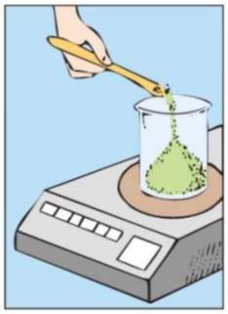
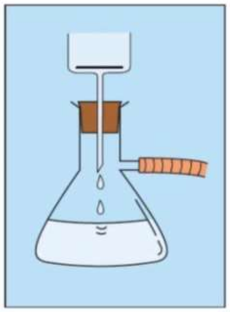
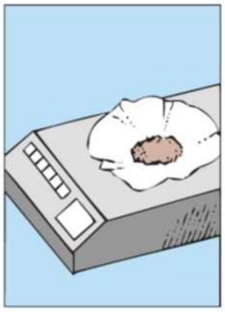
Quick Guide to Experimental Procedures

Gravimetric Analysis:



Weighing the sample to be analyzed.	**Dissolving this sample in water.**	**Adding a suitable chemical to form a precipitate.**	**Filtering to collect the precipitate**	**Repeated drying and weighing until a constant mass of precipitate is obtained.**

Common Mistakes:

- **Precipitate is not dry when you take the final mass.**
    - Results in the appearance of more precipitate than was actually produced because some mass is water.
    - Percent yield would be higher than it should be.

Common Applications:

- Mixtures of solids—determining the amount of a particular ion in a solution

Important to Remember:

- All sodium, nitrate, ammonium, and potassium compounds are soluble. Net ionic equations would not include these ions.

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Making a solution:

The AP Readers would prefer that you use a pipet at this point to make sure you hit the line perfectly.

(a) An amount of solute is weighed out on an analytical balance and then

(b) A portion of the solvent is added to the volumetric flask.

(c) The mixture is swirled until *all* of the solute is dissolved.

(d) Additional solvent is added up to the mark on the volumetric flask.

Common Mistakes:

- **Overfilling the volumetric flask**
    - Results in a dilute solution
- **Not using distilled water.**
    - Other ions could affect the experiment for which the solution is used
- **Not using a volumetric flask (beaker or Erlenmeyer instead)**
    - Loss of precision in concentration of prepared solution

Common Applications:

- Making solutions to dissolve substances for analysis, particularly in titrations.

Important to Remember:

- Molarity = moles solute/L of solution

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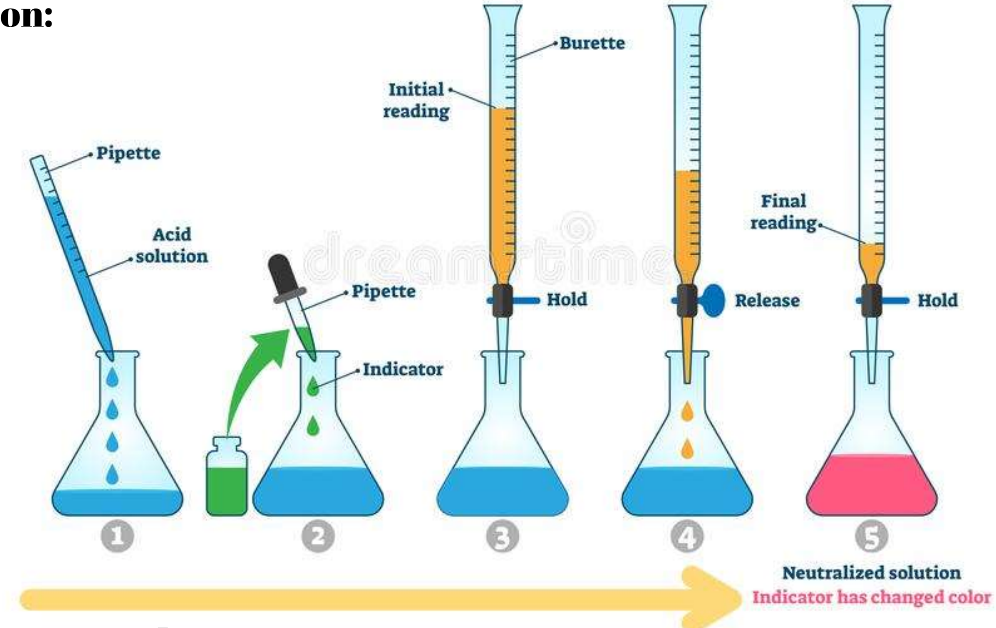
Titration:



Common Mistakes:

- **Overshooting the titration (too dark of a color at the end)**
  - Results in the concentration of the unknown solution in the flask appearing to be higher than it actually is, since too much titrant has been added.
- **Not using indicator.**
  - No perceivable endpoint.
- **Using incorrect indicator.**
  - pH at the equivalence point should be approximately equal to the pKa of the indicator.

Common Applications:

- Solving for the concentration of an unknown substance (analyte).
- Acid/Base, Redox

Important to Remember:

- Molarity = moles solute/L of solution
- **Analyte:** substance in flask
- **Titrant:** substance in buret
- **Standard solution:** solution of known concentration, usually goes into the buret.
- **$M_1V_1 = M_2V_2$** is helpful for solving for the concentration of the analyte solution at the equivalence point.
- **Endpoint:** point in titration where flask solution changes color
- **Equivalence point:** point in the titration where the moles of acid are equal to the moles of base

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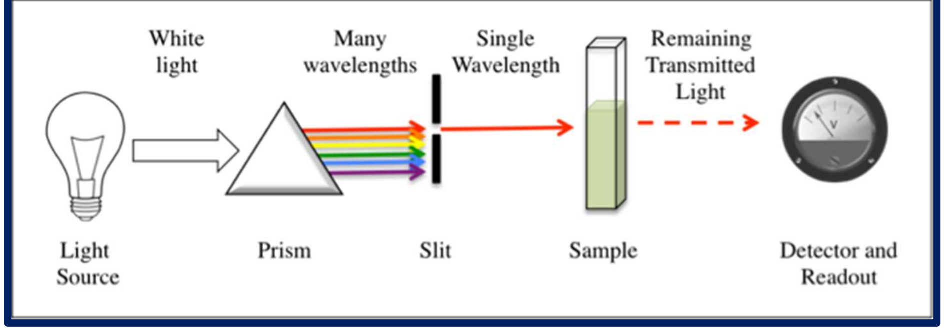
Analyzing Concentration of Solutions Using Beer's Law):

Diagram 1: Spectrophotometer



Step 1: Pick the wavelength for the solution where absorbance is highest (for solute).

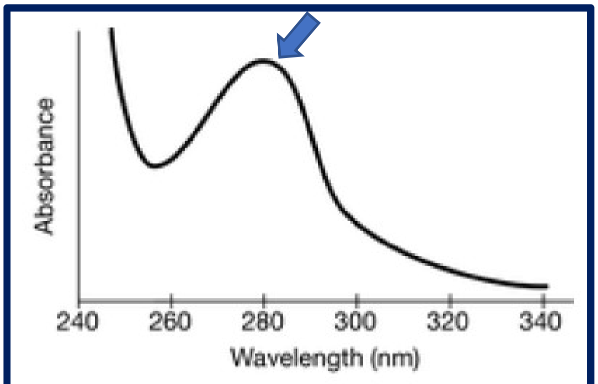
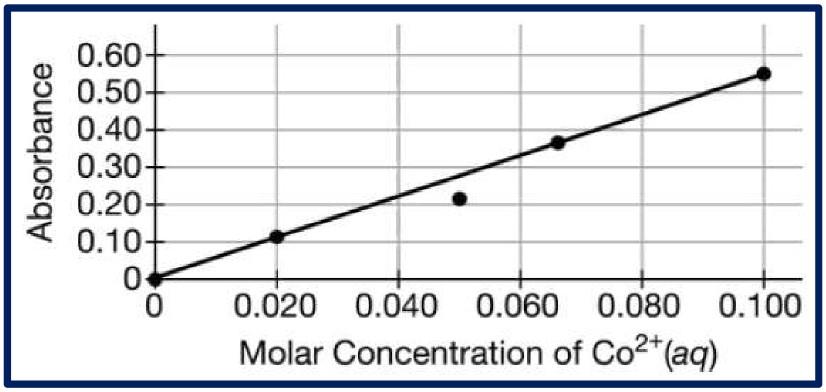


Diagram 2

Step 2: Measure absorbance for different concentrations at that wavelength. Graph the results.



$$A = \varepsilon bc$$

Absorbance = (molar absorptivity)(cuvette pathway length)(concentration)

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Common Mistakes:

- **Absorbance is lower than it should be (point falls below the line)**
  - Cuvette was cleaned with distilled water and then immediately filled with solution, creating a more dilute solution
  - Too little solute in the prepared solution
- **Absorbance is higher than it should be (point falls above the line)**
  - Cuvette is dirty with fingerprints/dust, etc.
  - Too much solute in the prepared solution
  - Contamination with a more concentrated solution
  - Used a cuvette with a longer path for one data point
- **Did not use the correct wavelength of maximum absorbance for the solute.**
  - Absorbances could be too low especially for dilute solutions
- **Overfilled the cuvette**
  - Should not have an impact on data
- **Picked a wavelength where it is high absorbance for the solvent**
  - Won't be able to distinguish absorbance due to solvent vs. solute

Common Applications:

- Determining the concentration of a solution of unknown concentration using solutions of known concentration
- Kinetics reactions (like bleach + blue food dye)

Important to Remember:

- Before using, you need to calibrate the spectrophotometer with a blank of just solvent (in order to account for any absorbance due to solvent and cuvette itself)
- Molarity = moles solute/L of solution
- Absorbance is the amount of light the solution absorbs at a specific wavelength
- Molar absorptivity (1/M*cm) describes how intensely a sample absorbs light at a specific wavelength (constant unique to the substance at a specific wavelength)
- Path length of sample is the length of the cuvette where the light will travel (cm)
- Concentration is molarity

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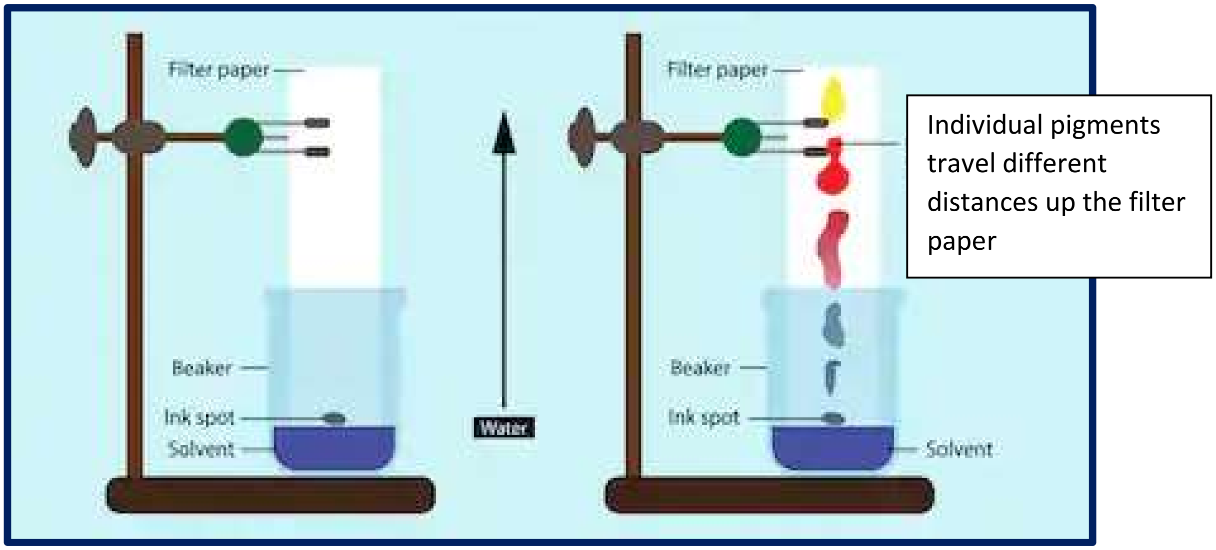
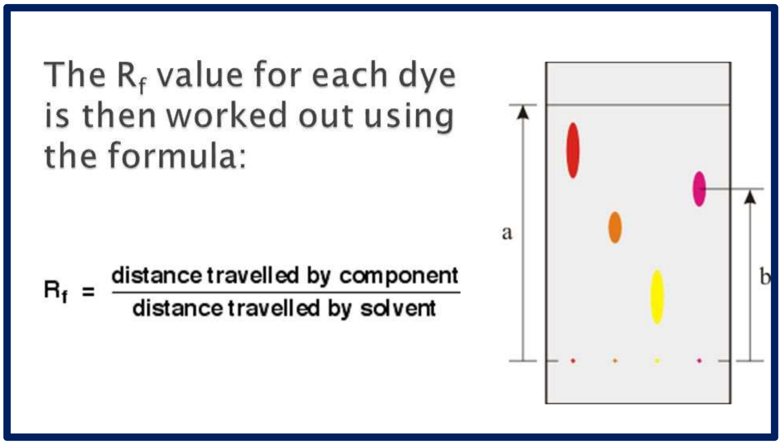
Chromatography



Individual pigments travel different distances up the filter paper



The R_f value for each dye is then worked out using the formula:

$$R_f = \frac{\text{distance travelled by component}}{\text{distance travelled by solvent}}$$

Common Mistakes:

- **Solvent reaches the top of the paper strip.**
    - Rf values cannot be calculated as we do not know how far the solvent would have traveled had their been more paper.
- **No major difference in polarity between paper and solvent**
    - Substances cannot be adequately separated
- **No major differences in polarity of components of mixture**
    - Substances cannot be adequately separated

Common Applications:

- Determining the components of a mixture

Important to Remember:

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- Paper is usually relatively nonpolar in comparison to the solvent.
- The substance that travels further up the paper is more attracted to the solvent.
- The substance that travels the least is most attracted to the paper.
- If multiple trials are run, compare Rf values, not relative heights.
- Polar substances tend to lack symmetry, have polar bonds, and have lone pairs on the central atom. They are most soluble in other polar substances.
- Nonpolar substances tend to be symmetrical, have identical bonds, and have no lone pairs on the central atom. They are most soluble in other nonpolar substances.

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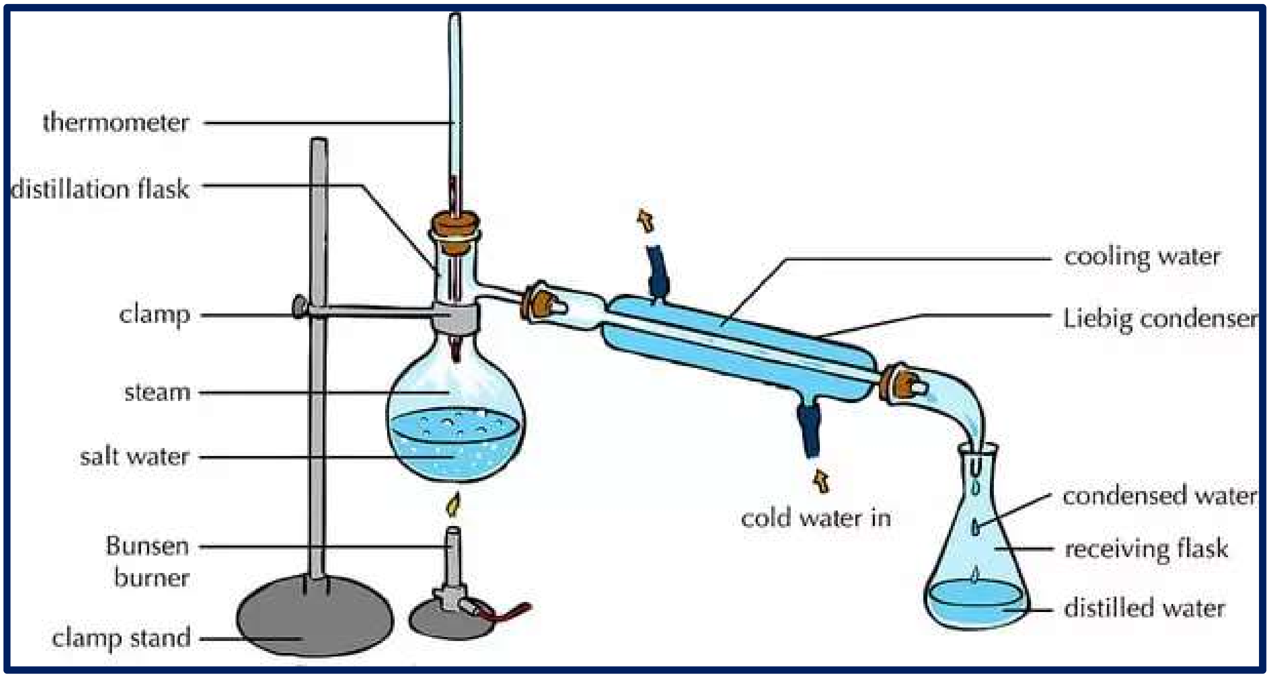
Fractional Distillation



Common Applications:

- Separating components in a solution/mixture based on differences in boiling point

Important to Remember:

- The substance with the lower boiling point has a greater vapor pressure and weaker intermolecular forces
- The substance with the higher boiling point has a lower vapor pressure and stronger intermolecular forces
- The temperature of the solution will remain constant while a component is boiling off

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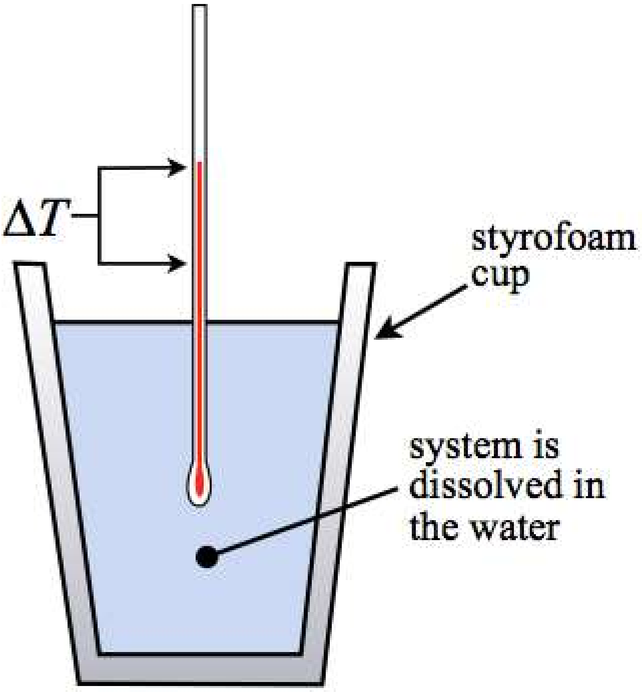
Coffee Cup Calorimetry



Common Mistakes:

- The final temperature is the highest (for exothermic) or lowest (for endothermic) temperature recorded during the reaction/process

Applications:

- Solving for the specific heat of a metal or the heat of reaction

Important to Remember:

- Endothermic processes have a drop in temperature.
- Exothermic processes have an increase in temperature.
- The water is not part of the system. It is part of the surroundings.
- $q = mC\Delta T$
    - q = heat in Joules or calories
    - m = mass of entire solution OR object, grams or kilograms
    - C = specific heat capacity, J/g°C (or a variation of the above)
    - $\Delta T = T_{final} - T_{initial}$