl solution, measure the mL of NaOH required to react completely with the HCl.

Solution Stoichiometry and Chemical Analysis

Titrations

What do we get?

- **Volume of NaOH. We know molarity of the NaOH, we can calculate moles of NaOH.**

Next step?

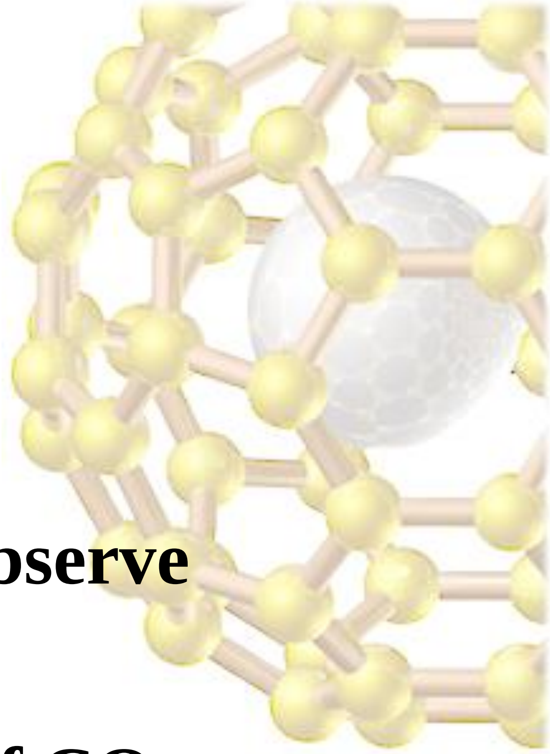
- **We also know HCl + NaOH $\rightarrow$ NaCl + H_2O. Therefore, we know moles of HCl.**

Can we finish?

- **Knowing mol(HCl) and volume of HCl (20.0 mL above), we can calculate the molarity.**

Exp. 2; Qualitative analysis

Cl^-, CO_3^{2-}, PO_4^{3-}, SO_4^{2-}



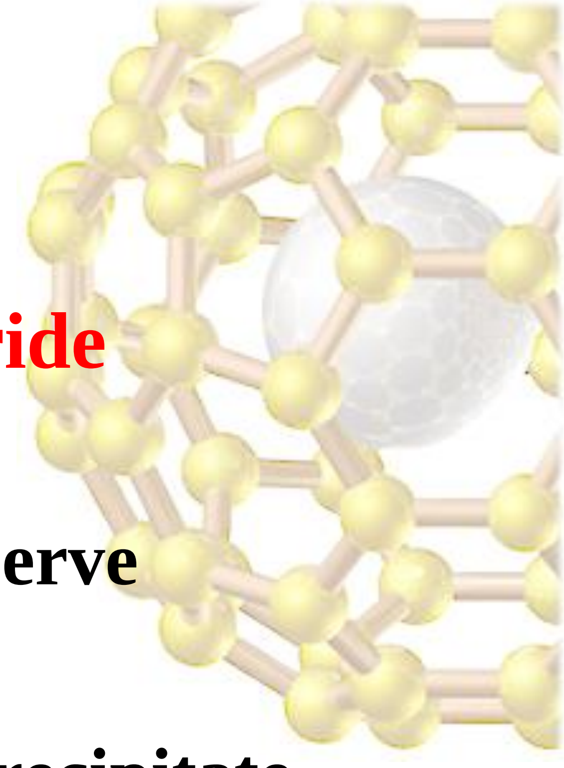
Reaction with dil. HCl

- **Adding few drops of dil. HCl and observe**

- **HCl + CO_3^{-2} → CO_2(g) Evolution of CO_2 gas**
- **HCl + Cl^- → No reaction**
- **HCl + PO_4^{3-} → No reaction**
- **HCl + SO_4^{2-} → No reaction**

Reaction with silver nitrate

- **Adding few drops of $AgNO_3$ and observe**
- **$Ag^+ + Cl^- \rightarrow AgCl(s)$ White Precipitate; turn violet in light**
- **$Ag^+ + PO_4^{3-} \rightarrow Ag_3PO_4(s)$ Yellow precipitate**
- **$Ag^+ + SO_4^{2-} \rightarrow Ag_2SO_4(s)$ Turbid-White precipitate**



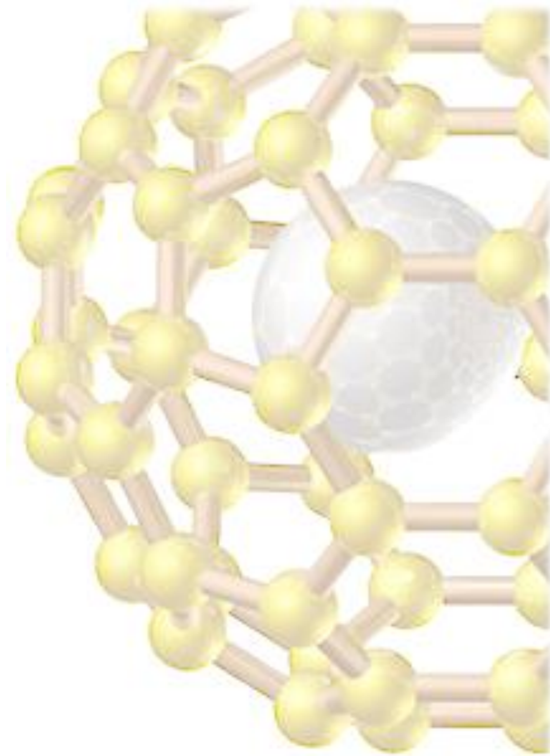
Reaction with Barium chloride

- **Adding few drops of $BaCl_2$ and observe**
- **$BaCl_2 + SO_4^{2-} \rightarrow BaSO_4(s)$ White precipitate**

Identification of Unknown ANION

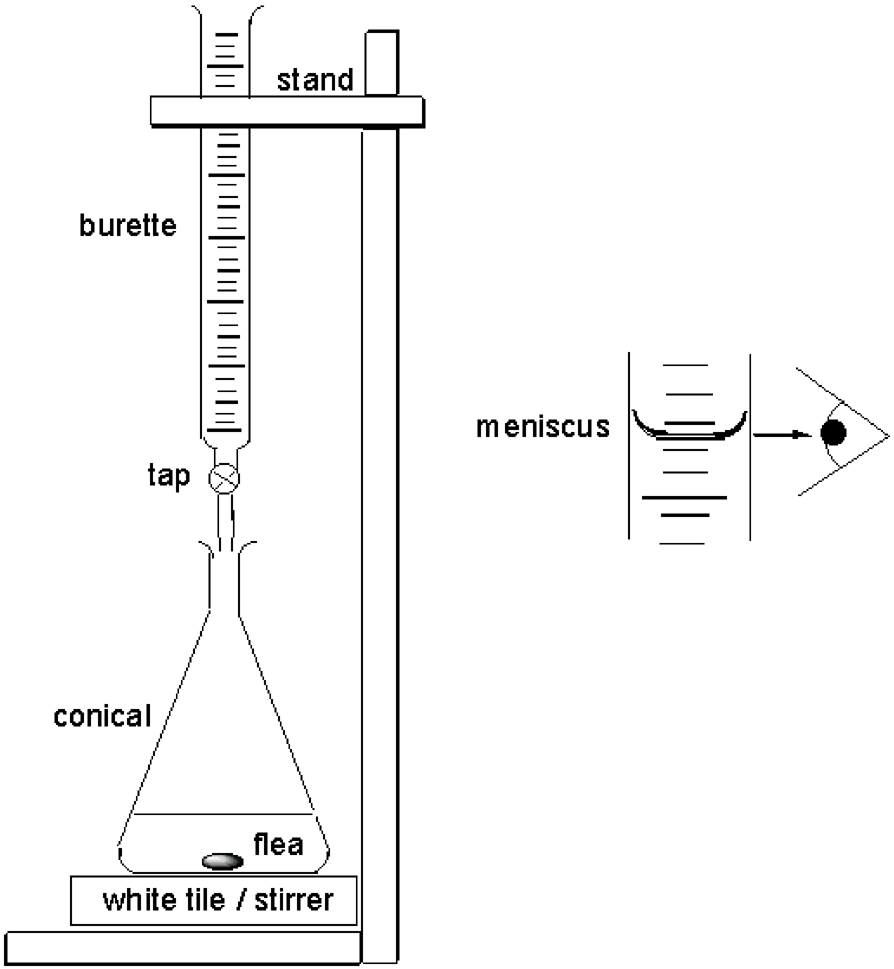
Anion/ Reagent	Dil HCl	$AgNO_3$	$BaCl_2$
CO_3^{2-}	effervescence and Evolution of CO_2		
Cl^-	-ve No reaction	White ppt turned violet in light	
PO_4^{3-}	-ve No reaction	Yellow ppt	
SO_4^{2-}	-ve No reaction	Turbid	White ppt

Exp. 3; ACID BASE TITRATION



INDICATORS

Titration



- **Method of analysis that will allow you to determine the precise endpoint of a reaction and therefore the precise quantity of reactant in the titration flask.**
- **The burette is used to deliver the second reactant to the flask.**
- **Indicator or pH Meter is used to detect the endpoint of the reaction.**

Doing a Titration

<u>**Begin by preparing your burette by**</u>

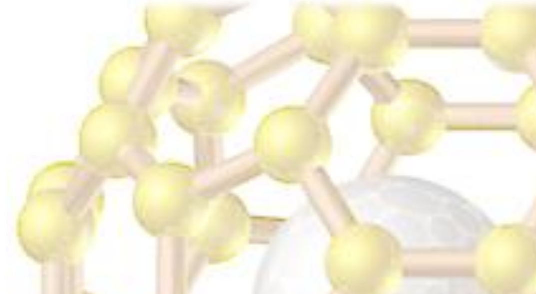
1. **washing with water**
2. **rinsing with tap and then distilled water**
3. **and rinsing with the titrant solution**

- ***You should check for air bubbles and leaks, before proceeding with the titration.***
- ***Be sure the tip of the burette is filled.***

- **Never dispense so that liquid is below the last calibration that you can read.**

How you read the burette

- **Take an initial volume reading and record it.**
- **Before beginning a titration, you should always calculate the expected endpoint volume.**



- ***<u>Preparation of TITRAND</u>***

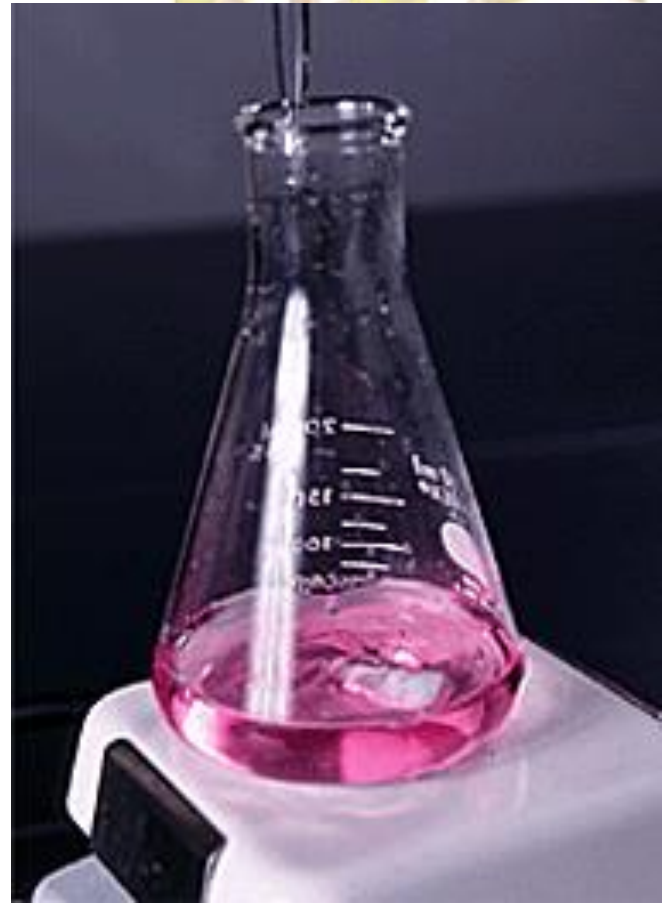
- **Prepare the solution to be analyzed by placing it in a clean Erlenmeyer flask or beaker.**
- **If your sample is a solid, make sure it is completely dissolved.**
- **Add indicator.**

- **Use the burrete to deliver a stream of titrant to within a couple of mL of your expected endpoint.**

- **You will see the indicator change color when the titrant hits the solution in the flask, but the color change disappears upon stirring.**

- ***Approach the endpoint more slowly and watch the color of your flask carefully.***
- **Use a wash bottle to rinse the sides of the flask and the tip of the buret, to be sure all titrant is mixed in the flask.**
- **Make sure you know what the endpoint should look like.**
- **For phenolphthalein, the endpoint is the first permanent pale pink.**
- **The pale pink fades in 10 to 20 minutes.**

- **If you think you *might* have reached the endpoint, you can record the volume reading and add another partial drop.**
- **Sometimes it is easier to tell when you have gone past the endpoint.**
- **If the flask looks like this, you have gone too far!**



Acid-Base Indicators

- **An acid-base indicator is a weak acid or a weak base.**
- **The undissociated form of the indicator is a different color than the original form of the indicator.**
- **An Indicator does not change color from pure acid to pure alkaline at specific hydrogen ion concentration, but rather, color change occurs over a range of hydrogen ion concentrations.**
- **This range is termed the *color change interval*. It is expressed as a pH range.**

Litmus Papers

- **Litmus is a weak acid. It has a seriously complicated molecule which we will simplify to HLit. The "H" is the proton which can be given away to something else. The "Lit" is the rest of the weak acid molecule.**
- **There will be an equilibrium established when this acid dissolves in water. Taking the simplified version of this equilibrium:**

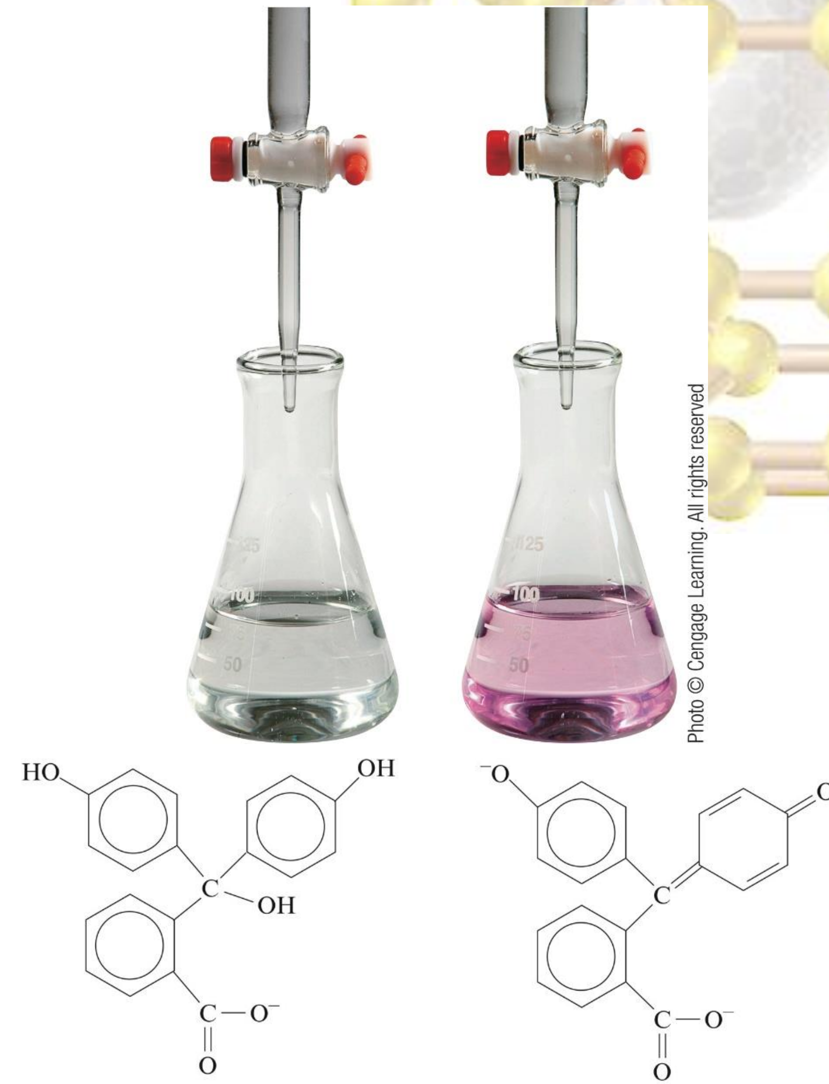
$$HLit_{(aq)} \rightleftharpoons H^+_{(aq)} + Lit^-_{(aq)}$$

- **The un-ionized litmus is red, whereas the ion is blue.**

End Point of an Indicator

- **should not be confused with the equivalence point of a titration**
- **indicator changes color at its endpoint**
- **equivalence point is the stoichiometric point where neutralization takes place**
- **ideally, the end point of the indicator and the stoichiometric equivalence point should coincide**

The Acid and Base Forms of the Indicator Phenolphthalein

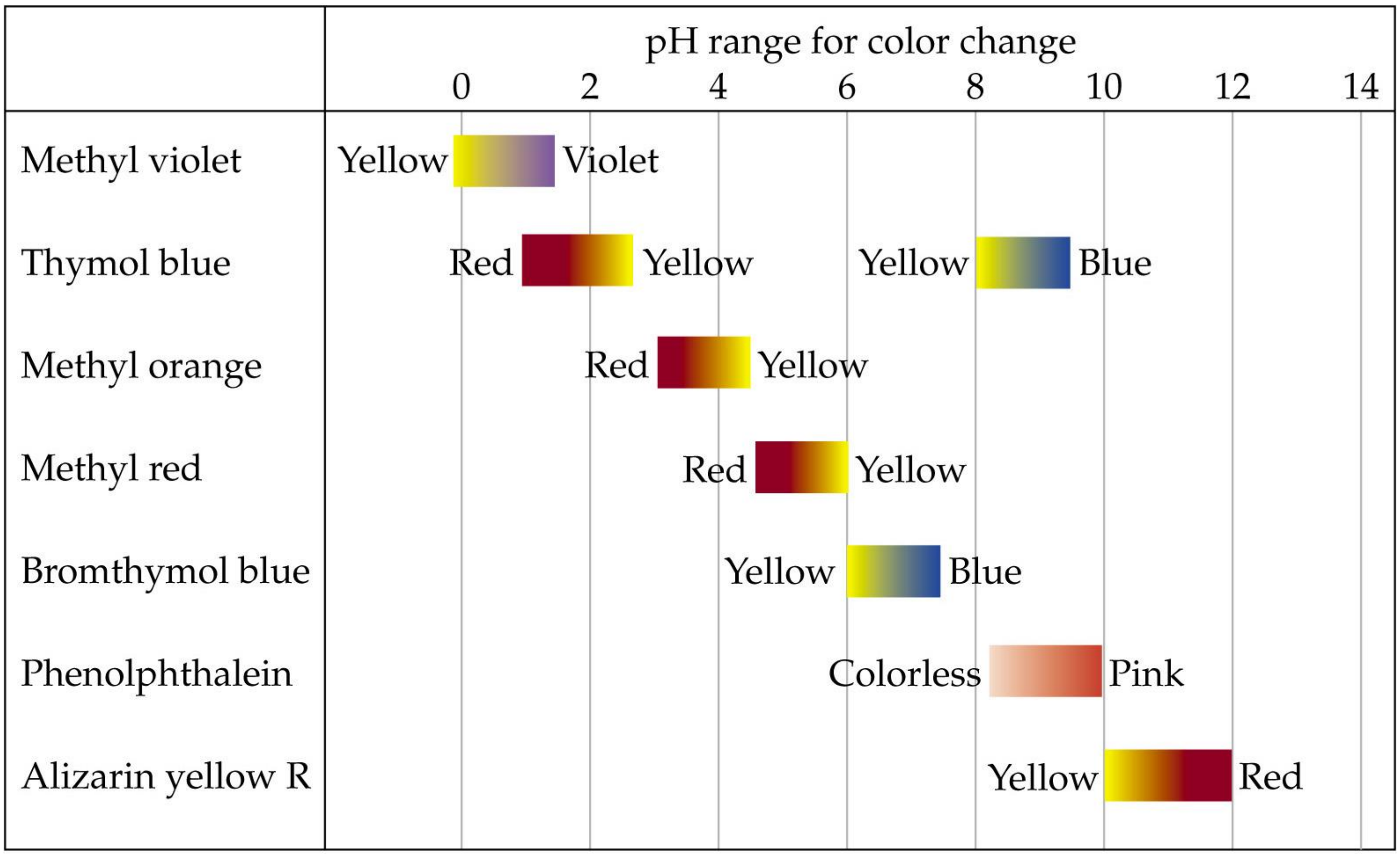


Colorless acid form, HIn

Pink base form, In^{-}

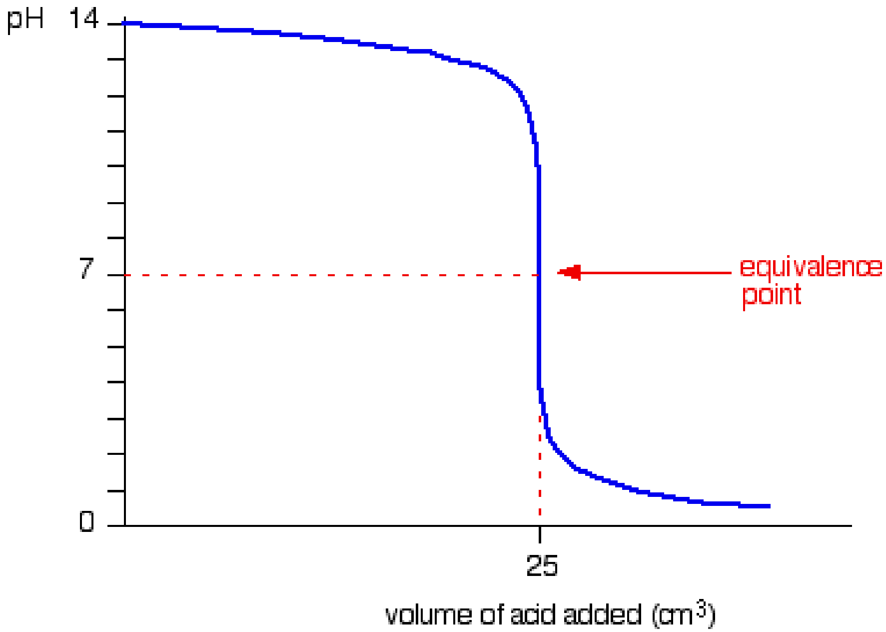
The Methyl Orange Indicator is Yellow in Basic Solution and Red in Acidic Solution

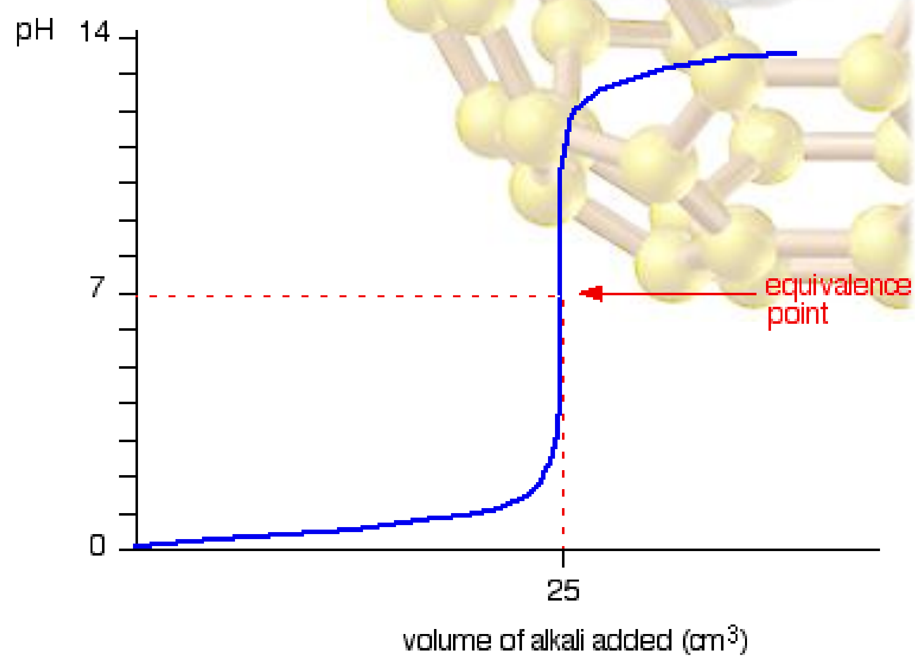
Some common indicators



Titration curves for strong acid vs. strong base

$$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)}$$



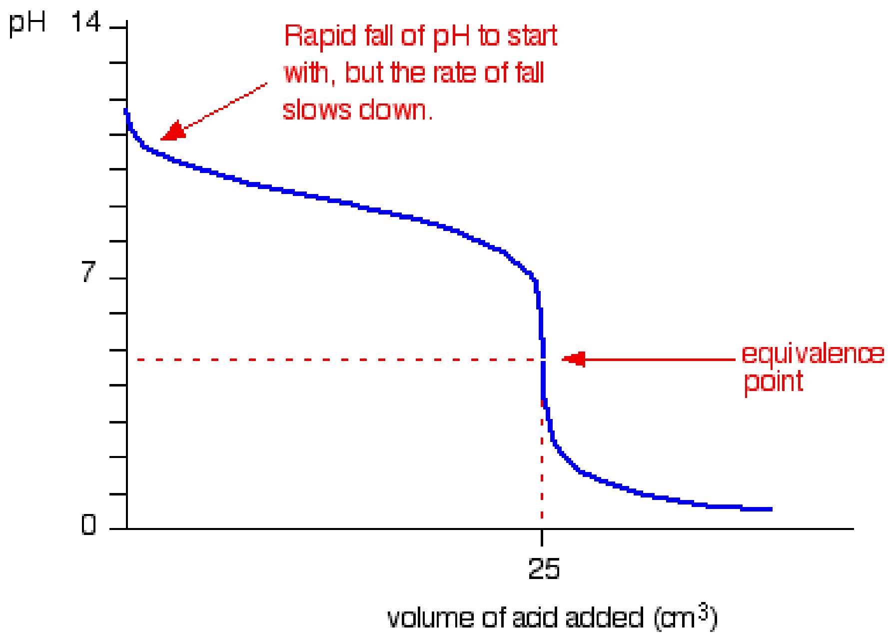


Running acid into the base

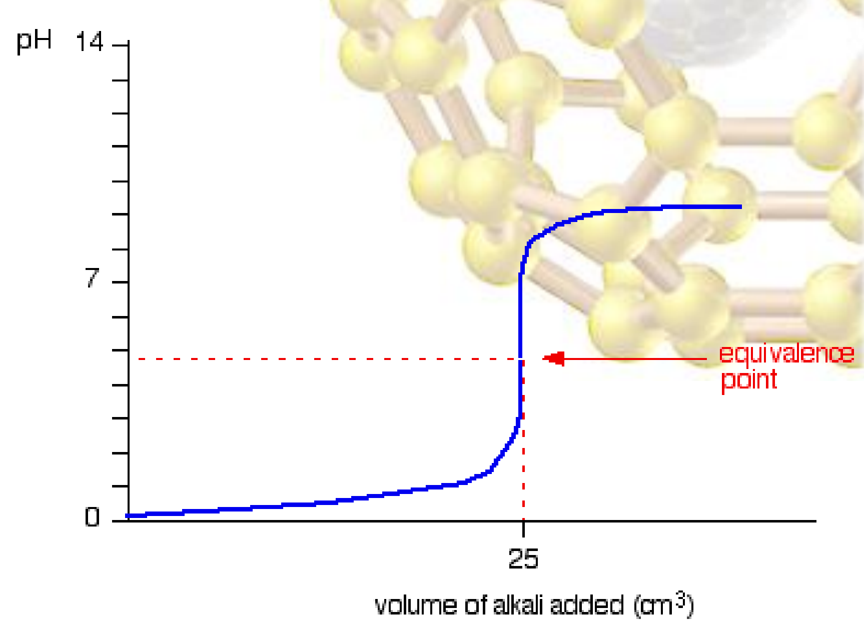
Running base into the acid

Titration curves for strong acid vs. weak base

$$NH_{3(aq)} + HCl_{(aq)} \longrightarrow NH_4Cl_{(aq)}$$



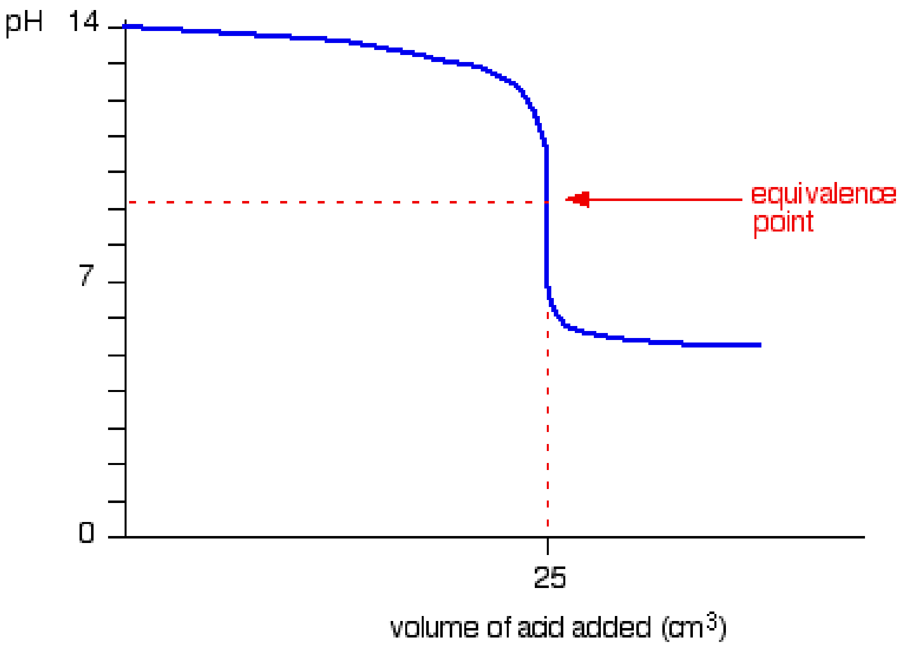
Running acid into the base

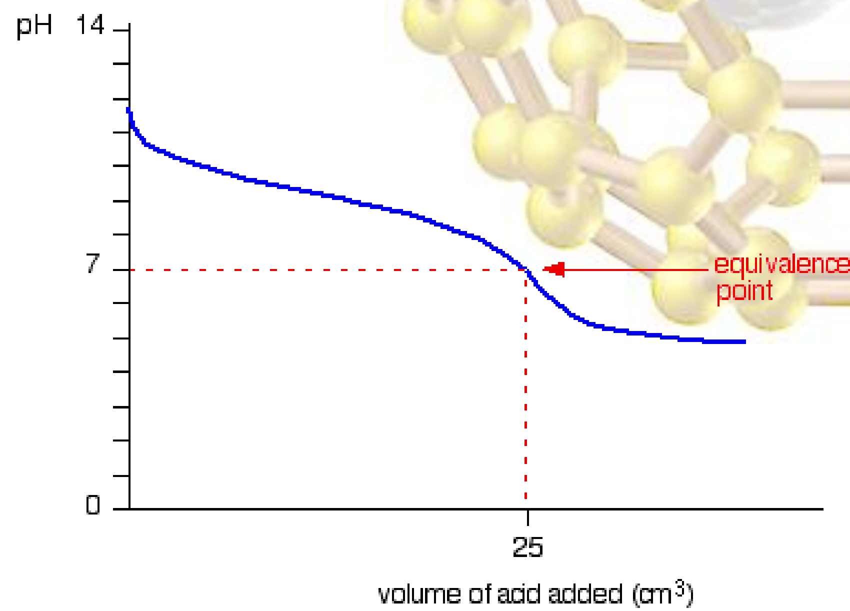


Running base into the acid

Titration curves for weak acid vs. strong base

$$CH_3COOH_{(aq)} + NaOH_{(aq)} \longrightarrow CH_3COONa_{(aq)} + H_2O_{(l)}$$



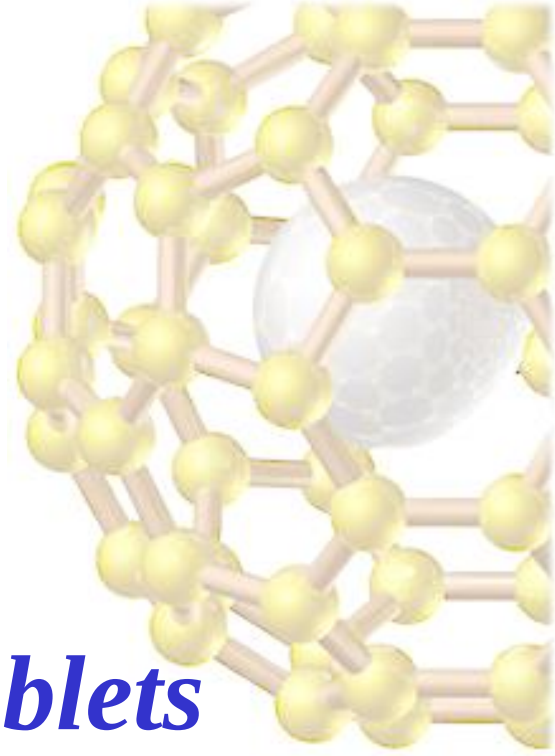


Running acid into the base

Running base into the acid

Acidic, Basic and Neutral Salts

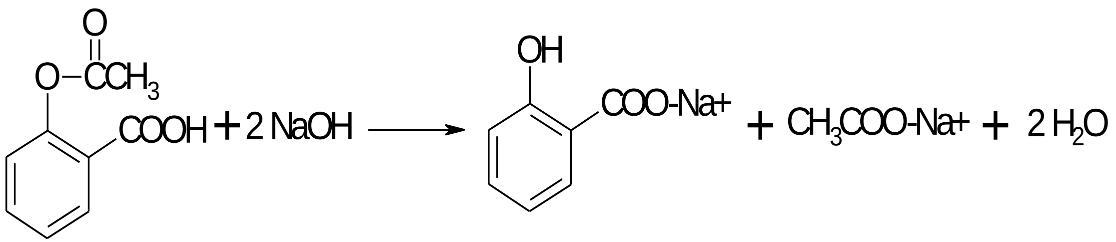
- **A salt is formed between the reaction of an acid and a base.**
- **Usually, a neutral salt is formed when a strong acid and a strong base is neutralized in the reaction: $H^+ + OH^- \Rightarrow H_2O$**
- **When weak acids and bases react, the relative strength of the conjugated acid-base pair in the salt determines the pH of its solutions.**
- **The salt, or its solution, so formed can be acidic, neutral or basic.**
- **A salt formed between a strong acid and a weak base is an acid salt, for example NH_4Cl.**
- **A salt formed between a weak acid and a strong base is a basic salt, for example CH_3COONa.**
- **These salts are acidic or basic due to their acidic or basic ions as shown in the tables here.**



The analysis of aspirin tablets

- **<u>Objective</u>**
- **To determine the amount of acetylsalicylic acid ($C_9H_8O_4$) in aspirin tablet and compare the result with the manufacturer's specification.**
- **<u>Introduction</u>**
- **In all pharmaceutical preparations, the manufacturer is required by law to state on the packaging the maximum amount of each active ingredient present. In many preparations the active ingredient will only form a small percentage of the pill or tablet as a whole. For example, in the case of many tablets, the tablet would simply disintegrate into a powder unless additives were put in to assist in making the ingredients cohere into tablet form. In this experiment students will have the opportunity to carry out a consumer survey on the aspirin’s content of a number of commercial preparations, to see whether the manufacturer's claims are justified.**

- **Theory**
- **Acetylsalicylic acid (aspirin) can be readily hydrolysed by sodium hydroxide into the sodium salts of two weak acids, acetic acid (ethanoic acid) and salicylic acid. In this experiment, the hydrolysis is effected using an excess of sodium hydroxide solution, the excess alkali will be later found by titration with standard sulphuric acid. The equation for the hydrolysis reaction is:**



$CH_3COOC_6H_4COOH + 2\ NaOH \rightarrow CH_3COO^-Na^+ + HOC_6H_4COO^-Na^+ + H_2O$

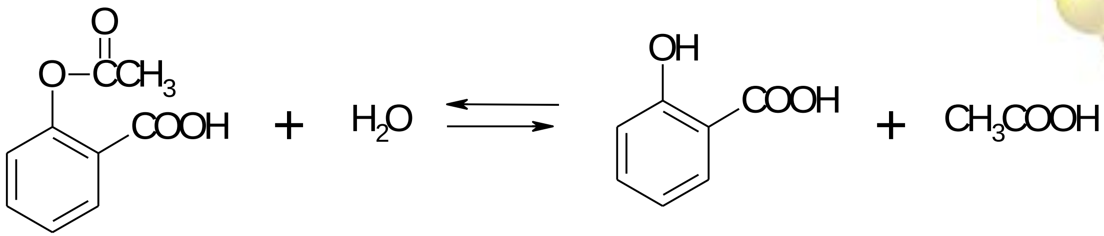
- **<u>Procedure</u>**
- **Weigh accurately a definite number of aspirin tablets (two, three tablets or not weighing more than 1.5g) into a 250cm^3 conical flask. Initiate the hydrolysis of aspirin by adding 25cm^3 of 1.0M sodium hydroxide by pipette, diluting with approximately the same volume of distilled water. Warm the flask over a tripod and gauze for ten minutes to complete the hydrolysis.**

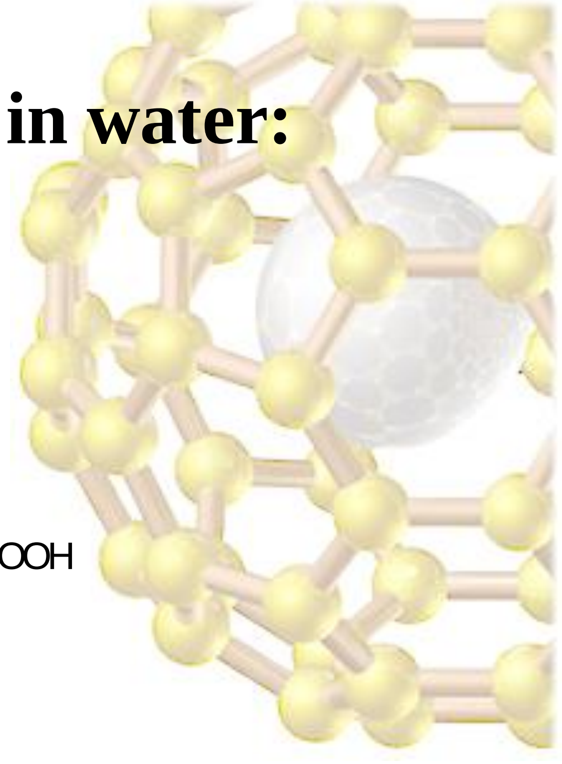
- **Cool the reaction mixture and transfer with washings to a 250cm^3 volumetric flask. Dilute to the mark with distilled water and then ensure that the contents of the flask are well mixed by repeated shakings.**
-
- **Titrate 25cm^3 portions of the diluted reaction mixture with the standard 0.05M sulphuric acid provided, using phenol red indicator. This indicator (with pH range 6.8 to 8.4) is most suitable for this titration. Repeat the titration twice.**

- **Calculation**
- **Calculate the mass of acetylsalicylic acid in each tablet and compare your result with the manufacturer's specification.**
- **Assume that the manufacturer’s specification is correct, calculate the percentage error of your experiment.**

- **<u>Questions</u>**
-
- **State the function of the sodium hydroxide solution used in the hydrolysis* of aspirin tablets. Does the initial amount (concentration and volume) of this alkaline solution be certain?**
- **State the reason why hydrolysis of aspirin is needed instead of direct titration of aspirin against NaOH.**
- **Explain why the hydrolysis product solution must be cooled down to room temperature before it is transferred to the volumetric flask.**
- **Explain the choices of indicators used in the titration process of this experiment.**

- * **Consider the hydrolysis of aspirin in water:**





- **Title:** **The analysis of aspirin tablets**

- **Objective**
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- **Calculation**

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