ORGANIC I TECHNIQUES

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Organic I Techniques

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This text was compiled on 08/22/2025

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CHAPTER OVERVIEW

1: Safety

1.1: Pictograms and SDS

Communication of Chemical Hazards

The Hazard Communication Regulation of the Occupational Safety and Health Administration (OSHA) requires that anyone who might come in contact with chemicals be informed of potential hazards. Communication of chemical hazards is through labels and safety data sheets as described below.

The Fire Diamond

A quick visual representation of hazards is provided by the Health, Flammability, Instability (HFI) or “Fire” diamond developed by the National Fire Protection Association (NFPA). The diamond consists of four quadrants with numbers or special symbols representing the degree of certain hazards. The top three quadrants contain NFPA codes (numbers from 0 to 4) indicating the degree of a particular risk.

4 – Extreme Risk

3 – Serious Risk

2 – Moderate Risk

1 – Slight Risk

0 – Minimal Risk

Quadrant	Type of Hazard	Color	Example
left	health	blue	3 = severe health hazard
top	flammability	red	2 = moderately flammable
right	instability[1]	yellow	1 = Slight reactivity
bottom	special	white	W = use no water

Safety Data Sheet (SDS)

Chemical manufacturers are required by law to provide a Safety Data Sheet (or SDS) for each chemical. Prior to 2015, this document was called a Material Safety Data Sheet or MSDS. Although your instructor or laboratory manual will always inform you of any hazards associated with the chemicals you will be using, you may view the SDS for any chemical you will use in the laboratory. The format of a Safety Data Sheet is standardized to contain 16 sections, each providing a specific type of information.

Label Element	Content
Product Identifier	Chemical name of the hazardous chemical
Supplier Identification	Contact information of manufacturer or distributor
Signal Word	Severity level of hazard is indicated using two words **DANGER:** more severe hazard **WARNING:** less severe
Hazard statements	Nature and degree of hazard
Precautionary statements	Recommended measures to minimize adverse effects
Supplemental Information	
Pictograms	Nine different pictograms are used to designate distinct physical, health and environmental hazards.

Hazard Pictograms

OSHA required hazard pictograms on labels consist of a black hazard symbol on a white background framed within a red border in the shape of a square set on a point. You should be familiar with these symbols.

Health Hazard	Flame	Exclamation Mark

		
Carcinogen Mutagenicity Reproductive Toxicity Respiratory Sensitizer Target Organ Toxicity Aspiration Toxicity	Flammables Pyrophorics Self-Heating Emits Flammable Gas Self-Reactives Organic Peroxides	Irritant (skin and eye) Skin Sensitizer Acute Toxicity Narcotic Effects Respiratory Tract Irritant Hazardous to Ozone Layer (Non-Mandatory)
Corrosion	**Exploding Bomb**	**Gas Cylinder**
		
Skin Corrosion/Burns Eye Damage Corrosive to Metals	Explosives Self-Reactives Organic Peroxides	Gases Under Pressure
Flame Over Circle	**Skull and Crossbones**	**Environment**
		
Oxidizers	Acute Toxicity (fatal or toxic)	(Non-Mandatory) Aquatic Toxicity

For a more complete discussion of hazard pictograms, visit the OSHA Web site. www.osha.gov

1.2: Gloves

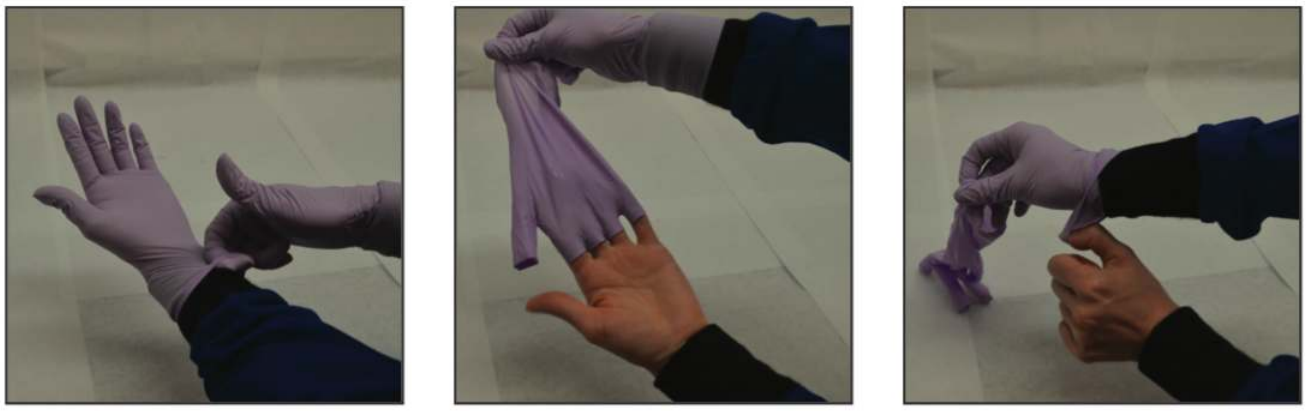
When you start to think about ways to handle chemicals to ensure your safety, you likely think about wearing gloves to minimize contact between your skin and the chemical. In reality, gloves may offer only limited protection. A common misconception is that the gloves you find in lab effectively protect the skin from solvents and other substances handled. This is simply not true.

To give you one example, the nitrile gloves that are typically made available are easily penetrated, and eventually broken down, by a common (and hazardous) solvent, acetone. There are many different gloves, but the better protection they offer, the more cumbersome they are to wear. Unfortunately, there are no gloves on the market that are both practical and offer protection against most organic chemicals encountered in undergraduate organic chemistry labs.

There are several excellent online resources to find information about the permeability of various solvents and reactants. Ansell.com has a very good overview of types of gloves and their effectiveness.

Common examples of poor glove hygiene include: touching (and contaminating) your face, skin, hair, or clothes with your gloves, touching (and contaminating) objects like phones, pens, and notebooks, that you will use outside the lab, failing to check gloves periodically for leaks and cracks, and failing to remove gloves each time you leave the lab.

In general, gloves should only be worn for short periods of time, and exchanged routinely, especially whenever a chemical has been in contact with them. The contaminated exterior of the gloves should never come in contact with your skin. Whenever you are done with your gloves, you should remove them in a safe manner and wash your hands.



(A) Start by hooking a finger into the outside of the glove.

(B) Gently pulling upwards, remove the first glove.

(C) Using your free hand, hook your thumb under the second glove.

Pulling upwards, remove the glove completely.

Figure 3.1. A safe way to remove laboratory gloves.

Adapted from *Gloves* by Alexander Sandtorv.

1.3: Obtaining Chemicals

Laboratory chemicals are typically stored, and dispensed, in a variety of containers and locations. Therefore, early hurdles in every experiment are 1) figuring out how to obtain the desired amount of the chemical and 2) figuring out how to transfer this material safely from the location where it is obtained to the location where it will be used.

Solids: Solids are usually stored in wide-mouthed bottles and weighed before use. The solid is weighed using a piece of weighing paper, and it is usually possible to transfer the solid from the bottle to the paper using a clean spatula or lab scoop. It is helpful to crease the weighing paper before weighing your solid to make transferring from the paper to your flask easier.

Liquids: Liquids are typically more difficult to handle, and the appropriate method will depend on how much liquid is needed and the specific properties of that compound. In most cases, it will be safer and more convenient to obtain a specific volume instead of weighing the liquid. For reagent quantities, a syringe, volumetric pipette, or a graduated cylinder can be used. Solvents tend to be used in larger quantities and a graduated cylinder will be used.

Adapted from *Getting chemicals to your reaction flask* by Alexander Sandtorv.

1.4: Fume Hoods

You have previously used fume hoods in general chemistry. Proper use of a fume hood is even more important in the organic laboratory due to the volatile nature of the chemicals used. The level of protection provided by a fume hood is affected by the manner in which it is used. No fume hood, however well designed, can provide adequate containment unless good laboratory practices are used.

The following practices should be followed when using a fume hood:

- Equipment and materials placed in the hood must not block hood slots, airfoils, or otherwise interfere with the smooth flow of air into the hood
- Keep all equipment and materials at least 6 inches behind the plane of the sash
- While working inside the hood, keep the hood sash closed as much as possible while still allowing comfortable working conditions. It should never be lifted above the black stopping mechanism.
- Completely close the hood sash or panels at the end of the day, when leaving experiments or chemicals unattended, and when the hood is not in use.
- Do not make quick motions into or out of the hood or walk quickly by the hood opening. These conditions will cause airflow disturbances which reduce the effectiveness of the fume hood.

1.5: Cleaning glassware

To clean glassware, use the following procedures:

- Use 2-3 mL of solvent to rinse residual organic compounds from glassware into a waste container. Most compounds should be highly soluble in the solvent. The default solvent is often **acetone** as it is inexpensive, relatively nontoxic, and dissolves most organic compounds. Residual acetone will likely evaporate from the flask, but it is acceptable for small quantities of residual acetone to be washed down the drain.
- After a preliminary rinse with acetone, or for glassware that did not contain an organic compound, the glassware should then be washed with soap and water at the bench. A spray bottle containing a dilute soap solution (Alconox) is available at each sink.
- Once clean and wiped mostly dry with a paper towel, place glassware back into the appropriate storage drawer at your hood. DO NOT leave on the drying rack when you are leaving for the day.

1.6: Waste

In the organic laboratory, there are three distinct types of waste:

Organic waste: Anything organic goes in this waste container, including organic solvents, solids that are left over from reactions or analyses, products obtained, organic extraction solvents and so on. Halogenated (chloroform) and nonhalogenated solvents (ethanol) should be kept separate.

Aqueous waste: Some aqueous solutions can be neutralized and poured down the sink. Others will be need to be disposed of in a provided container.

Solid waste: Solids, such as drying agents, and TLC plates should be placed in the fume hood to allow any solvents to be removed. Then, the solid materials can be thrown in the trash.

It is important to follow the waste disposal as directed by your instructor or TA.

The handling and disposal of chemicals is not only an important aspect of the day-to-day work in the lab, but also for the environment, ecology, and the workers and other organisms that surround us.

1.7: Stress and Anxiety

It is not unusual to feel worried and anxious about working in organic chemistry labs. Your concerns might range from anticipated difficulties in the lab (handling of chemicals, dangerous substances) or just your desire to do a good job.

Here are a few strategies that can help mitigate any feelings of anxiety and stress:

1. **Plan ahead.** It sounds simple, but as we have seen from the first chapter, planning is the most important step before undertaking an experiment. This will prepare ou for the things you need to do and the decisions you will need to make.
2. **Communicate with your TA.** If you find yourself feeling overwhelmed or deeply uncomfortable while in the lab, communicate this to your TA. If necessary, the TA, student, and professor will work together to find a plan to assist the student.
3. **Take a break.** If the experience simply gets overwhelming, tell your TA, and take a short break. Leave your lab coat and goggles behind and go out in the hall, or take a bathroom break.

Remember, lab should be a positive, engaging and interactive learning experience. The primary objective in the lab is to focus on what you are doing and learning.

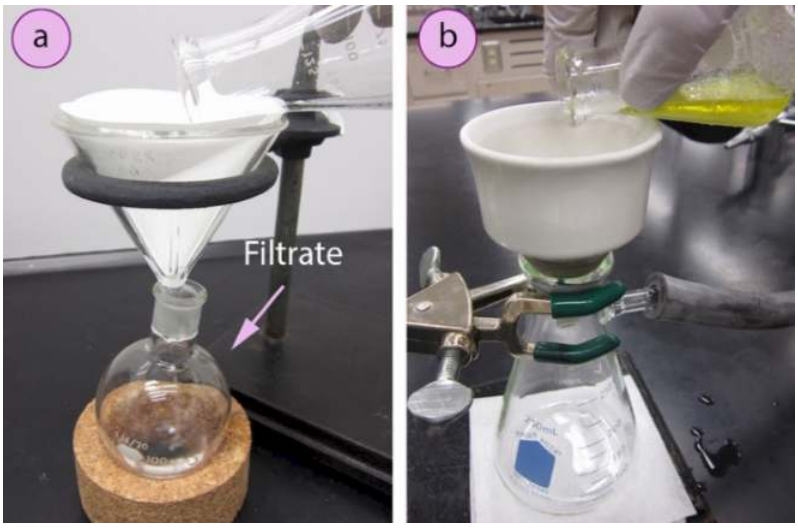
CHAPTER OVERVIEW

2: Filtration

2.1: Filtration Overview

There are many methods used to separate a mixture containing a solid and liquid. If the solid settles well, the liquid can sometimes be poured off (decanted).

The most common methods of solid-liquid separation in the organic lab are gravity and vacuum filtration. Gravity filtration refers to pouring a solid-liquid mixture through a funnel containing a filter paper, allowing the liquid to seep through while trapping the solid on the paper. Vacuum filtration is a similar process with the difference being the application of a vacuum beneath the funnel in order to pull liquid through the filter paper with suction.



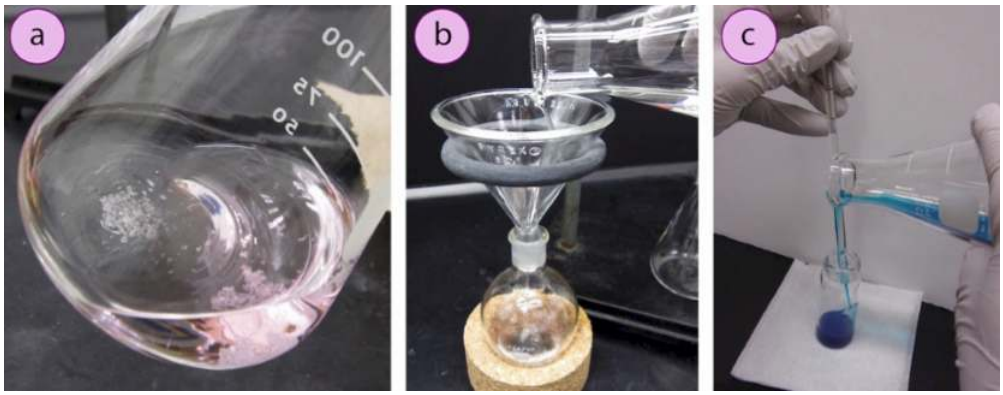
a) Gravity filtration, b) Vacuum filtration

Gravity and vacuum filtration have pros and cons, but what helps decide which method to use is generally whether the solid or filtrate is to be retained. The "**filtrate**" refers to the liquid that has passed through a filter paper. Gravity filtration is typically used when the filtrate is retained, while vacuum filtration is used when the solid is retained.

Adapted from *Overview of Methods* by Lisa Nichols.

2.2: Decanting

When there is a need to separate a solid-liquid mixture, on occasion it is possible to pour off the liquid while leaving the solid behind. This process is called decanting, and is the simplest separation method. Decanting is often used to remove hydrated sodium sulfate (Na_2SO_4) from an organic solution. The sodium sulfate often clings to the glassware, enabling the liquid to be poured off. If liquid is to be poured into a small vessel, a funnel could be used or liquid poured down a glass stirring rod to direct the flow. Unfortunately, there are many mixtures that do not decant well.



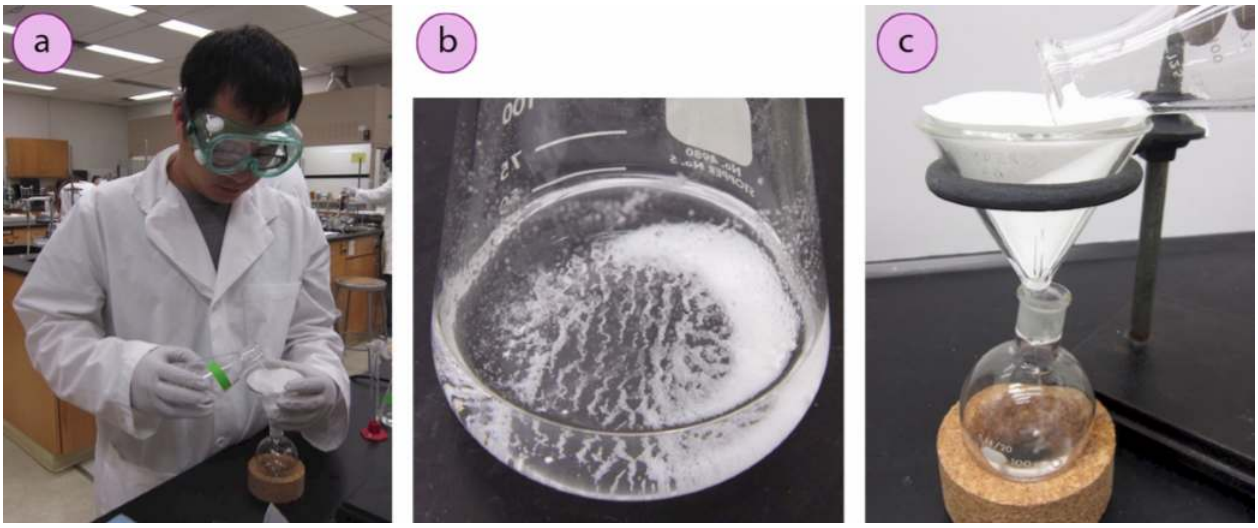
a) Sodium sulfate sticking to the glassware, b) Decanting a solid-liquid mixture, c) Using a glass stirring rod during decanting.

Adapted from *Decanting* by Lisa Nichols.

2.3: Gravity Filtration

When there is a need to separate a solid-liquid mixture, it is common that the particles are so fine that they swirl and disperse when the flask is tilted. These mixtures cannot be decanted, and an alternative method is gravity filtration. Gravity filtration is generally used when the filtrate (liquid that has passed through the filter paper) will be retained, while the solid on the filter paper will be discarded.

A common use for gravity filtration is for separating anhydrous magnesium sulfate ($MgSO_4$) from an organic solution that it has dried. Anhydrous magnesium sulfate is powdery, and with swirling in an organic solvent creates a fine dispersal of particles like a snow globe.



a) An organic solution dried with anhydrous magnesium sulfate, b) Gravity filtration of this solution.

To gravity filter a mixture, pour the mixture through a quadrant-folded filter paper or fluted filter paper in a funnel and allow the liquid to filter using only the force of gravity. It is best to pour as if attempting to decant, meaning to keep the solid settled in the flask for as long as possible. When solid begins to pour onto the filter paper, it has the possibility of clogging the filter paper pores or slowing filtration.

Adapted from *Gravity Filtration* by Lisa Nichols.

2.4: Vacuum Filtration

Vacuum filtration is the standard technique used for separating a solid-liquid mixture when the goal is to retain the solid (for example in crystallization). Similar to gravity filtration, a solid-liquid mixture is poured onto a filter paper, with the main difference being that the process is aided by suction beneath the funnel.

The process has advantages and disadvantages in comparison to gravity filtration.

Advantages:

1) Vacuum filtration is much faster than gravity filtration, often taking less than one minute with good seals and a good vacuum source.

2) Vacuum filtration is more efficient at removing residual liquid, leading to a purer solid.

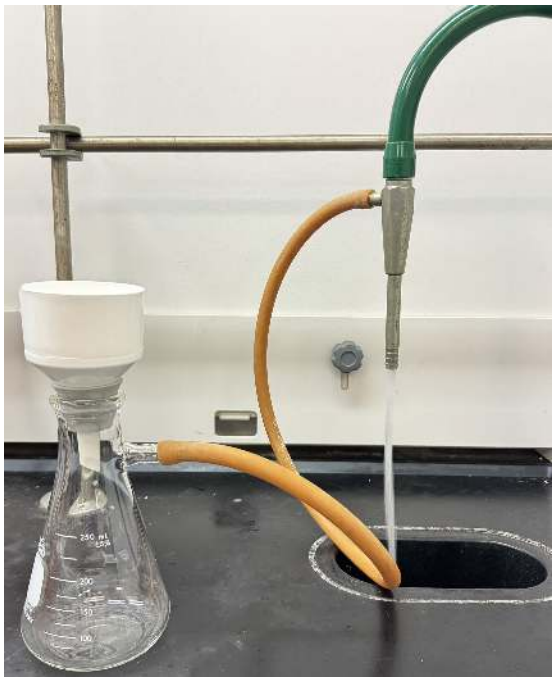
Disadvantages:

1) The force of suction may draw fine crystals through the filter paper pores, leading to a quantity of material that cannot be recovered from the filter paper, and possibly an additional quantity that is lost in the filtrate.

Vacuum filtration

Water Aspirator

A vacuum source is necessary for this filtration type. A water aspirator is an inexpensive attachment to a water spigot that creates a vacuum in the Erlenmeyer flask. As water flows through the faucet and the aspirator, suction is created in the flask.

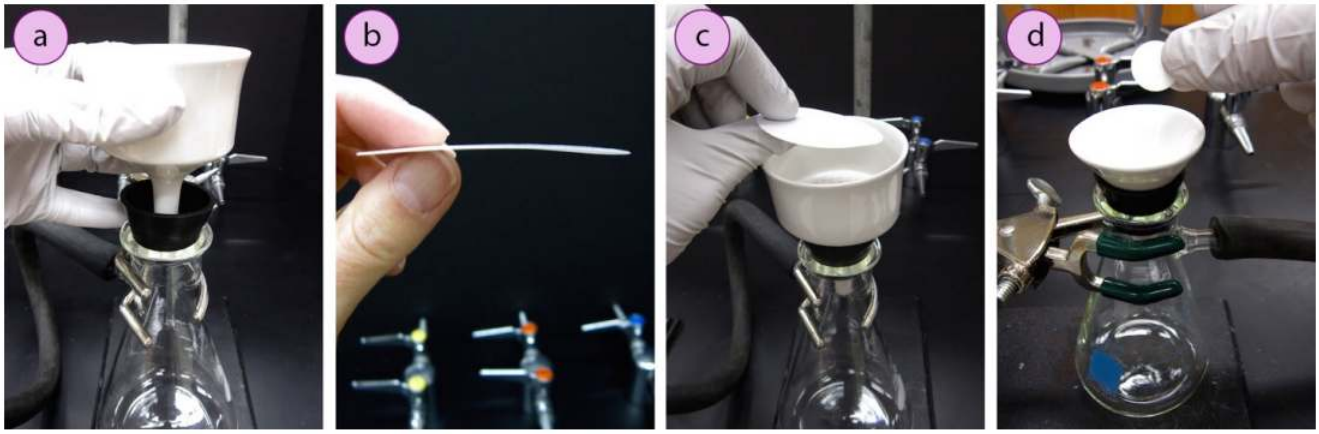


Water aspirator set up.

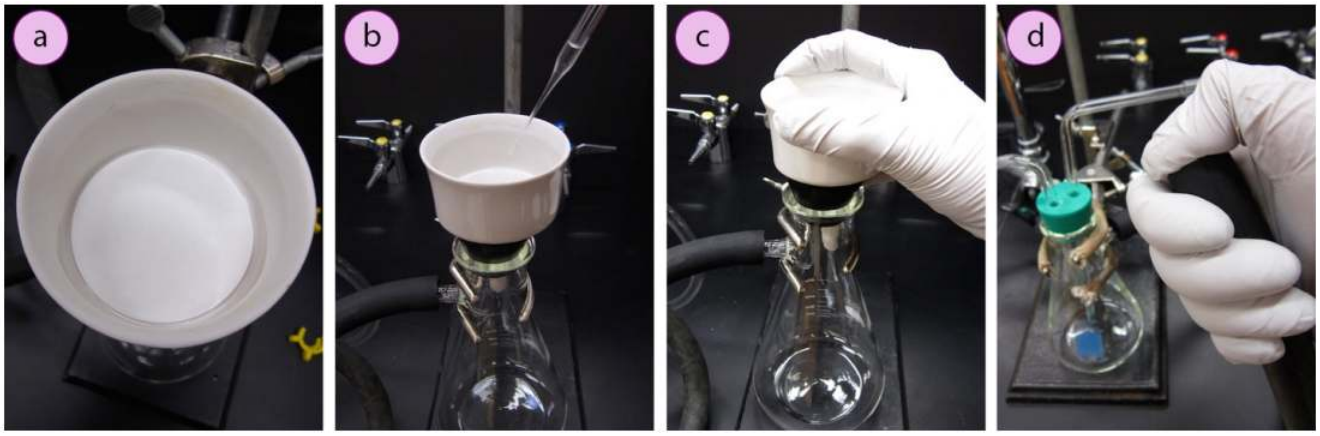
Step-by-Step Procedures

Assemble the suction filtration flask

1. Attach a thick-walled rubber hose to the Erlenmeyer's side arm. Connect this thick tubing to the water aspirator as well. It is best to not bend or strain the tubing as much as is practical, as this may cause poor suction.
2. Place a rubber sleeve and Buchner funnel atop the side-arm Erlenmeyer flask.
3. Obtain a filter paper that will fit perfectly into the Buchner funnel. The paper should cover all the holes in the funnel. Note: If you will be weighing the final product, be sure to weigh the top of the Buchner funnel with the filter paper in place.

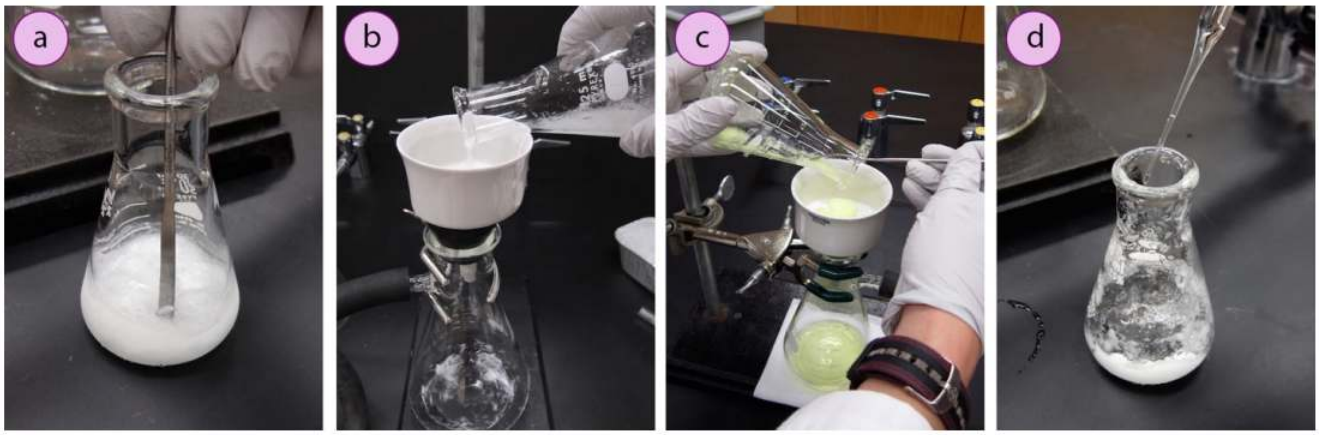


a) Placing the Buchner funnel in a rubber sleeve and Erlenmeyer flask, b) Concavity of a filter paper, c) Placing filter paper in a Buchner funnel, d) Placing filter paper in a Hirsch funnel.



a) Filter paper in the funnel, b) Wetting the filter paper with solvent, c) Pressing the Buchner funnel to create a good seal, d) Testing the suction of the aspirator.

4. Turn on the faucet connected to the water aspirator to create a strong flow of water (the degree of suction is related to the water flow). Wet the filter paper with cold solvent (using the same solvent used in crystallization, if applicable).
5. Suction should drain the liquid and hold the moist filter paper snugly over the holes in the filter. If the solvent does not drain or suction is not occurring, you may need to press down on the funnel to create a good seal between the glass and rubber sleeve.



a) Using a spatula to dislodge a thick solid from the glass, b) Filtering, c) Using a spatula to scoop a thick solid onto the filter paper, d) Rinsing residual solid from the flask using cold solvent.

Filter the Mixture

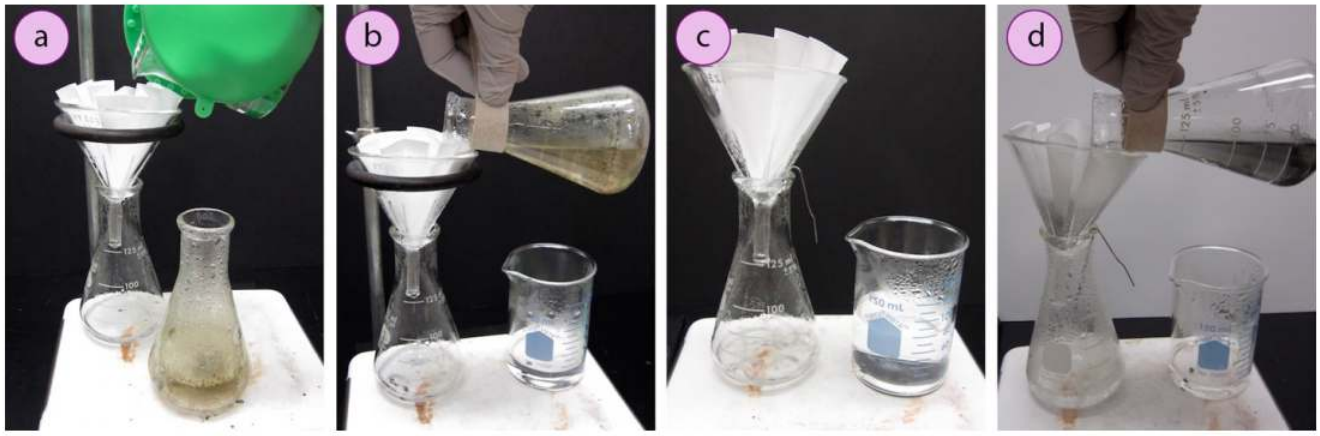
6. Swirl the mixture to be filtered in order to dislodge solid from the sides of the flask. If the solid is very thick, use a spatula or stirring rod to free it from the glass.
7. With a quick motion, swirl and dump the solid into the funnel in portions. If the solid is very thick, scoop it out of the flask onto the filter paper. It's best if the solid can be directed toward the middle of the filter paper, as solid near the edges may creep around the filter paper.
8. A small amount of chilled solvent (1-2mL for macroscale work) can be used to help rinse any residual solid from the flask into the funnel. In crystallization, it is not wise to use an excessive amount of solvent as it will decrease the yield by dissolving small amounts of crystals. Again, press on the funnel to create a good seal and efficient drainage if necessary.
9. After a few minutes, turn off the water and disconnect the flask. Take the top portion of the Buchner funnel off and cover with a Kimwipe. It is best to leave this to dry over the next week before completing any analysis.

Adapted from *Suction Filtration* by Lisa Nichols.

2.5: Hot Filtration

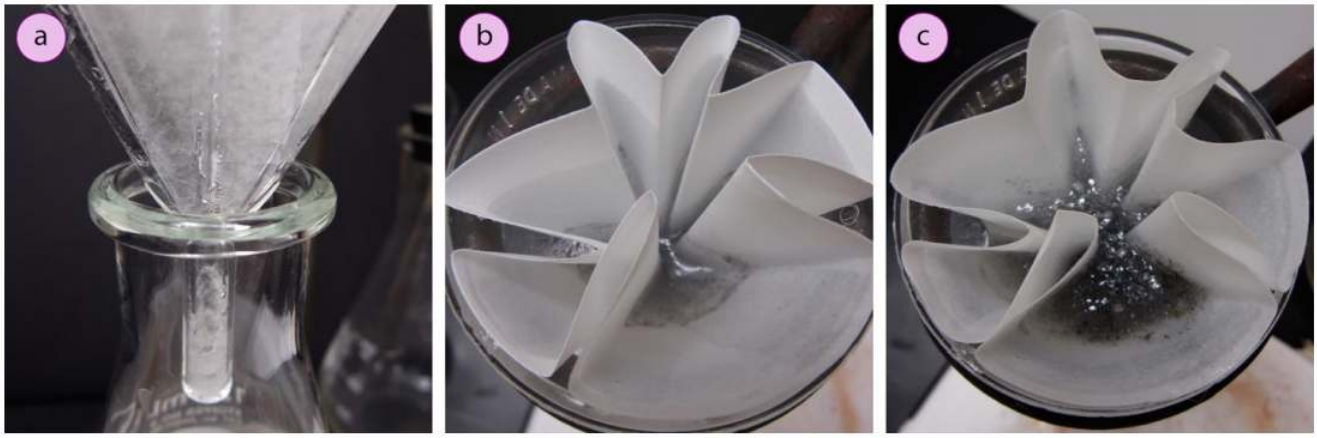
A hot filtration is generally used in crystallization, when a solid contains impurities that are insoluble in the crystallization solvent. It is also necessary in crystallization when charcoal is used to remove highly colored impurities from a solid, as charcoal is so fine that it cannot be removed by decanting.

A hot filtration is performed by first pouring a few mL of solvent through a funnel containing a "fluted filter paper". A fluted filter paper has many indentations and high surface area, which allows for a fast filtration. The funnel is allowed to get hot, while the mixture to be filtered is brought to a boil. The boiling mixture is then poured through the filter paper in portions.



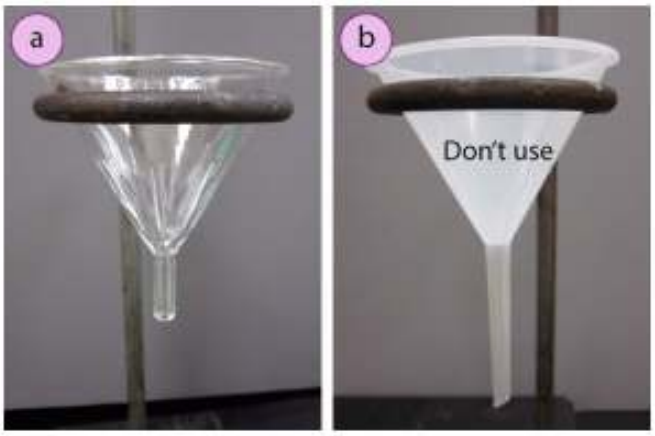
a) Pouring solvent through the funnel hot, b) Filtering a solution containing insoluble impurities

A hot filtration is used for filtering solutions that will crystallize when allowed to cool. It is therefore **important that the funnel is kept hot** during filtration through contact with hot solvent vapors, or crystals may prematurely form on the filter paper or in the stem of the funnel.



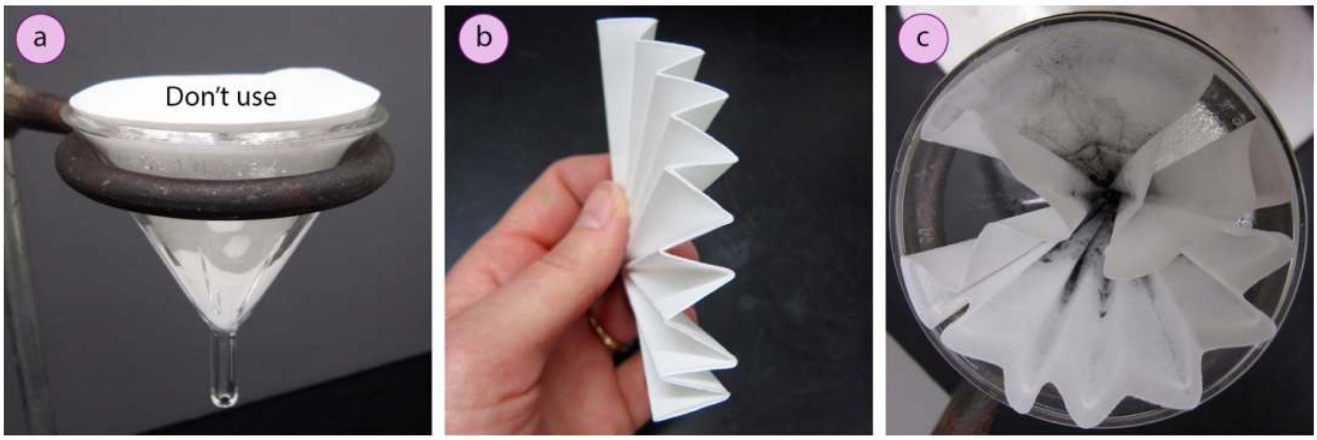
A filtration that was conducted at too low a temperature: a) Crystallization occurred in the stem of the funnel, b) A sheer can be seen on the filter paper, representing solid that has crystallized somewhat on the filter paper, c) Obvious crystals formed on the filter paper.

Crystallization on the filter paper can clog the setup and cause a loss of yield (as the filter paper will be later thrown away). Crystallization in the stem hinders filtration, and can act as a plug on the bottom of the funnel. An advantage of *hot* filtration is that the boiling solvent in the filter flask helps to dissolve crystals that prematurely form in the stem of the funnel. With hot filtration, it is advised to use a short-stemmed funnel, instead of a long-stemmed funnel, as material is less likely to crystallize in a short stem.



a) Short-stemmed funnel, b) Long-stemmed funnel (don't use for hot filtration).

As it is essential that a solution filters quickly before it has a chance to cool off in the funnel, a "**fluted filter paper**" is commonly used instead of the quadrant-folded filter paper sometimes used with gravity filtration. The greater number of bends on the fluted filter paper translate into increased surface area and quicker filtration. The folds also create space between the filter paper and glass funnel, allowing for displaced air to more easily exit the flask as liquid drains.

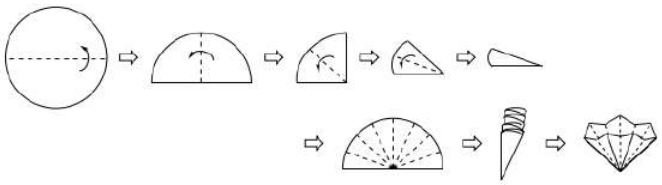


a) Quadrant-folded filter paper (not recommended for hot filtration), b) A fluted filter paper after folding, c) Unfolded fluted filter paper during hot filtration.

Step-by-Step Procedures

Prepare the Filtration Setup

1. Flute a filter paper of the correct size for your short stem funnel into an accordion shape. When placed in the funnel, the paper should not be shorter than the top of the funnel, or the solution might slip past the filter paper when poured.

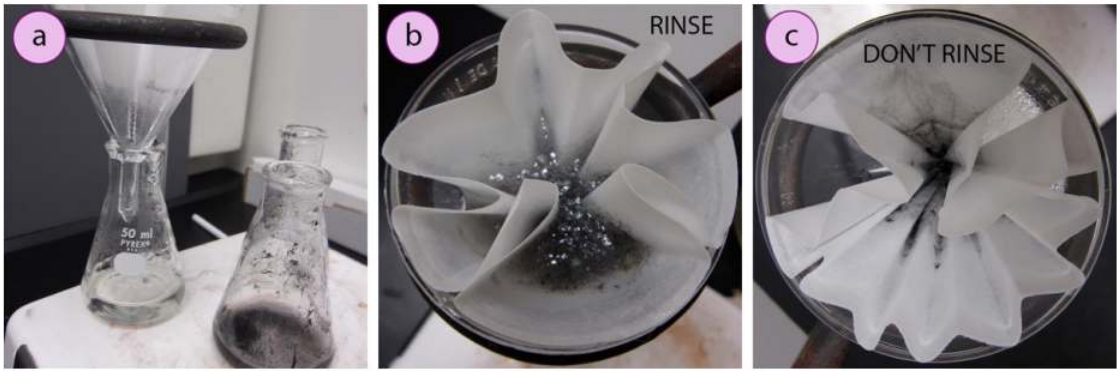


Creating a fluted filter paper. The dotted lines represent locations to crease and fold the filter paper. The arrows show the direction of folding.

2. With a clean Erlenmeyer flask of the correct size for the crystallization beneath the funnel and on the hot plate, pour a few mL of hot solvent into the funnel.
3. Allow the solvent to boil and get the entire setup hot.

Filter the Solution in Portions

4. When the filter flask is quite hot, and the solution to be filtered is boiling, pour the boiling mixture into the filter funnel in portions. Touch the flask to the filter paper in the funnel as you pour.
5. When not pouring the mixture to be filtered, return the flask to the heat source.
6. When the mixture is completely filtered, set the empty flask on the benchtop (**safety note:** do not heat an empty flask, or it may crack). Inspect the funnel: if crystals are seen on the filter paper, rinse with a few mL of boiling solvent to dissolve them.



a) Flask is returned to the heat source in between pouring, Filter papers that b) should be rinsed, c) should not be rinsed.

Adapted from *Hot Filtration* by Lisa Nichols.

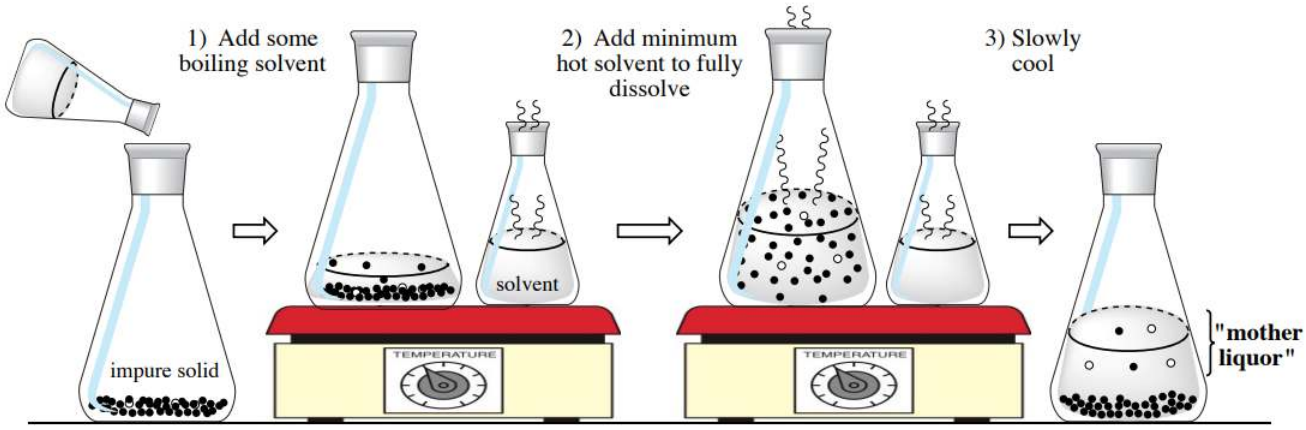
CHAPTER OVERVIEW

3: Recrystallization

3.1: Overview of Recrystallization

The uniformity and structural repetition differentiate a crystal from an amorphous solid. For this reason, there is a difference between precipitation (the rapid formation of a solid), and crystallization (the slow growth of a solid with regular microscopic structure). Precipitated solids tend to have lower purity than crystals, which is why crystallization is of interest to organic chemists.

Crystallization is used in the chemistry laboratory as a purification technique for solids. An impure solid is completely dissolved in a minimal amount of hot, boiling solvent, and the hot solution is allowed to slowly cool. The developing crystals ideally form with high purity, while impurities remain in the saturated solution surrounding the solid (called the "**mother liquor**"). The crystallized solid is then filtered away from the impurities. A diagram of the crystallization procedure is shown below.

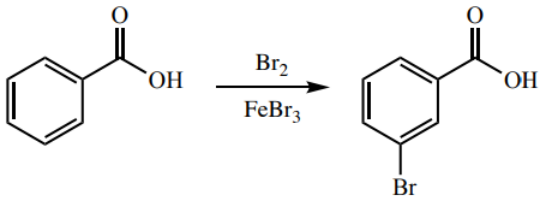


Procedural sequence for crystallization

Adapted from *Overview of crystallization* from Lisa Nichols.

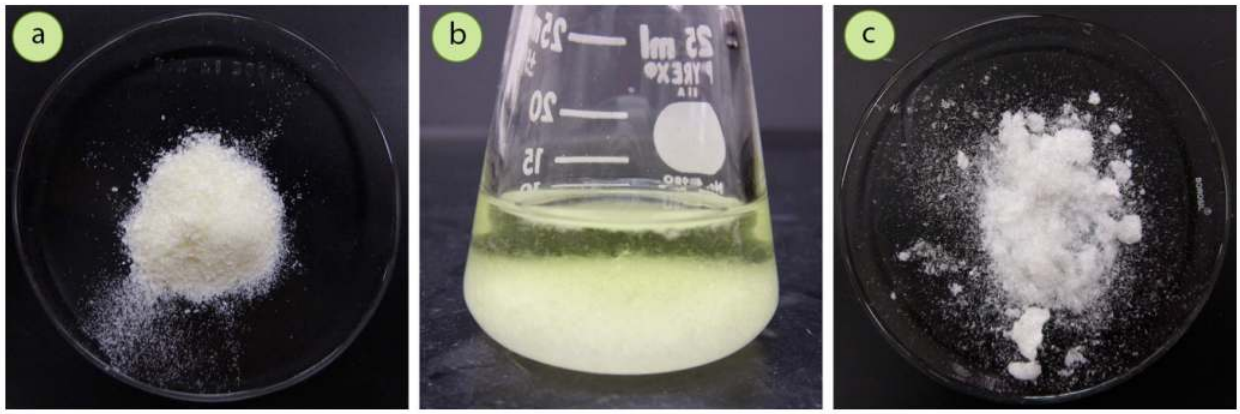
3.2: Purification of Products

If the crude product of a chemical reaction is a solid, it may be crystallized in order to remove impurities. For example, benzoic acid can be brominated to produce *m*-bromobenzoic acid. The crude solid product could very likely contain unreactive benzoic acid, and this impurity could be removed through crystallization.



Bromination of benzoic acid to produce m-bromobenzoic acid

To demonstrate, a mixture containing roughly 85 mol% *p*-bromobenzoic acid (a solid) contaminated with 15 mol% benzoic acid (another solid) had a yellow tint, and after crystallization the resulting solid was pure white. The crystallization appeared to purify the mixture based on the slight improvement in color. The purity of the recrystallized product can be determined by various techniques, such as determining the melting point.

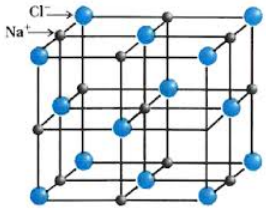


a) A mixed solid containing 85 mol% p-bromobenzoic acid and 15 mol% benzoic acid, b) Crystallization of the mixture using ethanol, c) Crystallized solid

Adapted from *Purification of Product Mixtures* by Lisa Nichols.

3.3: Crystallization Theory

Crystallization is an excellent purification technique for solids because a crystal slowly forming from a saturated solution tends to selectively incorporate particles of the same type into its crystal structure (modeled below). A developing solid will tend to incorporate particles of the same type in order to create the lowest energy solid, and exclude impurities that disrupt the idealized packing of the solid. This process works best if the impurity is present as a minor component of the crude solid.



Crystal structure of NaCl

The difference in crystal lattice energy between pure and impure solids is marginal, so a solution must be cooled SLOWLY to allow for differentiation. If a hot solution is plunged immediately into an ice bath, the system will favor the formation of a solid (any solid!) so strongly that there may be little preference for purity. Impurities can become engulfed in the developing solid and trapped as solutes are deposited unselectively onto the growing solid. It is only when enough time is allowed for equilibration between the developing solid and the solution that the lowest energy pure solid is formed.

The crystallization process can be though of as the crystal lattice "grabbing" solutes from solution. If the process is hurried, solutes may be "grabbed" indiscriminately, and once embedded in the interior of the solid they become trapped as the equilibrium between solid and solution happens only on the surface.

Along with increased purity, a slow crystallization process also encourages the growth of larger crystals. The pictures below shows crystallization of acetanilide from water with two different rates. The crystals grown in A were formed much more quickly, and are smaller than the slower grown, larger crystals shown in B. There are several benefits to larger crystals. Most importantly they tend to be purer than small crystals. They are also more easily collected by vacuum filtration, as very small crystals may pass through or wedge themselves into the pores of the filter paper.

a) These crystals of acetanilide formed from water are perfectly acceptable, but small, b) These crystals of acetanilide were grown much more slowly, and are larger (note crystal indicated with the arrow).

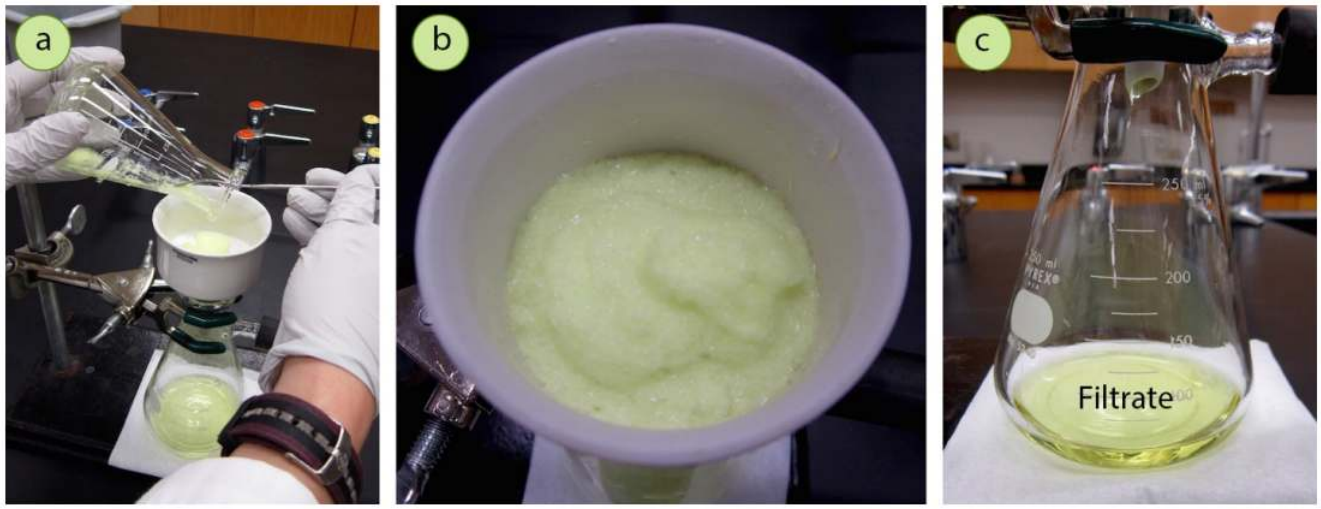
The quantity of solvent used in crystallization is usually kept to a minimum, to support the goal of recovering the **maximum amount of crystals**.

Every solid has partial solubility in the solvents used, even at cold temperatures. In the picture below, the cold solvent surrounding a yellow solid is tinted yellow as some compound dissolves. Solids that appear insoluble in a solvent do in fact have a (normally small) portion of material that dissolves.

Crystallization is most common with solids that have moderate solubility at low temperatures, so that heat can "tip them over the edge" to completely dissolve. This means that in practice there *will* be a quantity of compound that dissolves in the mother liquor at low temperatures, which can be significant depending on the compound's solubility. Use of the minimal amount of hot solvent lessens the quantity of compound that is lost to the mother liquid.

Yellow solid (benzil) in contact with cold ethanol

A loss of recovery should be expected when performing a crystallization. Although there are ways to maximize the return of crystals, a portion of the desired compound will always be lost. A portion of the compound of interest will remain dissolved in the mother liquor and be filtered away, as shown in the pictures below.



Loss of yield due to the solubility of compound in the cold solvent: a) Vacuum filtration, b) Recovered yellow solid, c) Yellow filtrate, indicating some yellow compound remained dissolved in the mother liquor

The loss of solid material can also be seen with every manipulation of the solid. Only the *majority* of the crystals can be delicately scraped off the glassware, Buchner funnel, and filter paper, and there will always be a residue that is left behind. When there is such an obvious loss of yield from solid clinging to glassware, it may seem wise to use solvent to rinse additional solid out of the flasks. A few rinses with cold solvent are indeed recommended, but it is not recommended to use solvent excessively in an attempt to recover every granule of solid. The more solvent that is used, the more compound will dissolve in the cold mother liquor, decreasing the yield. The loss of material due to residue on the glassware is an unfortunate, but accepted aspect of this technique.

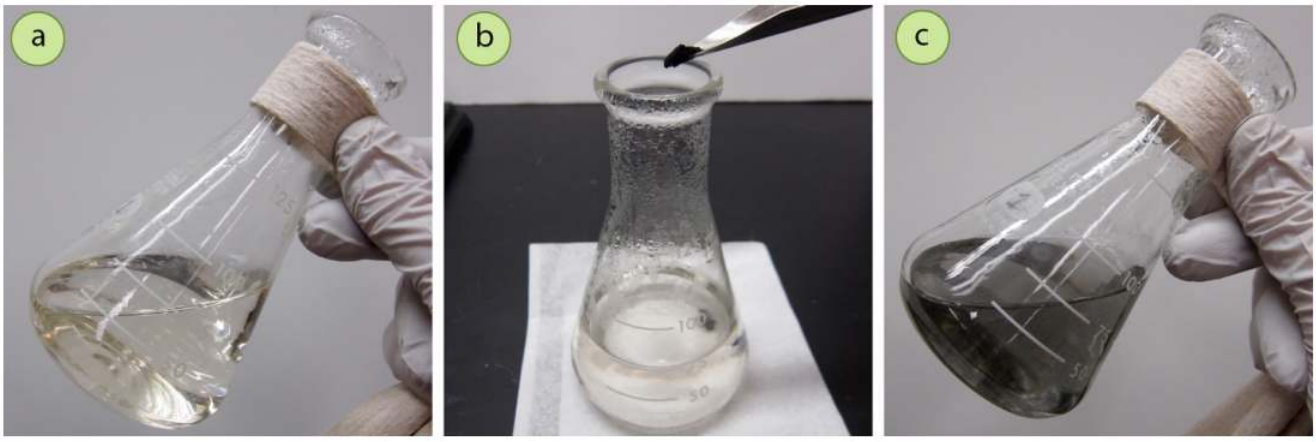
It is common for new organic chemistry students to be disappointed by a low yield and to worry that it is somehow their fault. Students are often inclined to cite "user error" as a main cause for a loss of yield in any process. Although spilling solid on the benchtop or addition of too much solvent will of course compromise the yield, the modest recovery of acetanilide in this section should demonstrate that sometimes low yields are inherent to the process and the chosen solvent.

Adapted from *Crystallization Theory* by Lisa Nichols.

3.4: Activated Charcoal

Activated charcoal is sometimes used to **remove small amounts of colored impurities** from solution. Activated charcoal has a high affinity for conjugated compounds, whose flat structures wedge themselves well between the graphene sheets. The quantity used should be limited, as charcoal adsorbs all compounds to some extent and could lead to a lower recovery of the desired compound. Charcoal should of course not be used if the product itself is colored.

Decolorizing charcoal (Norit) is added after a solid has been dissolved in the minimum amount of hot solvent. A small portion should be used at first; just as much that can fit on the tip of a spatula (the size of half a pea). **When adding charcoal, it's recommended to temporarily remove the solution from the heat source** or the charcoal's high surface area may cause superheated areas to immediately boil. If charcoal is added directly to a near boiling solution, the solution may boil over.



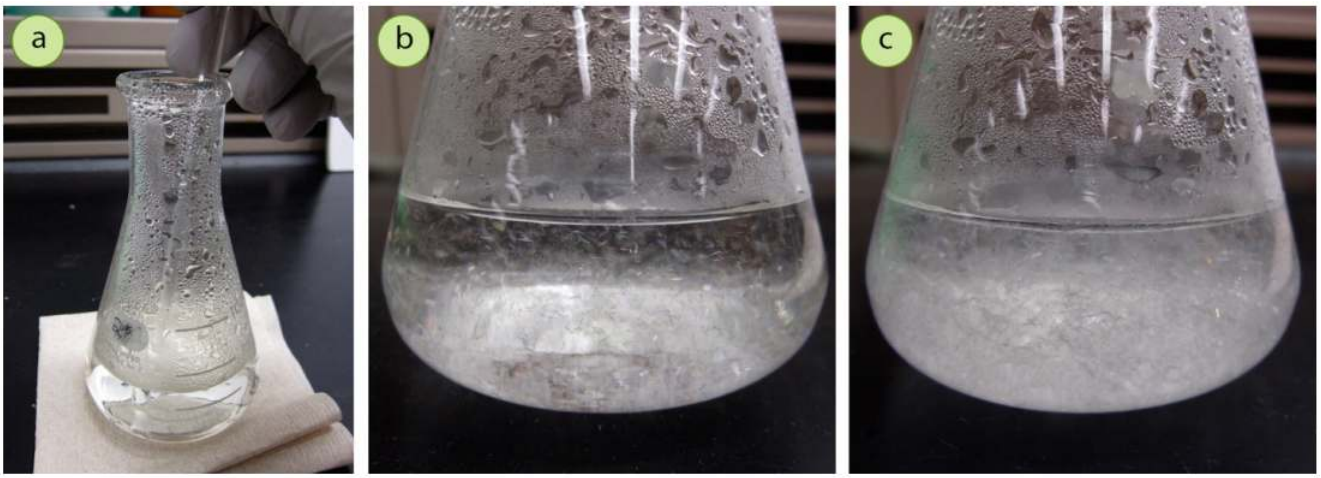
a) Solution before addition of charcoal, b) An appropriate quantity of charcoal to start with (tip of the spatula), c) Solution after swirling with charcoal.

Adapted from *Charcoal* by Lisa Nichols.

3.5: Inducing Recrystallization

At times, crystals will not form even when a solution is supersaturated. Crystallization may need to be initiated if the solution fails to produce crystals even when it is noticeably cooler than originally.

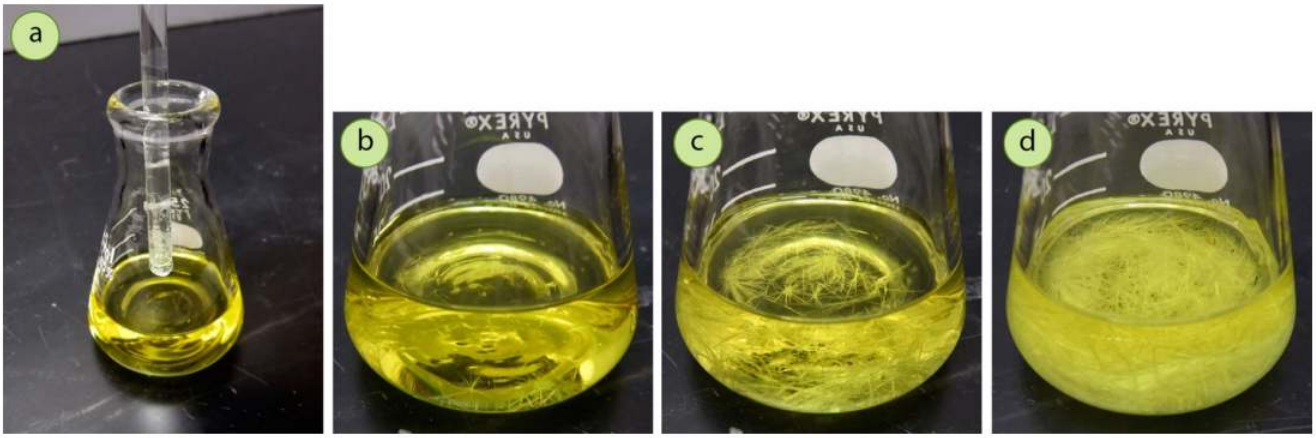
The easiest method to initiate crystallization is to scratch the bottom or side of the flask with a glass stirring rod, with enough force that the scratching is audible. Crystallization often begins immediately after scratching, and lines may be visible showing crystal growth in the areas of the glass that were scratched.



a) Scratching to initiate crystallization for acetanilide in water, b+c) Crystallization

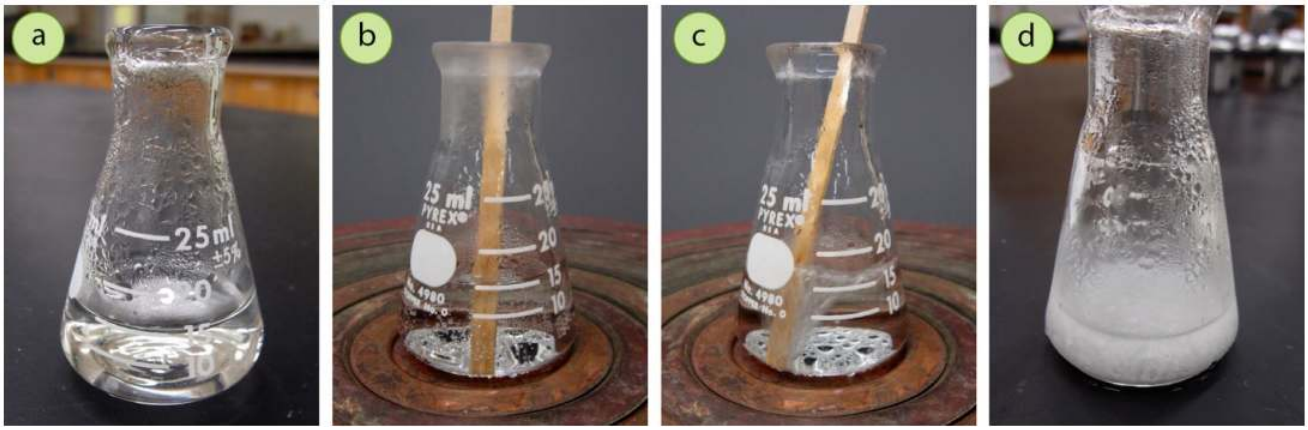
There are a few other methods that can be used with your TA or Instructor to initiate crystallization when scratching fails:

- Add a "seed crystal": a small speck of crude solid saved from before the crystallization was begun, or a bit of pure solid from a reagent jar. Seed crystals create a nucleation site where crystals can begin growth.
- Dip a glass stirring rod into the supersaturated solution, remove it, and allow the solvent to evaporate to produce a thin residue of crystals on the rod. Then touch the rod to the solution's surface, or stir the solution with the rod to dislodge small seed crystals.



a) glass stirring rod dipped into a solution of benzil, and residual solid formed on the rod after evaporation of the ethanol solvent, b-d) Crystallization, which began after the rod was stirred in the solution to dislodge the seed crystals.

If scratching and seed crystals do not initiate crystallization, it's possible there is too much solvent present that the compound remains completely soluble. To test if this is the case, return the solution to a boil and reduce the volume of solvent, perhaps by half. Allow the reduced solution to cool and see if solid forms. If it does, the quantity of solvent was definitely the problem.



a) No crystals formed when cooled, b+c) Reducing the solvent volume on a steam bath, d) Crystals then formed when cooled.

You can also use a lower temperature bath to try and encourage crystal formation. A salt water-ice bath provides cooler temperatures than water-ice baths alone.

Adapted from *Initiating Crystallization* by Lisa Nichols.

3.6: Procedure Summary

The crystallization in this section shows purification of a roughly 1 g sample of old N-bromosuccinimide (NBS), which was found in its reagent bottle as an orange powder. The crystallization uses **water as the solvent**, which has no flammability issues, and so a hotplate is used.

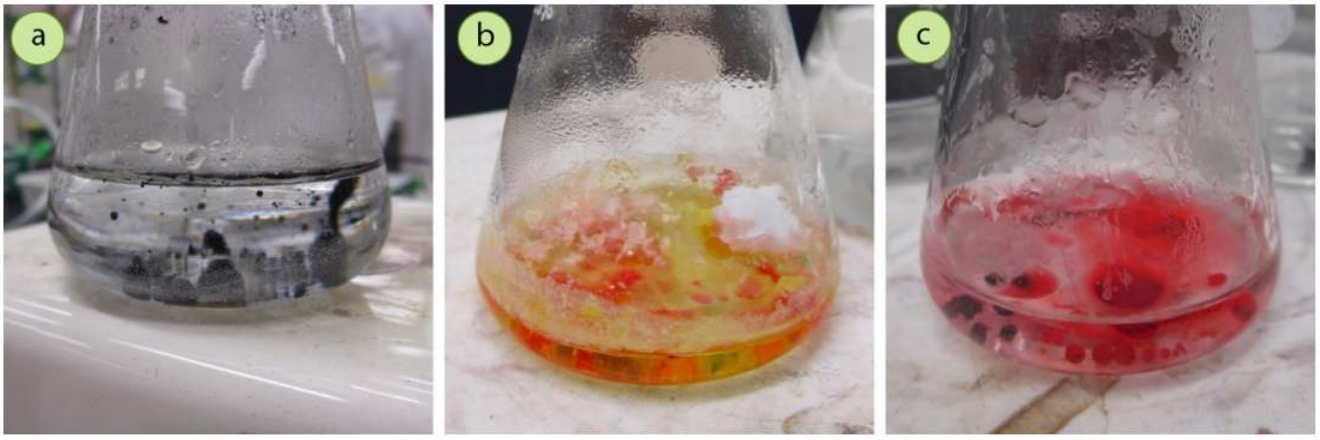
Prepare the Setup

1. Transfer the impure solid to be crystallized into an appropriately sized **Erlenmeyer flask**. If the solid is granular, first pulverize with a glass stirring rod. It is important that the flask be not too full or too empty during the crystallization. If the flask will be greater than half-full with hot solvent, it will be difficult to prevent the flask from boiling over. If the flask will contain solvent to a height less than 1 cm, the solution will cool too quickly. A rough guide is to use a flask where the sample just covers the bottom in a thin layer.
2. Place some solvent in a beaker or Erlenmeyer flask and bring to a gentle boil.

Add the Minimum Amount of Hot Solvent

3. When the solvent is boiling, grasp the beaker with a hot hand protector. Pour a small portion of boiling solvent into the flask containing the impure solid, to coat the bottom of the flask.
4. Place the flask containing the impure solid and solvent on the heat source and bring the solution to a gentle boil.
5. Add solvent in portions, swirling to aid in dissolution, until the solid just dissolves. Note that it may take time for a solid to completely dissolve. Each addition should be allowed to come completely to a boil before adding more solvent, and some time should be allowed between additions. Not allowing time for dissolution and consequently adding too much solvent is a main source of error in crystallization.
6. If insoluble solid impurities are present, the solution should be filtered. Colored impurities can also be removed at this point with charcoal.

Note: It is not uncommon for droplets of liquid to be seen during the heating process (see photos below). This is when the material "oils out", or melts before it dissolves. If this happens, the liquid droplets are now the compound you are crystallizing, so continue adding solvent in portions until the liquid droplets fully dissolve.



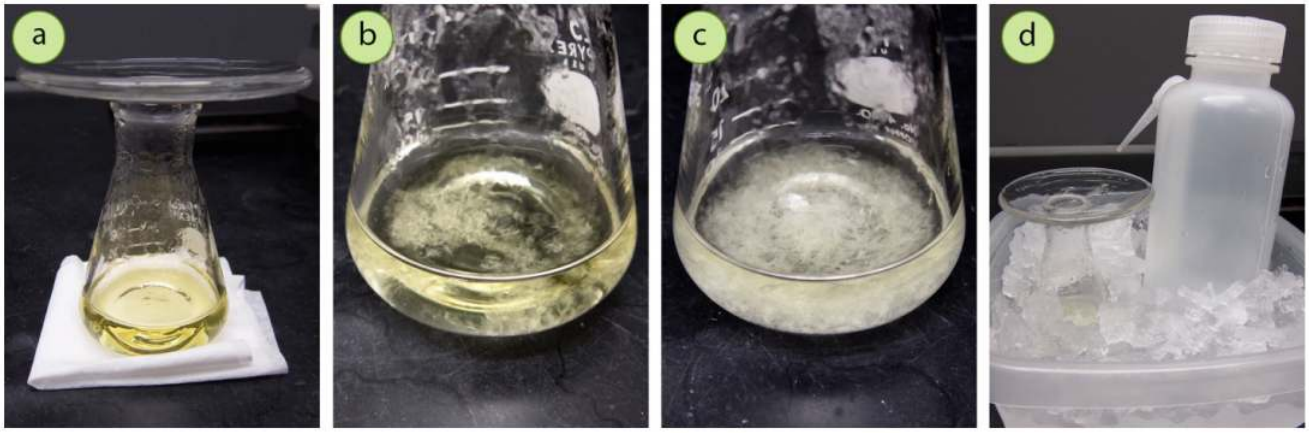
Examples of samples "oiling out" during dissolving: a) Oily droplets of acetanilide in a solution decolorized with charcoal, b+c) Oily droplets containing methyl red and acetanilide in water

Allow the Solution to Slowly Cool

When the solid is just dissolved, remove the flask from the heat source using a hot hand protector and set it aside to cool. Allow the solution to slowly come to room temperature. As the solution cools, eventually solid crystals should form. After crystallization has begun, the crystals should slowly grow as the temperature decreases. An ideal crystallization takes between 5-20 minutes to fully crystallize, depending on the scale.

When the solution is at room temperature, place the flask into an ice bath (ice-water slurry) for 10-20 minutes to lower the compound's solubility even more and maximize crystal formation. Also place a portion of solvent in the ice bath, to be used later for rinsing during vacuum filtration.

Use vacuum filtration to recover the solid from the mixture.



a-c) Cooling and crystallization of NBS, d) Further cooling in an ice bath along with some solvent (in the bottle) for rinsing.

Adapted from *Single Solvent Crystallization* by Lisa Nichols.

CHAPTER OVERVIEW

4: Melting Point

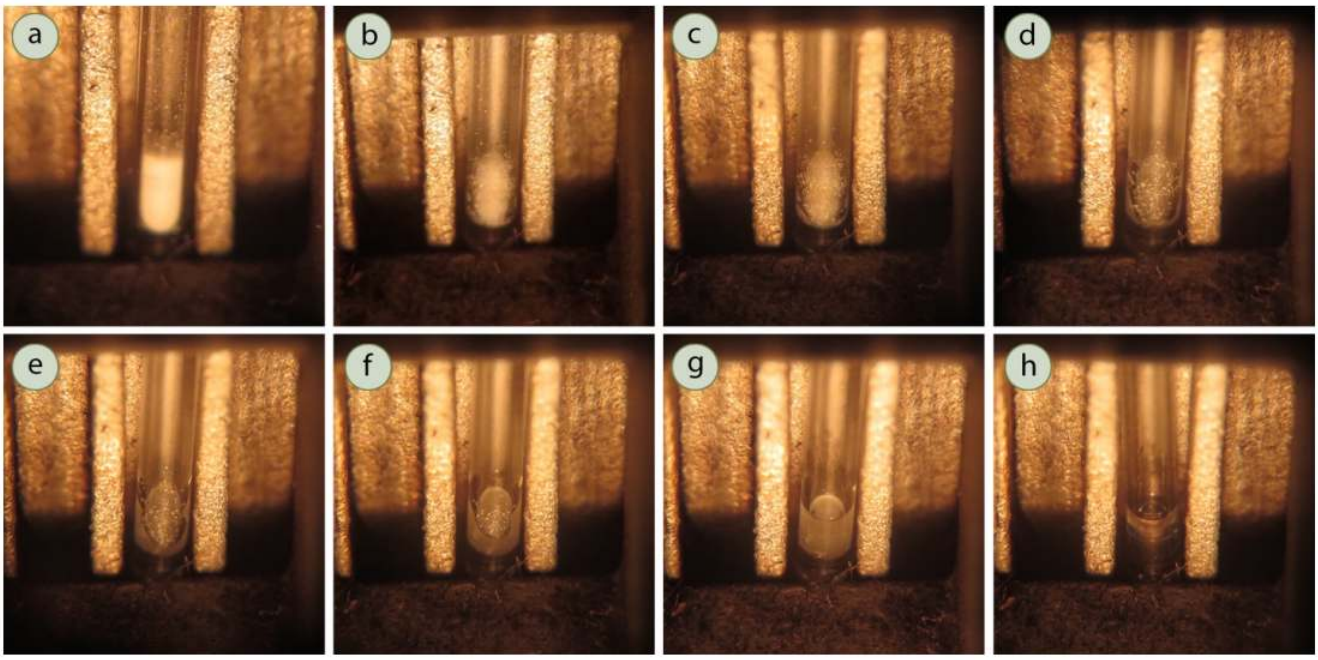
4.1: Melting Point Overview

Measurement of a solid compound's melting point is a standard practice in the organic chemistry laboratory. The melting point is the temperature where the solid-liquid phase change occurs. In some reference books it is listed as a single value (e.g. 98^{o}C), but in chemical catalogs it is more often listed as a range of values (e.g. 96-98^{o}C). The melting "point" is therefore more of a melting "range," and in part, reflects how melting points are experimentally determined.

A melting point is determined by loading a small amount of sample into a capillary tube (shown below), and then slowly heating the sample. The series of photos below shows a close-up view of a sample inside a melting point apparatus, where the sample is slowly heated through contact with hot vertical metal blocks on either side of the capillary tube. The sample is kept small in this technique to ensure adequate heat transfer between the metal and sample.

White solid (indicated with an arrow) in a capillary tube

The first value recorded for the melting range is with the very first appearance of liquid. As this temperature is approached, the solid may begin to glisten, and the temperature is recorded with the first hint of liquid movement (a droplet) inside the tube. The second value recorded for the melting range is with the melting of the entire sample, which occurs when all areas of opaque solid have turned into a transparent liquid.



Time-lapse melting of benzoic acid: a) Well below the melting point, b) "Glistening" of the solid, c) First liquid droplet is seen (the temperature is recorded as the lower value of the melting range), d-g) Melting, h) Sample is completely melted (the temperature is recorded as the upper value of the melting range)

Adapted from *Overview of Melting Point* by Lisa Nichols.

4.2: Uses of Melting Points

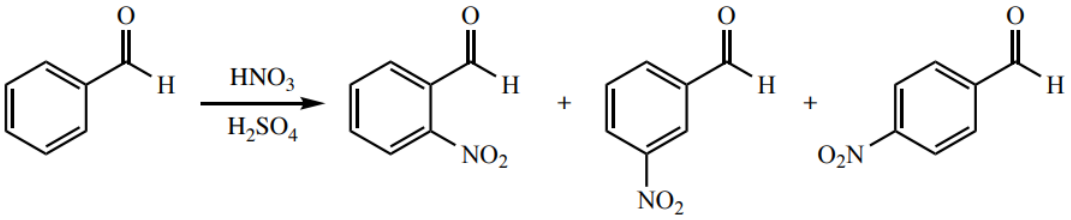
There are several reasons to determine a compound's melting point: it is useful in supporting the identification of a compound, as well as serving as a rough guide to the relative purity of the sample.

Identification

As a compound's melting point is a physical constant, it can be used to support the identity of an unknown solid. The melting point can be looked up in a reference book (this value would then be called the "**literature melting point**"), and compared to the experimental melting point.

Care must be taken to refrain from jumping to conclusions about the identity of a compound based *solely* on a melting point. Millions of solid organic compounds exist, and most have melting points below 250°C. It is not uncommon for two different compounds to have coincidentally similar or identical melting points. Therefore, a melting point should be used as simply one piece of data to *support* the identification of an unknown.

Although coincidentally similar melting points are not unheard of, when used in the context of assessing the product of a chemical reaction, melting points can be a powerful identification tool. For example, three possible products of the nitration of benzaldehyde are 2, 3, or 4-nitrobenzaldehyde. Since these products have very different melting points, the melting point of the resulting solid (if pure) could be used to strongly suggest which product was formed.



Benzaldehyde | 2-nitrobenzaldehyde mp 42-44 °C | 3-nitrobenzaldehyde mp 55-58 °C | 4-nitrobenzaldehyde mp 103-106 °C

Nitration of benzaldehyde.

Assessing Purity

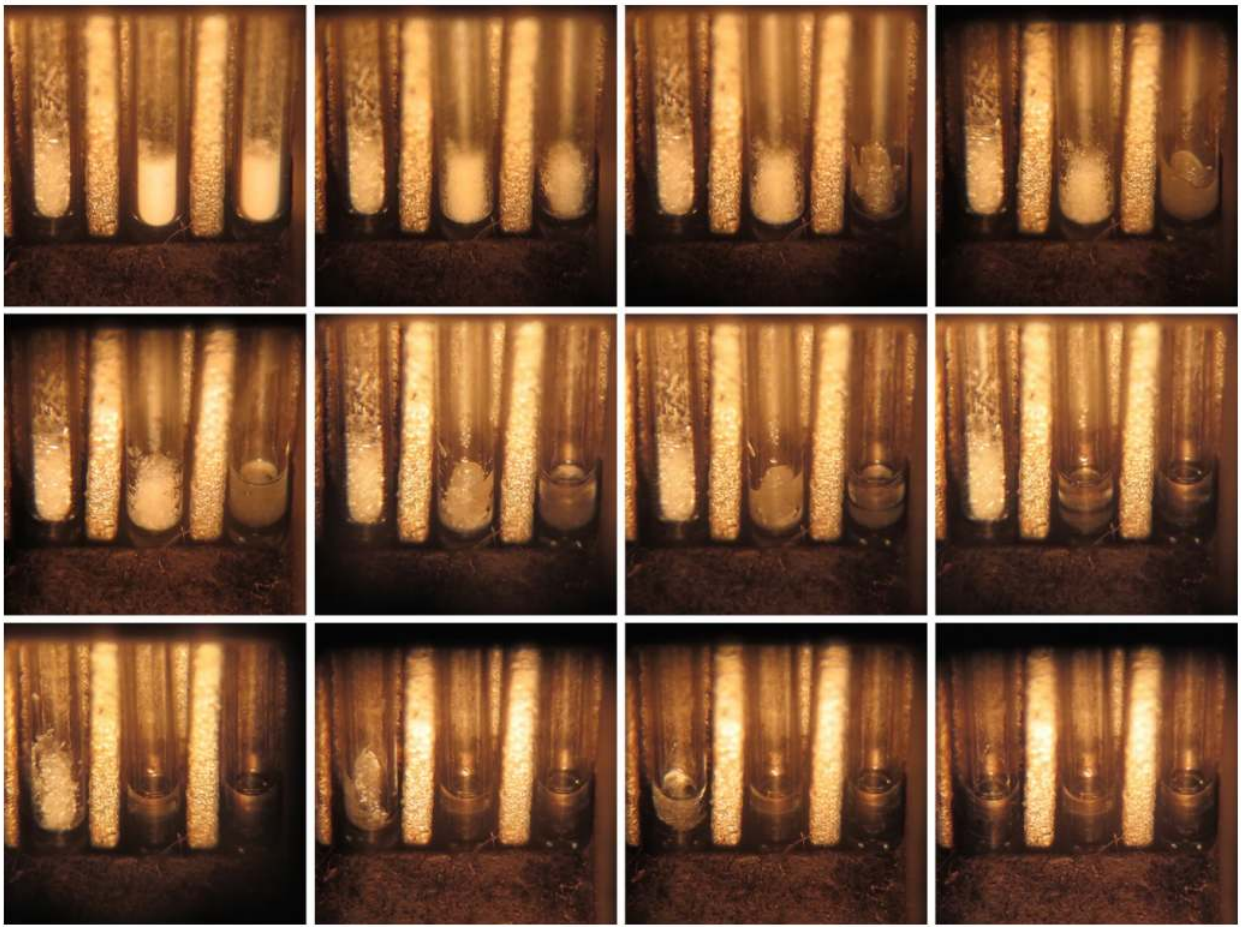
A second reason to determine a compound's melting point is for a rough measure of purity. In general, **impurities lower and broaden the melting range**.

For example, the melting points of samples of benzoic acid contaminated with known quantities of acetanilide are summarized in table below. As the quantity of impurity increased, melting began at a lower temperature, and the breadth of the melting range increased.

Melting points of benzoic acid/acetanilide mixtures

mol% Benzoic Acid	mol% Acetanilide	Melting Point (°C)
100%	0%	120 - 122
95%	5%	114 - 121
90%	10%	109 - 120
85%	15%	105 - 117
80%	20%	94 - 116

The series of photos below show the time-lapse melting of three samples side by side in a melting point apparatus: pure benzoic acid (left), benzoic acid with 10 mol% acetanilide impurity (middle), and benzoic acid with 20 mol% acetanilide impurity (right). As the samples are heated, the sample with the greatest impurity (on the right) melts first. Interestingly, both of the impure samples complete melting before the pure sample (on the left) begins to melt.



Time-lapse melting of three samples side by side in a melting point apparatus. Pure benzoic acid (left), benzoic acid with 10 mol% acetanilide (middle), benzoic acid with 20 mol% acetanilide (right)

Adapted from *Uses of Melting Points* by Lisa Nichols.

4.3: Melting Point Determination Procedure

There are a variety of methods by which a sample's melting point can be measured. In our lab we will be using a melting point apparatus called a DigiMelt.

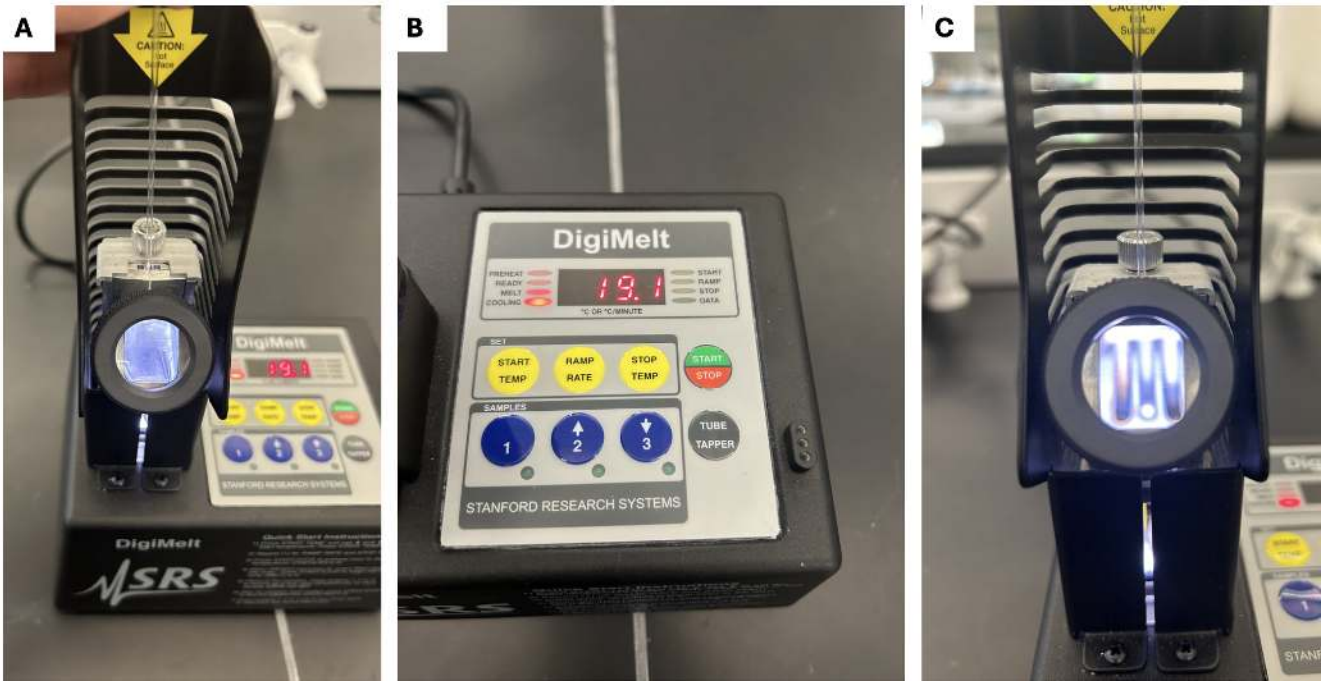
Sample Preparation

1. Obtain a glass capillary melting point tube, which has one end sealed and the other end open. Jab the open end of the tube into a pile of the solid to be analyzed. The solid must be dry or the results will be affected as solvent/water can act as an impurity and affect the melting range.
2. Invert the capillary tube and gently tap the tube on the benchtop to cause the solid to fall to the closed end. Then, drop the capillary tube closed side down several times through a long narrow tube. The capillary tube will bounce as it hits the benchtop, and pack the solid into the bottom of the tube. Failure to pack the solid well may cause it to shrink when heating, which can cause confusion as to the correct melting temperature.
3. If needed, repeat the previous steps to load sample until it is a height of 2-3 mm in the tube. It is important that the sample be no higher than 3 mm or the melting range will be artificially broad.

a) Depositing sample into the open end of a capillary tube, b) Inverting and tapping the tube on the benchtop, c) Dropping the sample through a long tube, d) Correct height of sample in the tube

Melting Point Apparatus

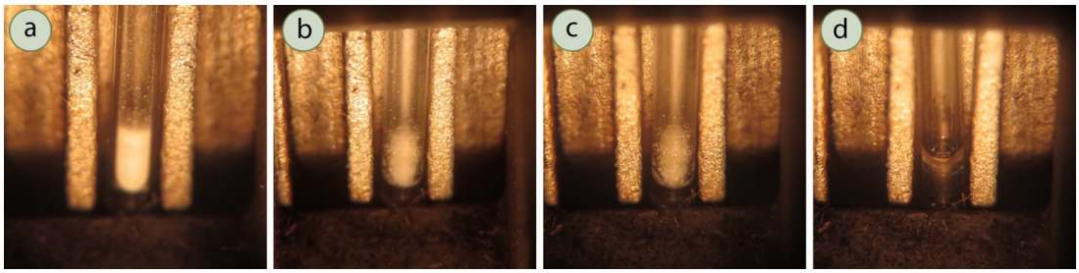
1. Turn on the apparatus and adjust the setting to an appropriate heating rate.
   - It is very important to adjust your starting temperature as you are going to heat at a slower rate to hone in on your melting point range. The starting temperature should be 15 °C below the start of melting point range and your ramp rate should be 1 °C/minute.
   - The end temperature can be set to 15 °C above the literature melting point.
2. Insert the capillary tube containing the sample into a slot behind the viewfinder of a melting point apparatus. There are three slots in each apparatus, and multiple melting points can be taken simultaneously after gaining experience with the technique.
3. Look through the viewfinder to see a magnified view of the sample in the apparatus, which should be illuminated.



a) Insertion of capillary sample into the melting point apparatus, b) Control for heating rate, start and end temperatures, c) Monitoring of the sample through the viewfinder

Recording the Melting Point

1. The solid may be approaching its melting point if the solid is seen pulling away from the walls of the tube to form a cone of solid, which is called "**sintering**." Melting will normally occur within a few degrees of this point. The solid may also shrink or compact before melting.
2. Record the first temperature of the melting range with the appearance of the first visible drop of liquid. At first it may seem as if the sides of the solid glisten, and the temperature should be recorded when a droplet is seen on the side or bottom of the tube (a hint of movement will be noticed in the tube).
3. Record the second temperature of the melting range when the entire sample has just melted, which occurs when all portions of the opaque solid have turned to a transparent liquid.
4. If another melting point trial is to be performed directly after the first, the metal block should be rapidly cooled to at least 30ºC below the next melting point.
5. Although unusual, the sample may begin to darken, which indicates decomposition is occurring before the sample is melting. Take note of the decomposition temperature, as it is sometimes as reliable a reference point as a compound's melting point. Use the letter "d" after a melting point to indicate decomposition (e.g. 251ºC d).



Melting of benzoic acid: a) Well below the melting point, b) Glistening, c) First liquid droplet seen, d) Sample is completely melted.

Adapted from *Step-by-step Procedures for Melting Point Determination* by Lisa Nichols.

CHAPTER OVERVIEW

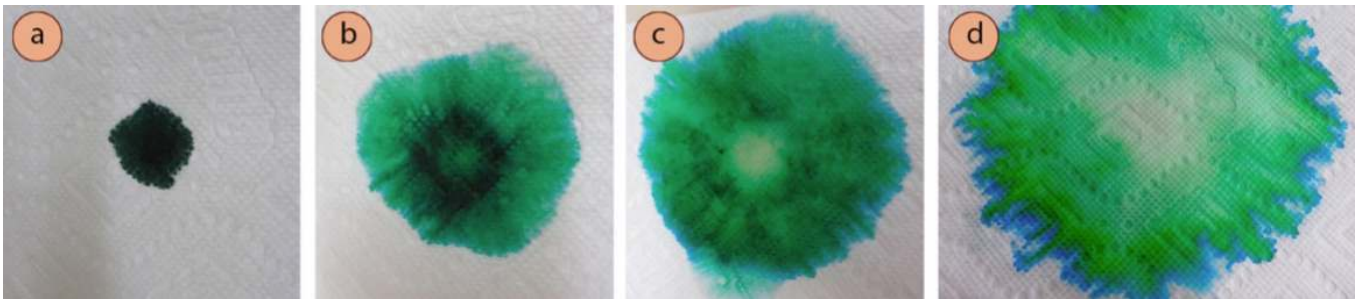
5: Thin Layer Chromatography

5.1: Overview of Chromatography

Chromatography is a technique used to separate the components of a mixture. It can be used as an analytical technique to gain information about what is present in a mixture, or as a purification technique to separate and collect the components of a mixture. Chromatography in the organic chemistry laboratory can be classified into several broad categories. These techniques follow the same general principles in terms of how they are able to separate mixtures, and so will be discussed collectively in this section.

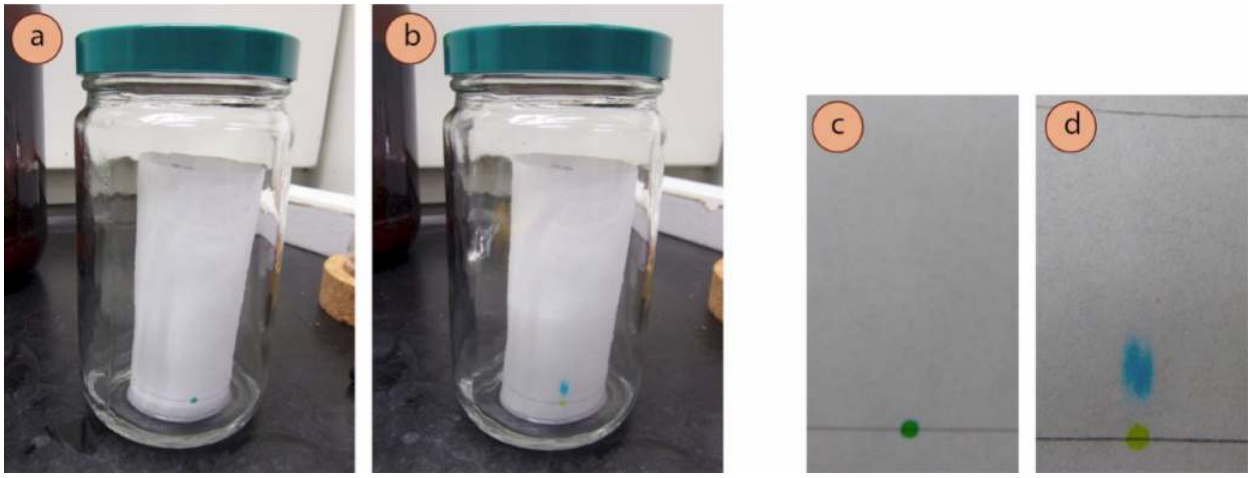
		
Thin Layer Chromatography (TLC) analytical method for testing mixtures	**Column Chromatography** purification method for separating and collecting components	**Gas Chromatography** instrumental technique for analyzing mixtures

The general idea of chromatography can be demonstrated with food dyes in your kitchen. Commercial green food dye does not contain any green colored components at all, and chromatography can show that green food dye is actually a mixture of blue and yellow dyes. If a drop of green food dye is placed in the middle of a paper towel followed by a few drops of water, the water will creep outwards as it wets the paper. As the water expands, the dye will travel with it. If you let the dye expand long enough, you'll see that the edges will be tinted with blue. This is the beginning of the separation of the blue and yellow components in the green dye by the paper and water.



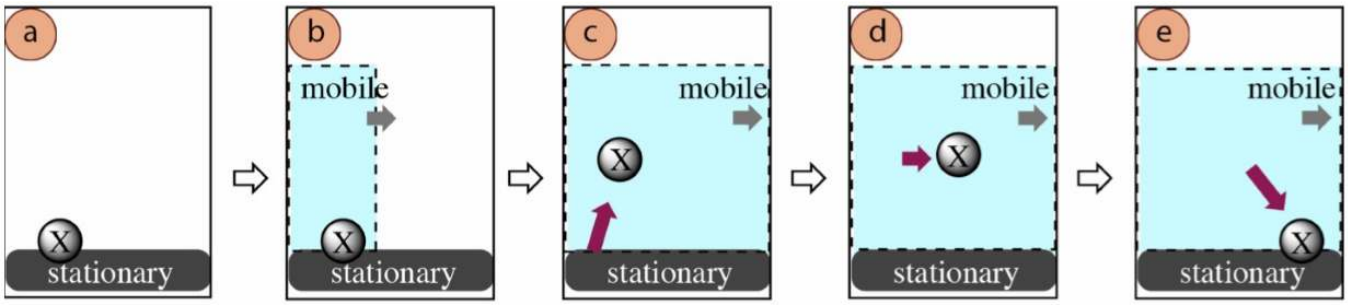
Time-lapse sequence of a drop of green food dye on a paper towel, followed by a drop of water.

A complete separation of the green food dye can be accomplished using paper chromatography. A dilute sample is deposited on the bottom edge of a piece of paper, the paper is rolled in a cylinder, stapled, and placed vertically in a closed container containing a small amount of solvent. The solvent is allowed to wick up the paper through capillary action (called "**elution**"), and through this method complete separation of the blue and yellow components can be achieved.



Paper chromatography of diluted green food dye: a) Before elution, b) After elution, c) Zoom in before elution, d) Zoom in after elution.

In all chromatographic methods, a sample is first applied onto a stationary material (**stationary phase**) that either **ab**sorbs or **ad**sorbs the sample: adsorption is when molecules or ions in a sample adhere to a surface, while absorption is when the sample particles penetrate into the interior of another material. The sample is then exposed to a liquid or gas traveling in one direction (**mobile phase**). The sample may overcome its intermolecular forces with the stationary surface and transfer into the moving material. The sample will later readsorb to the stationary material, and transition between the two materials in a constant equilibrium.



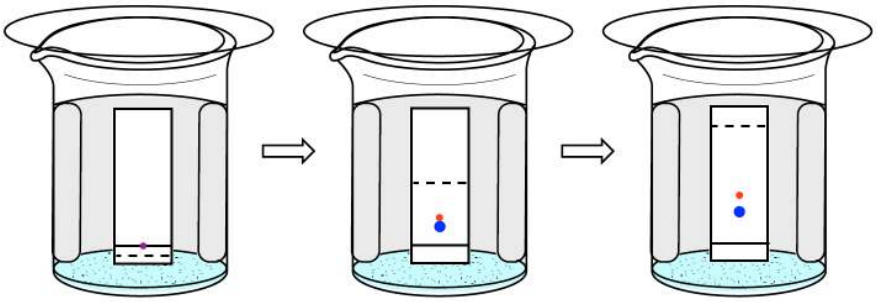
Generic chromatography sequence for compound X: a) Adsorption, b) Exposure to a mobile phase, c) Sample X breaks its attachment to the stationary phase, d) Movement with the mobile phase, e) Reattachment to the stationary phase. Steps c-e are equilibrium steps and constantly repeated.

The main principle that allows chromatography to separate components of a mixture is that components will spend different amounts of time interacting with the stationary and mobile phases. A compound that spends a large amount of time in the mobile phase will move quickly away from its original location, and will separate from a compound that spends a larger amount of time stationary. The main principle that determines the amount of time spent in the phases is the **strength of intermolecular forces** experienced in each phase. If a compound has strong intermolecular forces with the stationary phase it will remain adsorbed for a longer amount of time than a compound that has weaker intermolecular forces. This causes compounds with different strengths of intermolecular forces to move at different rates.

Adapted from *Chromatography Generalities* by Lisa Nichols.

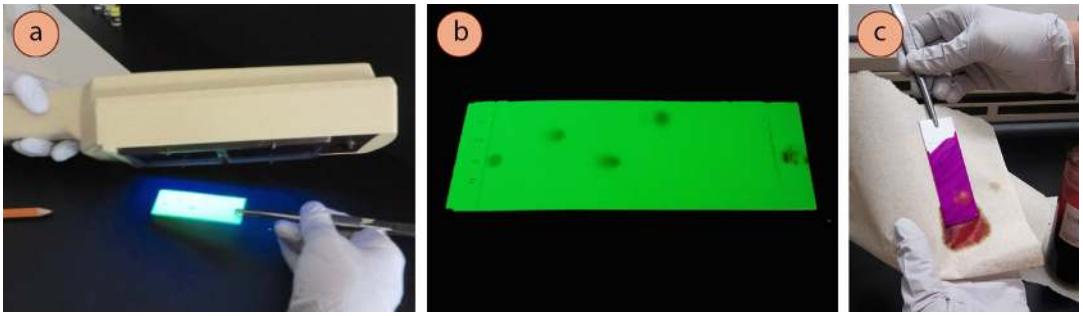
5.2: TLC Overview

Two commonly used stationary phases in chromatography are silica ($SiO_2 \cdot xH_2O$) and alumina ($Al_2O_3 \cdot xH_2O$). Both of these materials are fine white powders that, unlike paper, do not stand up by themselves, so they are deposited in a thin layer on some sort of backing (either glass, aluminum, or plastic). Using these thin layers of stationary phase for separations is called "**thin layer chromatography**" (TLC), and is procedurally performed much the same way as paper chromatography.



TLC sequence to separate the components of a purple sample.

It is very common for organic compounds to appear colorless on the white adsorbent background, which poses the challenge of *seeing* the separation that occurs in a TLC. TLC plates often have to be "visualized," meaning something has to be done to the plate in order to temporarily see the compounds. Common methods of visualization are to use UV light or a chemical stain.



a) Visualization of a TLC plate using UV light, b) Compounds appear dark against a fluorescent green background, c) Use of a chemical stain to visualize a plate.

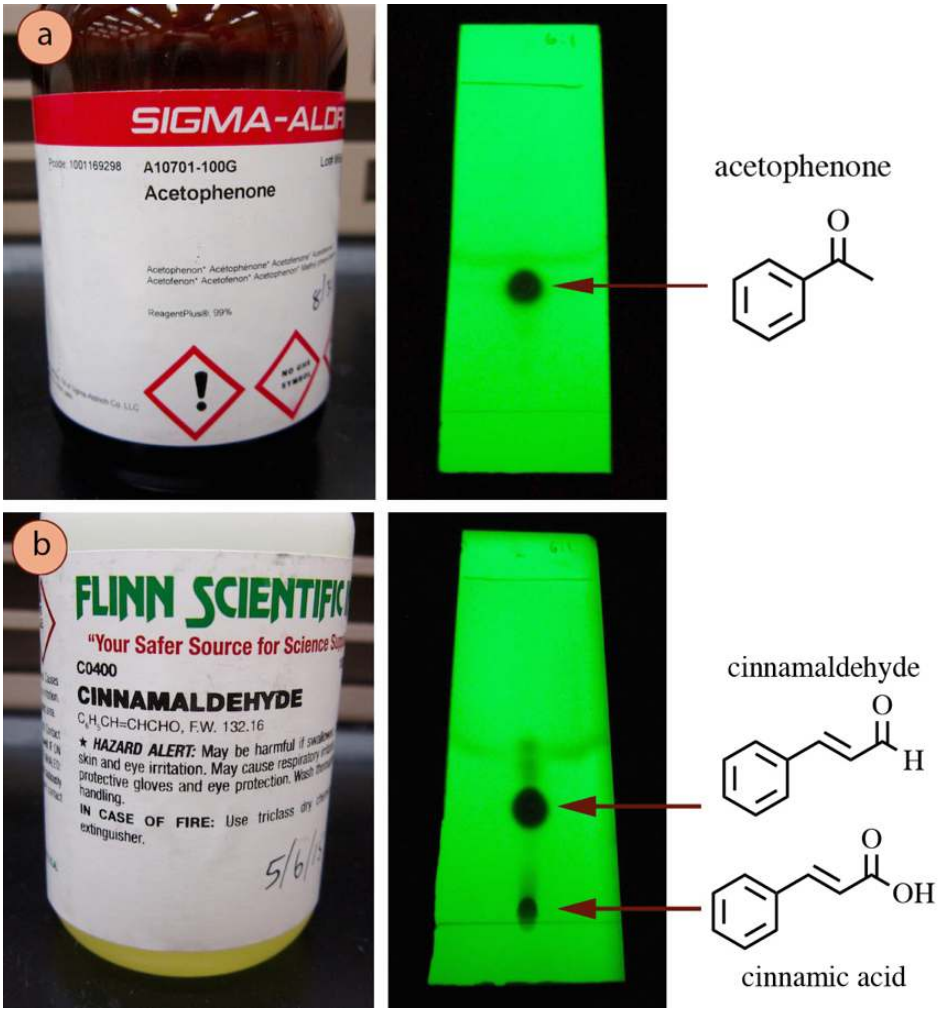
Adapted from *Overview of TLC* by Lisa Nichols.

5.3: TLC Uses

TLC is a common technique in the organic chemistry laboratory because it can give quick and useful information about the purity of a sample and whether or not a reaction in progress is complete. When low polarity solvents are used, a TLC plate can be complete in less than 5 minutes.

Assessing Purity

One of the uses of TLC is to assess the purity of a sample. The photos below show TLC plates of acetophenone and cinnamaldehyde: samples that were diluted from their reagent bottle, run, and visualized with UV light. Acetophenone appeared as only one spot on the TLC plate, indicating the reagent is likely pure. Conversely, cinnamaldehyde is unquestionably impure as its TLC showed two large spots, and had a few fainter spots as well. Aldehydes are prone to air oxidation, and it is common for aldehydes to be found in their reagent bottles alongside with their corresponding carboxylic acids. TLC is one method that can be used to determine how much an aldehyde has degraded.



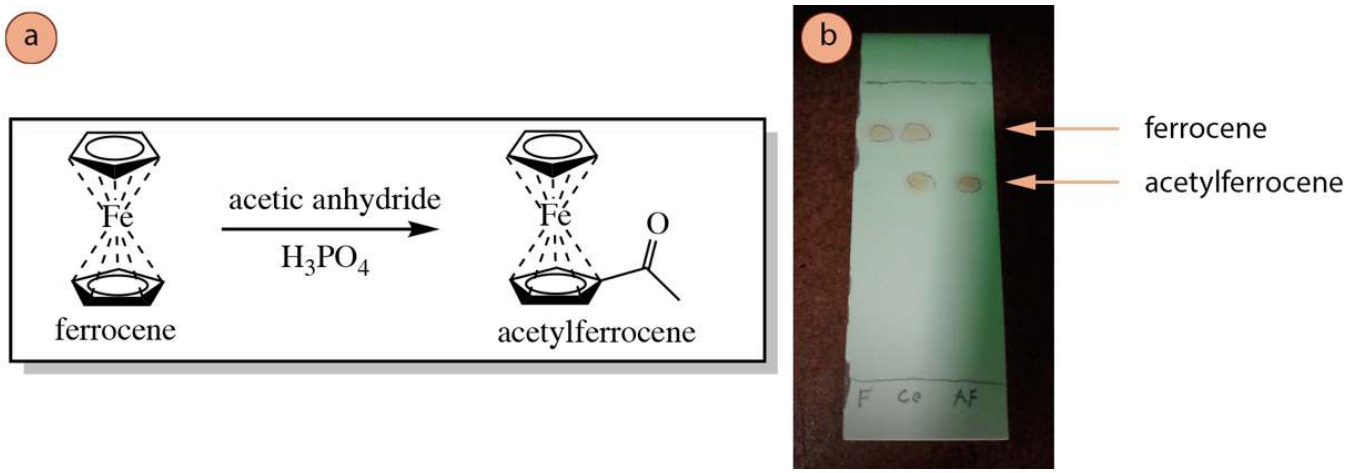
Eluted TLC plates visualized with UV light of: a) Acetophenone, b) Cinnamaldehyde

Assessing Reaction Progress

TLC can be used to analyze a chemical reaction, for example to determine if the reactants have been consumed and a new product has formed. A pure sample of the reactant can be spotted in one lane of a TLC, and the product mixture in another lane. Often the central lane is used for reference, where both reactant and product mixture are spotted over top of one another, in what is called the "**co-spot**."

For example, the reaction shown below is analyzed by TLC. In the first lane (labeled F) is spotted a pure sample of the reactant ferrocene. In the last lane (labeled AF) is spotted the product mixture, which is assumed to be acetylferrocene. In the central lane

(labeled co) is spotted both pure ferrocene and the product mixture. The right-most "AF" lane shows that the reaction appears to be a success: the higher spot of ferrocene is absent (meaning it has been consumed), and a new product spot is present. More tests would have to be done to confirm that the lower spot is the expected product of acetylferrocene, but the TLC results look promising.



a) Reaction scheme of the acetylation of ferrocene, b) TLC plate (visualized with UV light) with 3 lanes: F for pure ferrocene, Co for the co-spot, and AF for the acetylferrocene product.

Monitoring a Reaction by TLC

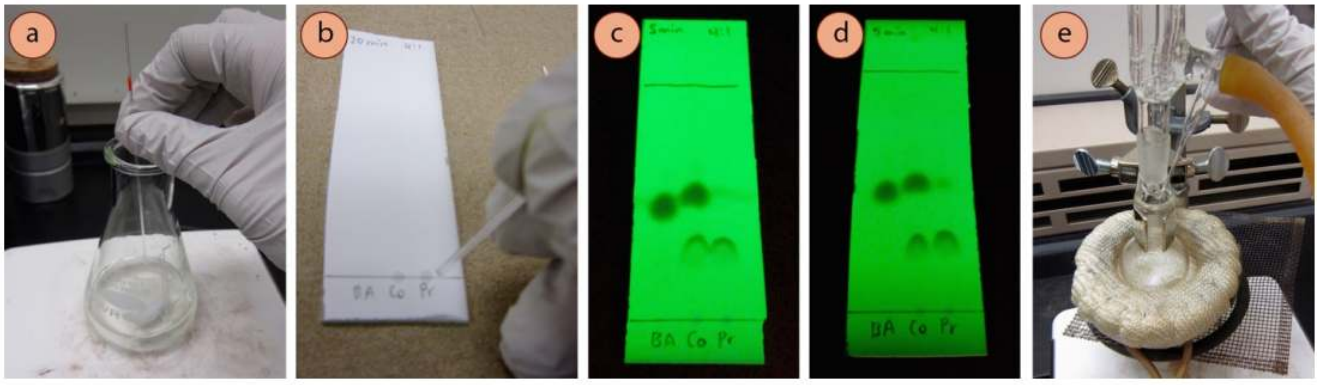
TLC can be used to monitor the progress of a reaction. To use TLC in this manner, three lanes are spotted on a TLC plate: one for the limiting reactant, one for the co-spot, and one for the reaction mixture. The goal is to note the disappearance of the limiting reactant in the reaction mixture lane and the appearance of a new product spot. When the limiting reactant has completely disappeared, the chemist deduces that the reaction is complete, and can then be "worked up".

Obtaining an Aliquot of a Reaction in Progress

To monitor a reaction's progress by TLC, an "**aliquot**" (or tiny sample) of the reaction mixture is necessary.

If the reaction is run at room temperature or with only mild heating, and the concentration of reactants is conducive to TLC, a capillary spotter can be directly inserted into the flask where the reaction is taking place. The aliquot can then be directly spotted on the TLC plate. It is important to fully allow a spot to dry on the TLC plate before placement in the TLC chamber.

To obtain an aliquot of a refluxing solution, briefly remove the condenser and insert a spotter into the reaction mixture. Immediately re-connect the condenser and adjust the clamps while holding the aliquot. Alternatively, lift the flask from the heat source to temporarily cease the reflux before inserting the spotter. The sample may be able to be spotted directly on the TLC plate, but if too concentrated it can be first diluted with acetone in a small vial.



a) Obtaining an "aliquot" from a reaction mixture, b) Spotting the reaction mixture multiple times in the same lane, c) The appearance of the reaction mixture when it was not fully dried before elution, d) When the sample was fully dried before elution, e) Obtaining an aliquot for a refluxing solution.

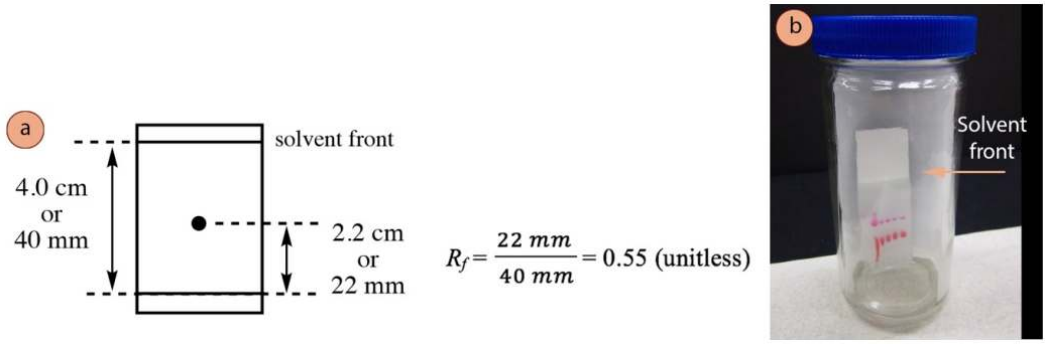
Adapted from *Uses of TLC* by Lisa Nichols.

5.4: Retention Factor

A convenient way for chemists to report the results of a TLC plate in lab notebooks is through a "**retention factor**", or **Rf value**, which quantitates a compound's movement.

Rf = distance traveled by the compound / distance traveled by the solvent front

To measure how far a compound traveled, the distance is measured from the compound's original location (the baseline marked with pencil) to the compound's location after elution (the approximate middle of the spot). To measure how far the solvent traveled, the distance is measured from the baseline to the solvent front. The solvent front is essential to this Rf calculation. When removing a TLC plate from its chamber, the solvent front needs to be marked *immediately* with pencil, as the solvent will often evaporate rapidly.



a) Sample Rf calculation, b) Appearance of the solvent front on an eluting TLC plate.

Since an Rf is essentially a percentage, it is not particularly important to let a TLC run to any particular height on the TLC plate. Slight variations in Rf arise from error associated with ruler measurements, but also different quantities of adsorbed water on the TLC plates that alter the properties of the adsorbent. Rf values should always be regarded as approximate.

If we analyze two compounds and they give approximately the same Rf with the same solvent system, the two compounds are likely the same.

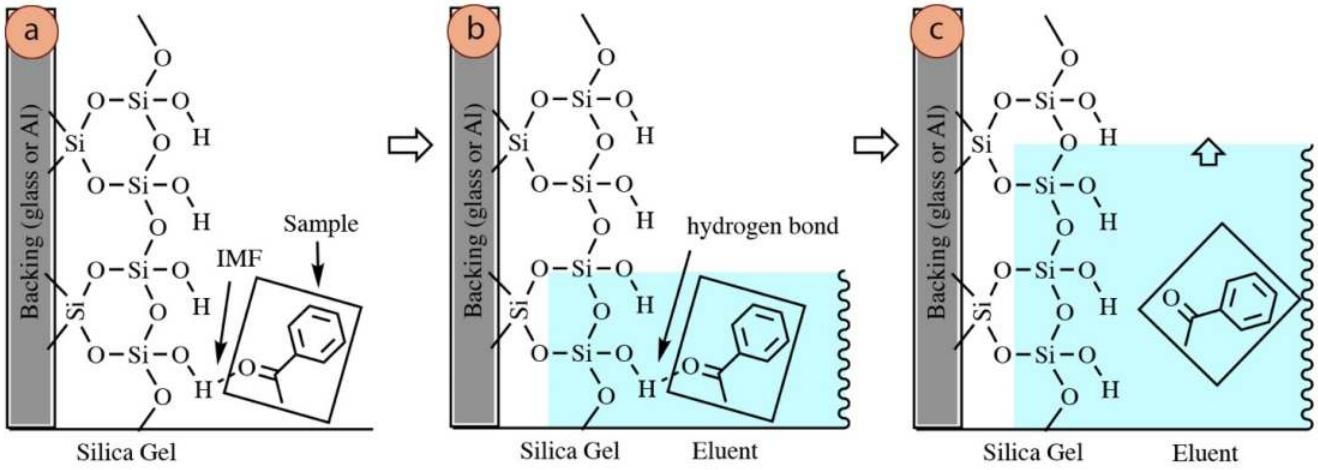
Although in theory a TLC can be run to any height, it's customary to let the solvent run approximately 0.5 cm from the top of the plate to minimize error in the Rf calculations, and to achieve the best separation of mixtures. A TLC plate should not be allowed to run completely to the top of the plate.

Adapted from *The Retention Factor* by Lisa Nichols.

5.5: Separation Theory

Silica gel is composed of a network of silicon-oxygen bonds, with O−H bonds on its surface, as well as a layer of water molecules. Silica gel ($SiO_2 \cdot xH_2O$) is used in this discussion, but is structurally analogous to alumina ($Al_2O_3 \cdot xH_2O$). This very polar stationary phase is paired with a relatively nonpolar mobile phase (an organic solvent or solution), in what is referred to as "normal phase" TLC. Although this is the most common form of TLC (and what will be focused on in this section), "reverse phase" TLC (with a nonpolar stationary phase and a polar mobile phase) is sometimes used.

The figure below shows how acetophenone would cling to the surface of silica gel through intermolecular forces (IMFs). In this case, acetophenone can hydrogen bond to the silica surface through its oxygen atom. As eluent flows over the sample, an equilibrium is established between the sample being adsorbed on the stationary phase and dissolved in the mobile phase. When in the mobile phase, the compound moves up the plate with the flow of liquid to later readsorb on the stationary phase further up the plate. The resulting Rf of the compound is dependent on the amount of time spent in the stationary and mobile phases.



Structural diagrams of compounds bound to a silica coated TLC plate: a) Acetophenone spotted on the baseline of the TLC plate, b) Eluent crawling up the TLC plate, c) Acetophenone in the mobile phase after breaking its IMFs with the silica surface.

The equilibrium distribution between the two phases depends on several factors:

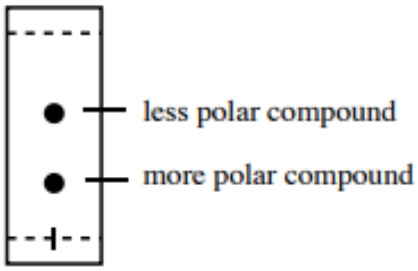
The strength of intermolecular forces between the sample and the **stationary phase:**

> A compound that forms strong IMFs with the silica or alumina will often favor the stationary phase, and will spend much of the elution time adhered to the plate. This means it will spend less time in the mobile phase (which is the only means for it to travel up the plate), causing it to end up low on the TLC plate, and have a low Rf.
>
> Compounds that have oxygen or nitrogen atoms should be able to hydrogen bond with the stationary phase (have strong IMFs with the stationary phase), and thus will have lower Rf values than compounds of similar size that can only interact through dispersion forces.

The strength of interaction between the sample and the **mobile phase:**

> As the mobile phase is always less polar than the stationary phase in normal phase TLC, polar compounds will tend to have a lesser affinity for the mobile phase than nonpolar compounds (based on the "like dissolves like" principle). Therefore, polar compounds tend to spend less of the elution time mobile than a nonpolar compound, so will travel "slower" up the plate, and have a low Rf.

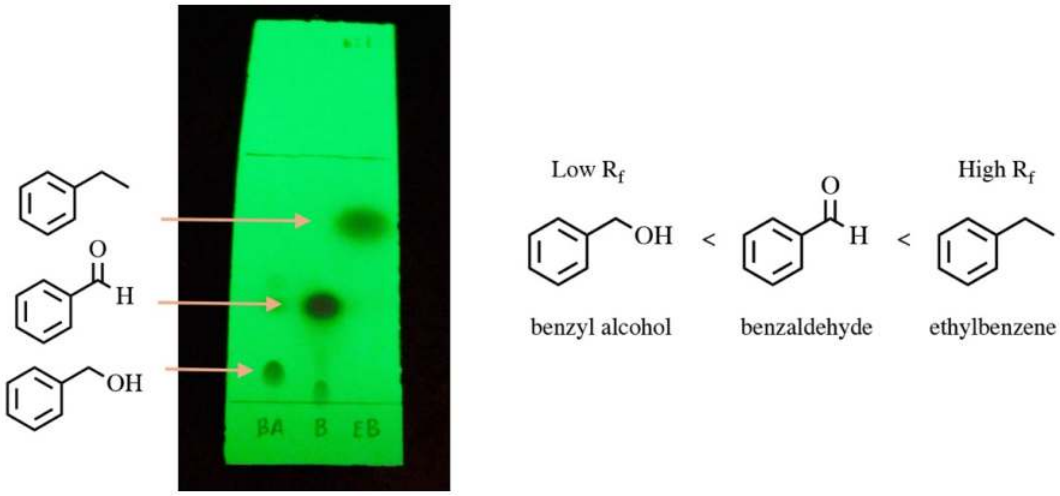
Thus, a compound with a lower Rf tends to have more polar functional groups than a compound with a higher Rf.



Relation of polarity to Rf

Structural Considerations

To demonstrate the effect of structural features on Rf, an eluted TLC plate of benzyl alcohol, benzaldehyde, and ethylbenzene is shown in Figure 2.18. The relative order of Rf reflects the polarity trend in the series.



TLC of benzyl alcohol (lane 1), benzaldehyde (lane 2), and ethylbenzene (lane 3)

Benzyl alcohol and benzaldehyde have polar functional groups so they had lower Rf values than ethylbenzene, which is completely nonpolar. Both compounds are able to hydrogen bond to the polar stationary phase, so are more strongly attracted to the stationary phase than ethylbenzene, which interacts only through weak dispersion forces. As the least "polar" of the series, ethylbenzene is also the best dissolved by the weakly polar eluent. For these reasons, ethylbenzene spent the least time in the stationary phase and the most time in the mobile phase, which is why it traveled the furthest up the plate and had the highest Rf of the series.

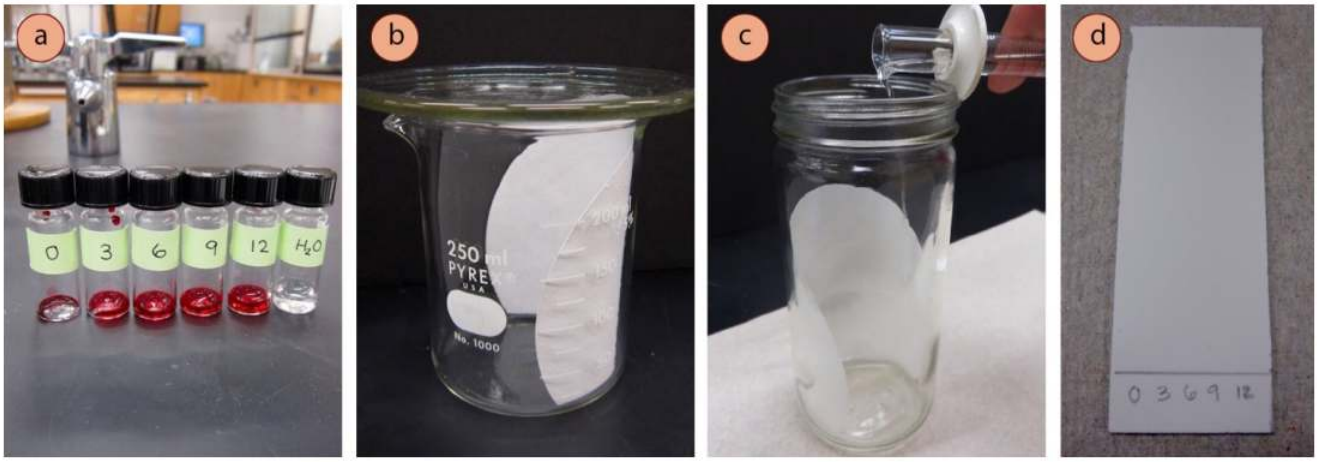
Both benzaldehyde and benzyl alcohol are capable of hydrogen bonding with the stationary phase, but benzyl alcohol had the lower Rf because it can form *more* hydrogen bonds (through both the oxygen and hydrogen atoms of the OH group). This caused benzyl alcohol to be more strongly adhered to the silica/alumina than benzaldehyde, causing it to spend more time in the stationary phase.

Adapted from *Separation Theory* by Lisa Nichols.

5.6: TLC Procedure

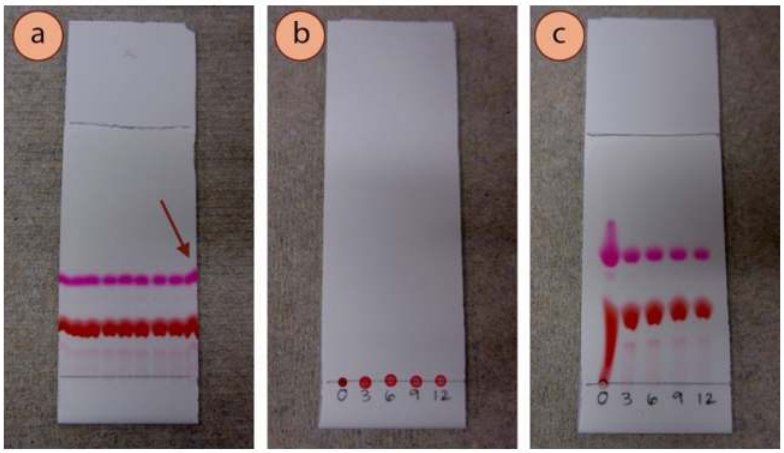
Prepare the Samples, TLC Chamber, and Plate

1. Dissolve a small amount of solid a small vial or test tube using acetone.
2. Obtain a TLC chamber with lid. Cut a piece of filter paper so that when placed in the chamber, the filter paper fits inside the chamber and is flat on the bottom but not obscuring your view of the inside. The filter paper keeps the chamber saturated with vapors so when the eluent rises on the plate it doesn't easily evaporate..
3. Add a portion of a prepared solvent for chromatography (5-10 mL for this type of TLC chamber). Close the lid and tilt the chamber to wet the filter paper.
4. Obtain a TLC plate, touching the plate only on the back or edges, but not on the white surface. Use a rule to lightly draw a straight line with a **pencil** roughly 1 cm from the bottom. Do not use pen as the ink will separate along with your samples and the plate cannot be used.



a) Prepared samples, b) TLC chamber made with a beaker and watch glass, c) Adding eluent to a TLC chamber, d) Prepared plate.

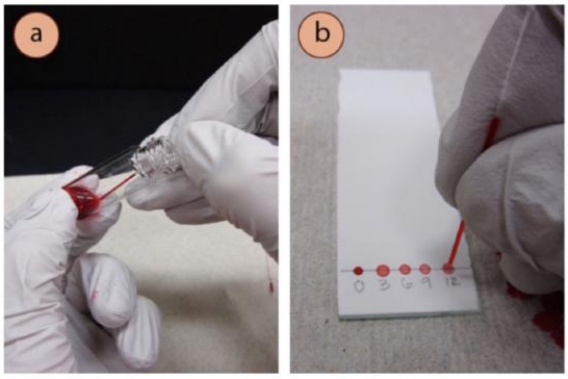
5. Label the areas with pencil where you plan to place the samples. The lanes should not be placed too close to the edge (keep at least 5mm away from each edge), as it is not unusual for solvent to travel slightly "fast" at the edge where capillary action of the solvent is greater. You should be able to place 3-5 spots on each plate. The lanes should also not be placed too close to one another, or the spots may overlap after elution. Spot broadening means that samples deposited right next to each other on the baseline of a TLC plate will probably bleed together during elution.



a) Red food dye sample spotted many times, with arrow indicating the creep of sample on the edges, b) before elution, c) after elution

Spot the TLC plate with sample

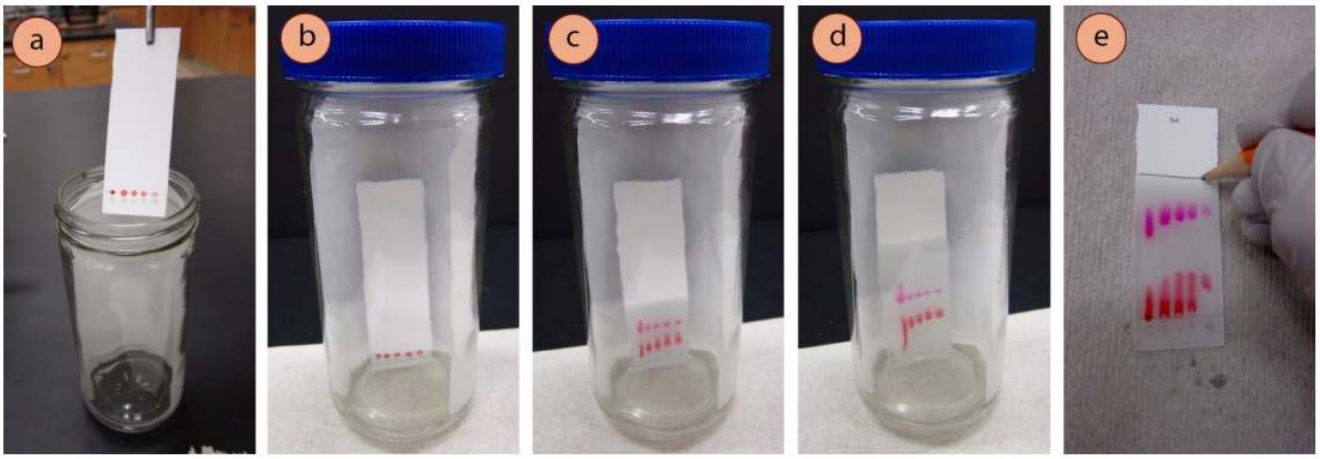
6. Obtain a capillary spotter (a very thin hollow piece of glass open at both ends). Place your spotter into the diluted sample you want to analyze to withdraw liquid into the spotter through capillary action.
7. Keeping the spotter mostly vertical, make a practice "spot" on a paper towel or scrap piece of silica or alumina to familiarize yourself with how the liquid delivers from the spotter. The spots should be very small, around **2 mm** in diameter. Deliver a very small spot of material on the pencil line of the appropriate lane using a quick "up and down" motion with your hand. Don't gouge the silica or alumina with the spotter.
8. Dispose of the spotter in the broken glass container and use a new spotter for each sample.



a) Tipping the vial to withdraw liquid into the spotter, b) Spotting samples onto a TLC plate.

Place the TLC plate in the chamber to "elute"

9. Use forceps to delicately place the TLC plate into the chamber. Be sure that the plate touches the solvent perpendicular to the solvent, and not at an angle. This can cause the solvent to travel up the plate unevenly. The liquid level must be below the pencil line where the samples are spotted or the compounds will dissolve in the pool of eluent instead of traveling up the plate. Cap the chamber delicately and don't touch it again until the TLC is complete.
10. Allow the TLC to develop. As liquid moves up the TLC plate it will appear transparent and wet. If the eluent is very polar (e.g. contains large amounts of ethanol or water), elution will take a relatively long time (can be 30-40 minutes). If the eluent is very nonpolar (e.g. contains large amounts of hexane or petroleum ether), elution will be relatively quick (can be 2-5 minutes for a 10 cm tall plate).
11. It is best to let the TLC run to around 0.5cm from the top of the plate to get the best separation of spots and to minimize error in Rf calculations.



a) Using forceps to place the TLC plate into the chamber, b-d) Elution, e) Marking the solvent line with a pencil.

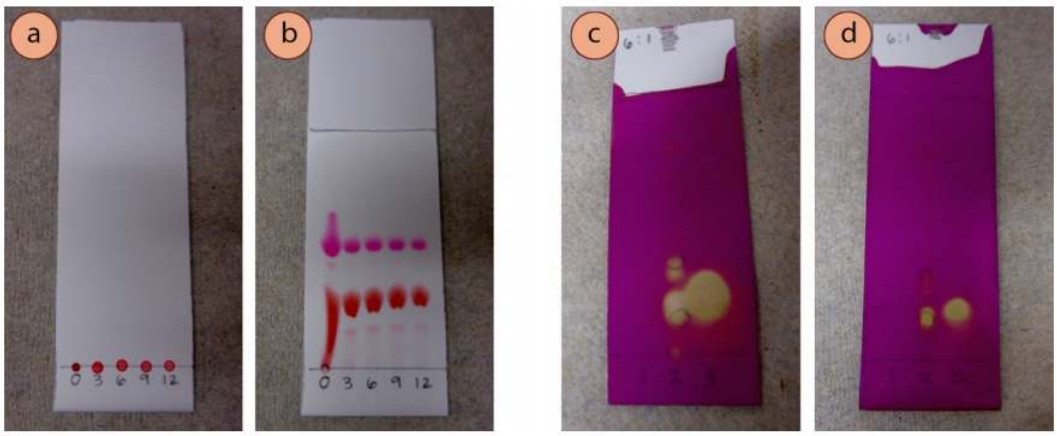
Remove the TLC plate from the chamber

12. Open the TLC chamber, and remove the TLC plate with forceps. *Immediately* mark the solvent line with a pencil to enable an Rf calculation, as the solvent often readily evaporates.
13. If the compounds on the TLC plate are colored, the process is complete. If the compounds are colorless, they need to next be visualized. Visualization techniques are described on the next page.

TLC Troubleshooting

The spots are streaky or "blobby"

The components of a sample can appear as long streaks or "blobby" spots on a TLC plate if the samples are run at too high a concentration. The TLC plate can become overload and proper equilibration between stationary and mobile phases does not occur. Running TLC on concentrated samples gives inaccurate Rf values and may hide multiple spots. If streaking is seen on a TLC plate, the sample should be diluted and run again.

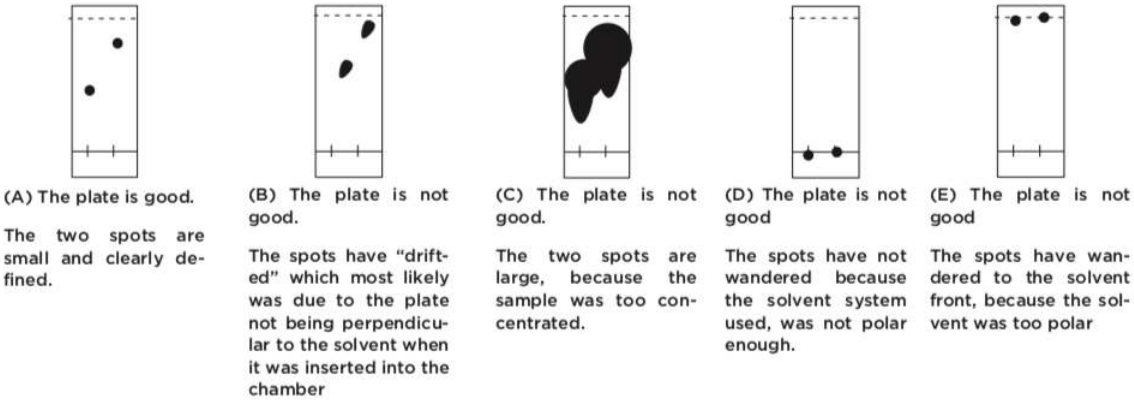


a) Before elution of red food dye samples at different aqueous dilutions, b) After elution, c) Samples too concentrated, d) Samples at proper dilution.

The spots ran unevenly

At times the solvent front may run unevenly on a TLC plate. This may occur if the plate was placed in the chamber at a slight tilt, if eluent splashed onto the plate during placement in the chamber, or if the chamber was jostled during elution. In cases where the front is dramatically different from one position to the other, the front should be measured for each *lane* of the plate (if calculating an Rf) instead of only once, or the plate should be remade and run again.

TLC analyses require a fair amount of care to get right. The most likely problems can be traced to the application and preparation of the TLC plate. Let us have a closer look at some TLC plates that are of inadequate quality and their causes.



(A) The plate is good.

The two spots are small and clearly defined.

(B) The plate is not good.

The spots have "drifted" which most likely was due to the plate not being perpendicular to the solvent when it was inserted into the chamber

(C) The plate is not good.

The two spots are large, because the sample was too concentrated.

(D) The plate is not good

The spots have not wandered because the solvent system used, was not polar enough.

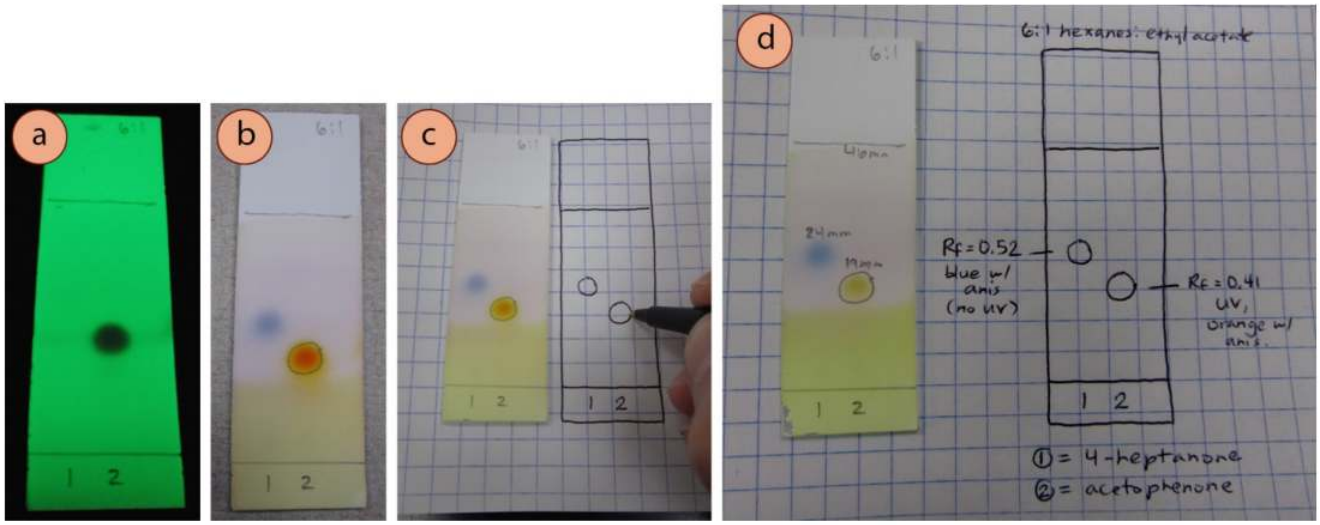
(E) The plate is not good

The spots have wandered to the solvent front, because the solvent was too polar

Figure 2.9. Various TLC scenarios.

Notebook Record of TLC

It is important to copy a TLC plate "to scale", meaning the dimensions should be the same in the notebook as they are in actuality. It is important to copy the TLC plate as accurately as possible, drawing the spots exactly as they appear, even if they are streaky or blobby. All spots seen in a lane should be recorded, even if they are faint. Good record keeping means to record all observations, even if the importance is unknown; a faint, unexpected spot may become relevant at a later time.



Record TLC plate into a lab notebook: a) Plate under UV light, b) Plate stained with p-anisaldehyde, as it appeared directly after heating, c) Copying the TLC into the notebook, d) Final record in the notebook.

Several other notations should be made along with the sketch of the TLC pate. The solvent system and identity of what was spotted in each lane must be recorded. For each spot, an Rf should be calculated along with notation of UV activity. If a spot changes appearance over time, as the orange spot in Figure 2.29b faded to light green over time, the initial appearance should be recorded.

Adapted from *Step-by-Step Procedures for Thin Layer Chromatography* by Lisa Nichols.

5.7: Visualizing TLC Plates

Organic compounds most commonly appear colorless on the white background of a TLC plate, which means that after running a TLC, chemists often cannot simply *see* where compounds are located. The compounds have to be "visualized" after elution, which means to temporarily convert them into something visible.

Ultraviolet Absorption

The most common non-destructive visualization method for TLC plates is ultraviolet (UV) light. A UV lamp can be used to shine either short-waved (254 nm) or long-waved (365 nm) ultraviolet light on a TLC plate with the touch of a button. Most commercially bought TLC plates contain a fluorescent material (e.g. zinc sulfide) in the silica or alumina, so the background of the plate will appear green when viewing with short-waved UV light. If a compound absorbs 254 nm UV light, it will appear dark, as the compound prevents the fluorescent material from receiving the UV light.

This method is so quick and easy that it is often the first visualization method tried. It is most useful for visualizing aromatic compounds and highly conjugated systems, as these strongly absorb UV. Since the compounds remain unchanged after viewing with UV light, a further visualization technique can be used afterwards on the same plate.

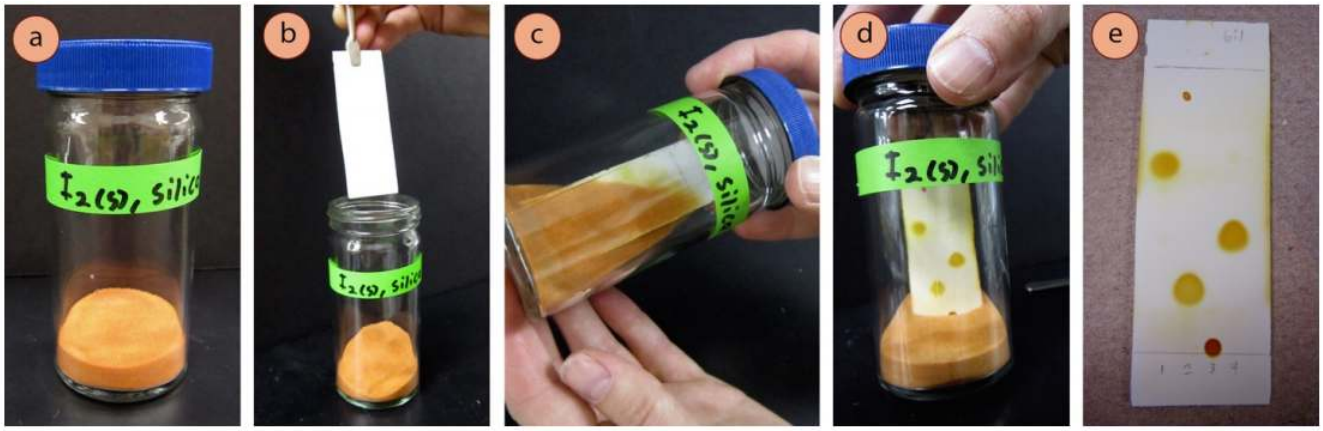
a) Tilting a UV lamp to visualize a TLC plate, b) Box to protect eyes from UV damage, c) Appearance under UV.

Procedure for UV visualization of TLC plate:

1. Use a UV lamp to look at your developed TLC plate by pressing the short-waved button. **Safety note:** Care should be taken to never look directly at the UV source, and to minimize exposure to eyes.
2. The plate background will appear green under short-waved UV light, and UV-active compounds will appear dark. Use a pencil to lightly circle spots, as they will disappear when the UV light is removed.
3. Some compounds themselves fluoresce, appearing a variety of colors when exposed to either short- or long-waved UV light (bright purple or blue is the most common). Record these types of observations in your notebook if you see them, as they are rare, and are therefore an excellent identification tool.

Iodine

A commonly used semi-destructive visualization method is to expose a developed TLC plate to iodine (I_2) vapor. An "iodine chamber" can be created by adding a few iodine crystals to a TLC chamber, or by adding a few iodine crystals to a chamber containing a portion of powdered silica or alumina. When a developed TLC plate is placed in the chamber and capped, the iodine sublimes and reacts with the compounds on the plate, forming yellow-brown spots. The coloration occurs because iodine forms colored complexes with many organic compounds.



a) An iodine chamber using silica gel, b-d) Inserting a plate into the chamber and jostling, e) Developed TLC plate with iodine.

Procedure for visualization of TLC plate with iodine:

1. If not already prepared, make an "iodine chamber": in a fume hood, place a few centimeters of powdered silica or alumina in a screw-capped TLC chamber and add a few crystals of solid iodine (**safety note:** silica and alumina are lung irritants, and iodine vapor is considered an irritant and toxic). Let the silica or alumina and iodine sit together for a while with periodic swirling, and eventually the powder will become orange from adsorbing the iodine vapor.
2. In a fume hood, place the developed TLC plate in the iodine chamber with forceps and close the lid. Gently shake the chamber to bury the TLC plate in the iodine-stained silica or alumina until the spots become colored. Many spots will appear yellow-brown almost immediately, and may darken with extended time.
3. Promptly record appropriate observations of the TLC in your notebook, or circle the spots with a pencil, as the colors will soon fade as the iodine evaporates from the plate.

Chemical Stains

There are a variety of destructive visualization methods that can turn colorless compounds on a TLC plate into colored spots. A plate is either sprayed with or dipped in a reagent that undergoes a chemical reaction with a compound on the TLC plate to convert it into a colored compound, enabling the spot to be seen with the naked eye. Since a chemical reaction is occurring in the process, it is common to gently heat a plate after exposure to the reagent to speed up the reaction, although this may be unnecessary with some stains. Not every compound can be visualized with every reagent if they do not react together, and stains are often designed to work with only certain functional groups. The specific stain should be chosen based on the presumed structure of the compounds you want to visualize.

Adapted from *Visualizing the Plates* by Lisa Nichols.

CHAPTER OVERVIEW

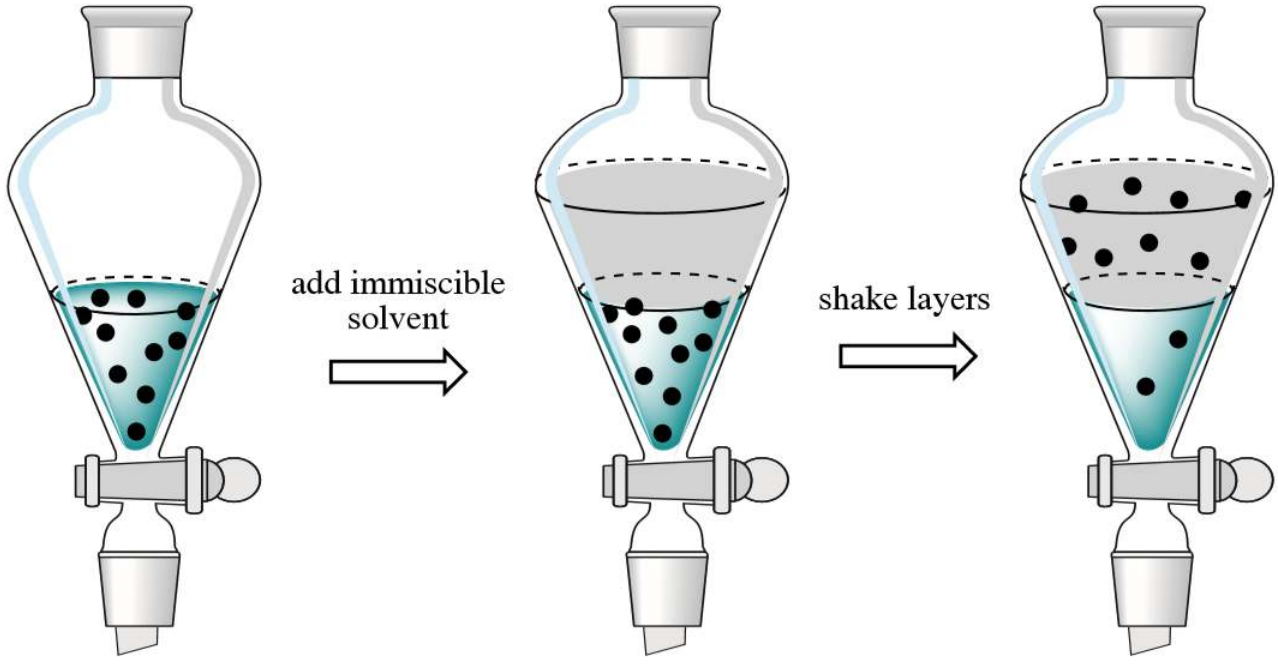
6: Extraction

6.1: Extraction Overview

"Extraction" refers to transference of compound(s) from a solid or liquid into a different solvent or phase. When a tea bag is added to hot water, the compounds responsible for the flavor and color of tea are extracted from the grounds into the water. Decaffeinated coffee is made by using solvents to extract the caffeine out of coffee beans. Bakers use the extract of vanilla, almond, orange, lemon, and peppermint in their dishes, essences that have been extracted from plant materials using alcohol.

Examples of extraction: a) Tea, b) Baking extracts, c) Plant pigments extracted into water droplets after sprinklers hit a fallen leaf on the sidewalk.

In the chemistry lab, it is most common to use liquid-liquid extraction, a process that occurs in a separatory funnel. A solution containing dissolved components is placed in the funnel and an immiscible solvent is added, resulting in two layers that are shaken together. It is most common for one layer to be aqueous and the other an organic solvent. Components are "extracted" when they move from one layer to the other. The shape of the separatory funnel allows for efficient drainage and separation of the two layers.



Schematic of extraction

Compounds move from one liquid to another depending on their **relative solubility in each liquid**. A quick guide to solubility is the "like dissolves like" principle, meaning that nonpolar compounds should be readily extracted into nonpolar solvents (and vice versa). When allowed to equilibrate between two liquids in a separatory funnel, the majority of a compound often ends up in the layer that it is more soluble.

Adapted from *Overview of Extraction* by Lisa Nichols.

6.2: Uses of Extraction

There are several reasons to use extraction in the chemistry lab. It is a principal method for isolating compounds from plant materials. Extraction moves compounds from one liquid to another, so that they can be more easily manipulated or concentrated. It also enables the selective removal of components in a mixture.

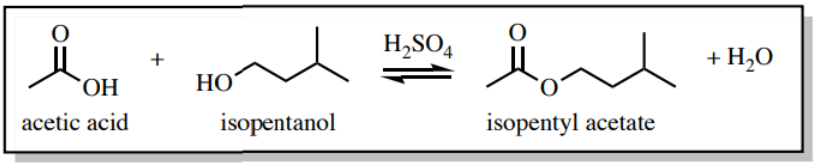
Extracting Natural Compounds

Fruit and plant leaves are primarily composed of cellulose and water, but also contain "essential oils." a greasy mixture of compounds that capture the "essence" of the plant material's smell and taste. Orange oil is roughly 95% limonene, and due to its nonpolar structure, can be extracted from its rind into an organic solvent like hexanes or dichloromethane. The oil can then be concentrated and used to flavor or scent foods, cleaning supplies, and candles.

Selective Removal of Components

When conducting an experiment that synthesizes a chemical product, a reaction is often complete whenever stirring or heating is complete. Directly afterwards more steps are needed to "**work-up**" the reaction in some way. A work-up refers to methods aimed at isolating the product from the reaction mixture, and often begins by using a separatory funnel and extractions.

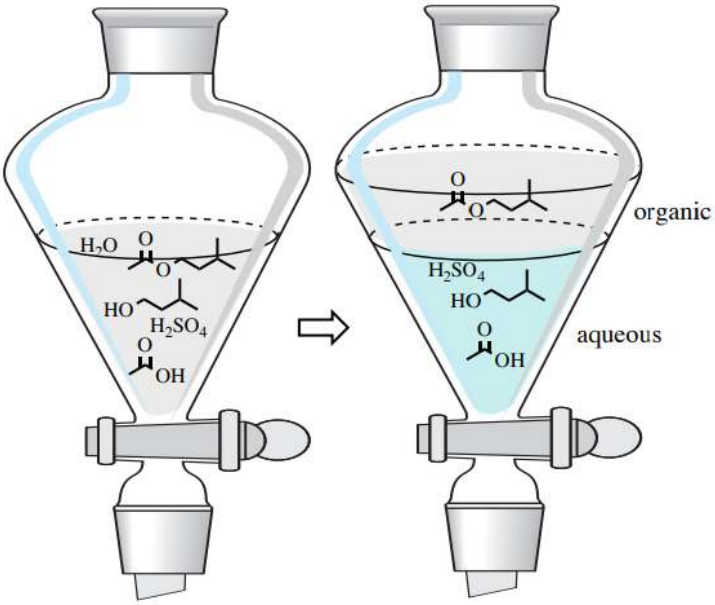
For example, imagine that acetic acid and isopentanol have been heated in the presence of an acid catalyst for one hour in order to make isopentyl acetate, an ester that smells of bananas. After the one-hour time period, there is unfortunately not *just* the banana-smelling ester in the flask. The flask will also contain byproducts, leftover starting materials if the reaction is incomplete, as well as any catalysts used. In this example, there could be *five* compounds in the reaction flask after heating is complete.



Reaction scheme to produce isopentyl acetate.

When "working up" this reaction, the resulting mixture is often poured into a separatory funnel along with some water and organic solvent. This produces two layers in the separatory funnel: an aqueous layer and an organic layer.

After shaking this heterogeneous mixture, the compounds distribute themselves based on their solubility. Compounds that have high water solubility favor the aqueous layer while less polar compounds favor the organic layer. In this example, the acid catalyst and residual carboxylic acid or alcohol would likely be drawn into the water layer. The ester would have a greater affinity for the organic layer than the aqueous layer, causing it to be isolated from the other components in the reaction mixture.



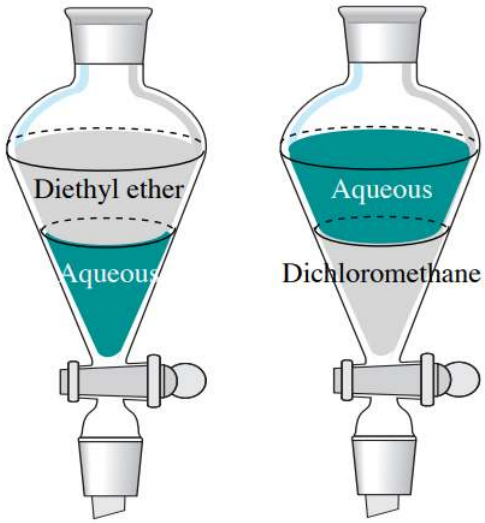
Extraction using water and an organic solvent to isolate isopentyl acetate from the reaction mixture.

Adapted from *Uses of Extraction* by Lisa Nichols.

6.3: Density

It is essential that you know whether the aqueous layer is above or below the organic layer in the separatory funnel, as it dictates which layer is kept and which is eventually discarded. Two immiscible solvents will stack atop one another based on differences in density. The solution with the lower density will rest on top, and the denser solution will rest on the bottom.

Most non-halogenated organic solvents have densities less than 1 g/mL, so will float atop an aqueous solution (if they are immiscible). Halogenated solvents are denser than water (have densities greater than 1 g/mL), and so will instead sink below aqueous solutions.



Relative position of aqueous and organic layers. Most organic solvents like diethyl ether are on top, except for halogenated solvents like dicholoromethane, which are typically on bottom

Density of common solvents at room temperature.

Solvent	Density (g/mL)
Pentane	0.626
Petroleum Ether	0.653
Hexanes	0.655
Diethyl ether	0.713
Ethyl acetate	0.902
Water	0.998
Dicholoromethane	1.33
Chloroform	1.49

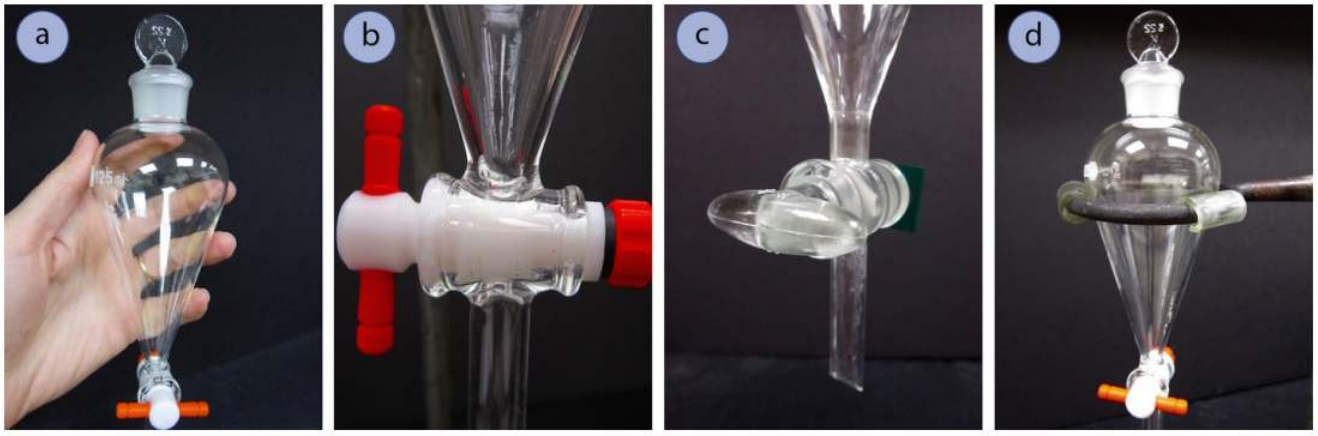
Many solutions used in separatory funnels are fairly dilute, so the density of the solution is approximately the same as the density of the solvent. For example, if mixing diethyl ether and a 10% NaOH(aq) solution in a separatory funnel, knowledge of the exact density of the 10% NaOH solution is not necessary. A 10% NaOH(aq) solution is 90% water (by mass), meaning the density should be fairly close to the density of water (approximately 1 g/mL). The actual density of a 10% NaOH(aq) solution is 1.1089 g/mL, a value only slightly greater than the density of water. The diethyl ether will be the top layer in this situation.

Adapted from *Which Layer is Which?* by Lisa Nichols.

6.4: Extraction Procedure

Prepare the Setup

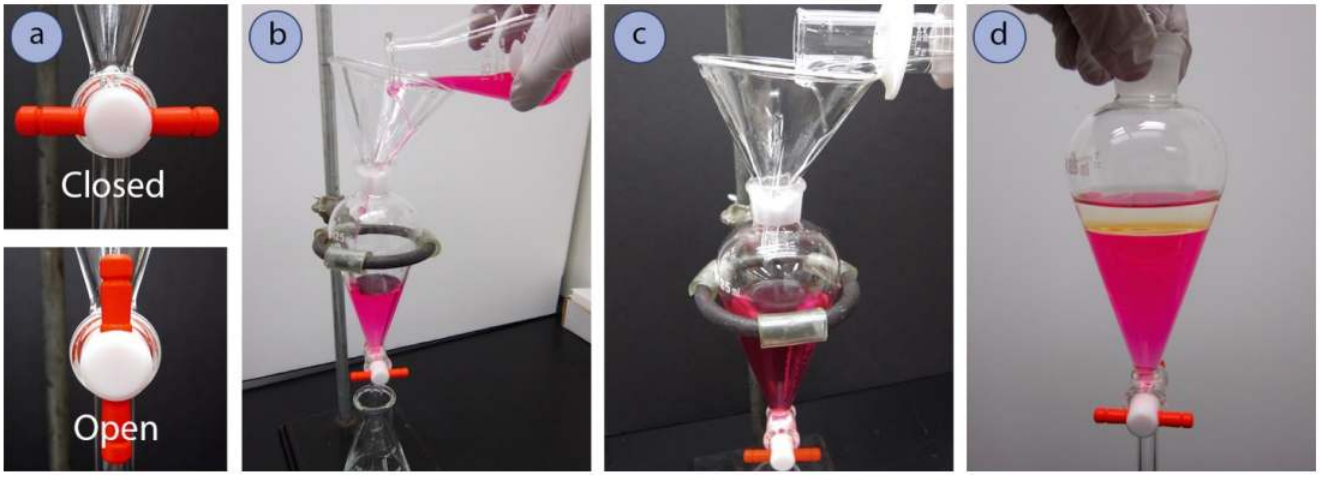
1. Obtain a separatory funnel.
    - A Teflon stopcock is typically used. If it was taken apart to dry, reassemble the stopcock, placing the parts in the appropriate order. Be sure that the Teflon stopcock is moderately tight so that it can still easily turn, but is not so loose that liquid can seep around the joint.
    - Also obtain a stopper that fits well in the top joint of the funnel.
2. Place the separatory funnel in a ring clamp attached to a ring stand or latticework.



a) Separatory funnel, b) Correct order of stopcock components, c) Glass stopcock, d) Funnel in ring clamp

Add the Solutions

3. Before pouring anything into a separatory funnel, be sure that the stopcock is in the "closed" position, where the stopcock is horizontal. As a fail-safe, always position an Erlenmeyer flask beneath the separatory funnel before pouring. This can catch liquid in case the stopcock is accidentally left open, or if the stopcock is loose and liquid leaks through unintentionally.
4. Using a funnel, pour the liquid to be extracted into the separatory funnel. A separatory funnel should never be used with a hot or warm liquid.
5. Pour a quantity of the extractive solvent into the separatory funnel, as indicated by the procedure. It is unnecessary to use precise quantities of solvent for extractions, and the volumes can be measured in a graduated cylinder. If a procedure calls for 20mL of solvent, it is acceptable if between 20-25mL is used.



a) Closed and open stopcocks, b) Pouring in liquid with a funnel: notice the Erlenmeyer flask positioned below as a fail-safe, c) Pouring in the organic solvent, d) Separatory funnel before mixing

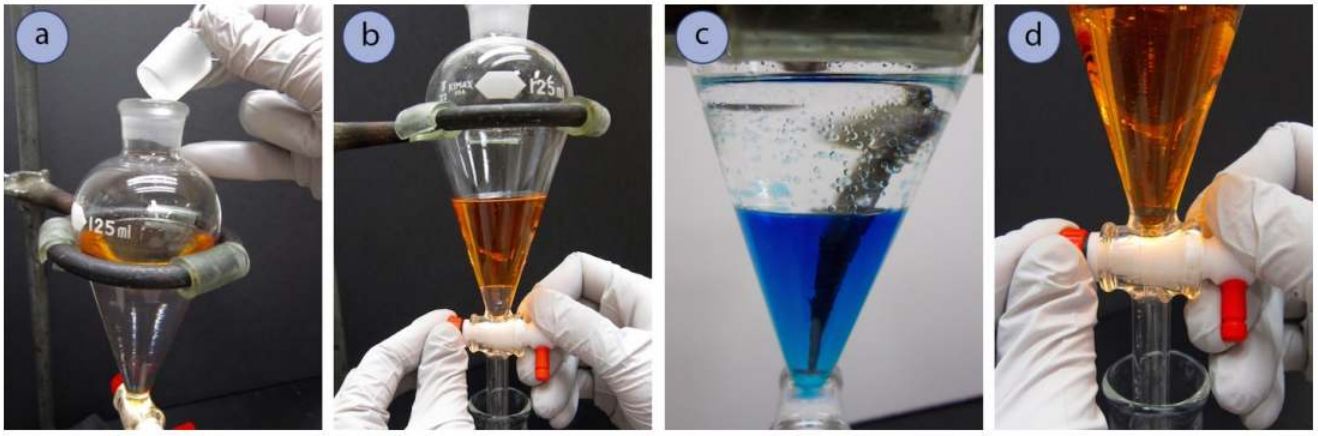
Mix the Solutions

6. Place the stopper on the funnel, and hold the funnel such that the fingers of one hand securely cover the stopper, while the other hand grips the bottom of the funnel.
7. Gently invert the funnel, and swirl the mixture a little. Although it is not uncommon for *some* liquid to creep into the ground glass joint when inverted, it should be minimal.
8. Pressure may build up inside the separatory funnel when solutions are mixed, so immediately after swirling, and with the funnel still inverted, "**vent**" the funnel by briefly opening the stopcock to allow for a release of pressure. With highly volatile solvents (like diethyl ether), a definite "swoosh" can be heard upon venting, and small amounts of liquid may even sputter out the stopcock. **Safety note:** Never point the stopcock toward someone as you vent, as it's possible some liquid may splatter onto him or her.
9. Close the stopcock and mix the solutions a bit more vigorously, periodically stopping to vent the system.
10. Place the separatory funnel upright in the ring clamp to allow the layers to fully separate. The interface between the layers should settle rather quickly, often within 10 seconds or so. If the interface is clouded or not well defined (an emulsion has formed), see the next page.

a) Holding the separatory funnel before shaking, b) Inverting the funnel to mix the components, c) Venting to release pressure.

Separate the Layers

11. Liquid will not drain well from a separatory funnel if the stopper remains on, as air cannot enter the funnel to replace the displaced liquid. Before draining liquid from a separatory funnel, **remove the stopper**.
12. Drain the majority of the bottom layer into a clean Erlenmeyer flask, positioning the ring clamp so that the tip of the separatory funnel is nestled in the Erlenmeyer flask to prevent splashing. Stop draining when the interface is within 1cm of the bottom of the stopcock.
13. Gently swirl the funnel to dislodge any droplets clinging to the glass.
14. Further drain the bottom layer, stopping when the interface just enters the stopcock chamber.



a) Taking the stopper off before draining the funnel, b) Draining to the interface, c) Clinging droplets, d) Stopping when the interface is in the stopcock

Notes:

- It is proper technique to drain the bottom layer through the stopcock, and to pour out the top layer from the top of the funnel. This method minimizes re-mixing the solutions, as only the lower layer touches the stem of the funnel.
- **Never throw away any liquids from an extraction** until you are *absolutely* sure that you have the desired compound. Undesired layers can be properly disposed of when the desired compound is *in your hands* (e.g. after the rotary evaporator has removed the solvent). Mistakes made during extractions (e.g. carrying on with the wrong layer), can be solved as long as the solutions have not been placed in the waste container.
- To clean a separatory funnel, first rinse it with acetone into a waste container. Then wash the funnel with soap and water at your benchtop. Disassemble the Teflon stopcock to dry. Then reassemble and put back into your lab drawer.

a) Pouring out the top layer, b) Labeled layers, c) Drying the separatory funnel with a disassembled stopcock

Adapted from *Step-by-step procedure for Extractions (Large Scale)* by Lisa Nichols.

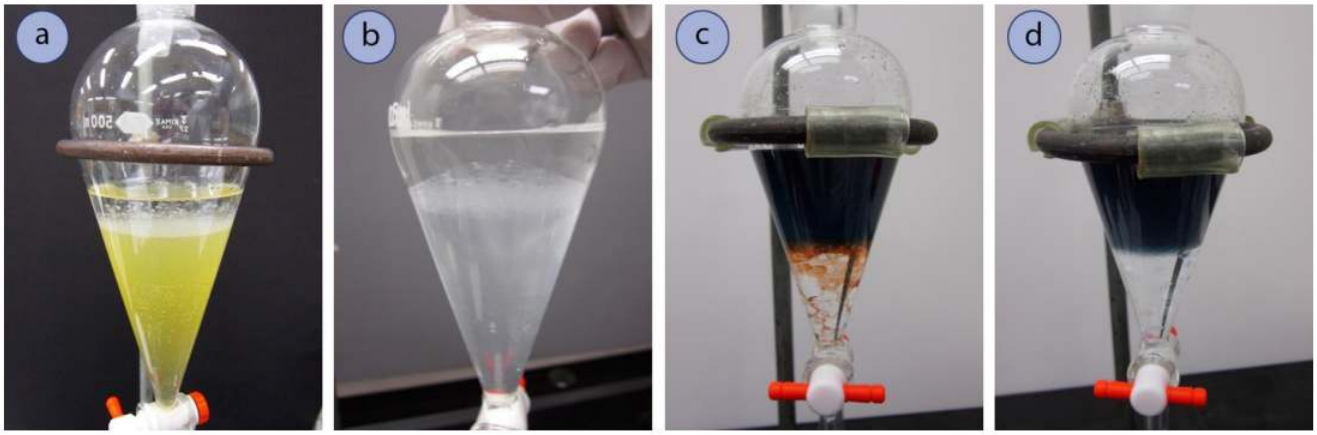
6.5: Emulsions

Emulssions are when tiny droplets of one layer are suspended in the other layer, resulting in no distinct interface between the two layers. Often an emulsion looks like a bubbly mess near the interface, and can even appear to be an odd-looking third layer.

Emulsion formed between dichloromethane and brine (with food coloring).

Emulsions can happen for several reasons:

1. The density of each layer may be so similar that there is weak motivation for the liquids to separate.
2. There may be soap-like compounds or other emulsifying agents present that dissolves some of the components in one another.



a) An emulsion with biodiesel and methanol, b) An emulsion with brine and ethyl acetate, c) An emulsion with dichloromethane and brine (as well as food coloring), d) The emulsion is resolved after addition of water that decreases the density of the top brine layer.

Emulsions can be very difficult to rectify, and it's best if they are avoided in the first place by shaking solutions that are prone to emulsions *gently* in the separatory funnel. Nonetheless, if an emulsion does form, there are some ways to attempt to clarify them:

1. For mild emulsions, gently swirl the layers and try to knock down suspended droplets with a glass stirring rod.
2. Allow the solution to sit for a period of time. With enough time, some solutions do settle out on their own. This of course may not be practical.
3. If an emulsion is formed because the two layers have similar densities, try to alter the density of each layer to make them more different. To help clarify an emulsion, try to decrease the density of the top layer or increase the density of the bottom layer.
4. Try decreasing the solubility of one component in the other. One method is to add NaCl or NH_4Cl to the separatory funnel, which dissolves in the aqueous layer and decreases the ability of organic compounds to dissolve in water ("salting out").

Adapted from *Step-by-step procedure for Extractions (Large Scale)* by Lisa Nichols.

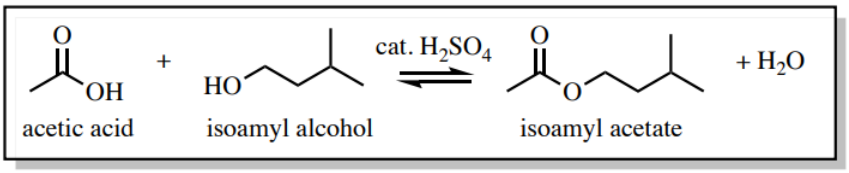
6.6: Common "work-up"

Common Washes

Water

The most common wash in separatory funnels is probably water. It is cheap, non-hazardous, and works well to remove may impurities found alongside a desired product. Water can potentially remove water-soluble impurities from an organic layer, as long as they are present in quantities that do not exceed their water solubility. The following are common materials that can be removed with a water wash: unconsumed acid or base, many ionic salts, and compounds that can hydrogen bond with water and are relatively small (e.g. CH_3CH_2OH or CH_3COCH_3).

To demonstrate the effectiveness of a water wash, a Fischer esterification reaction was conducted to produce isoamyl acetate. In this reaction, an excess of acetic acid is used to drive the reaction through Le Chatelier's principle, and the acetic acid had to be removed from the product during the purification process.



Reaction scheme for the synthesis of isoamyl acetate.

The reaction was then "worked up" by pouring the reaction mixture into a separatory funnel and washing the organic layer with water, sodium bicarbonate, and brine in succession. The main purpose of the water wash was to remove the majority of the catalytic sulfuric acid and the excess acetic acid, while the sodium bicarbonate wash neutralized the rest.

The sodium bicarbonate wash in this example was necessary (and discussed in the next section) because a water wash alone may not fully remove the acetic acid. It's important to know that when a compound is "water soluble" it does not necessarily mean it is "organic insoluble", a common misconception that arises from the "like dissolves like" principle. The ability of acetic acid and other polar compounds to dissolve in the organic layer of a separatory funnel should not be ignored.

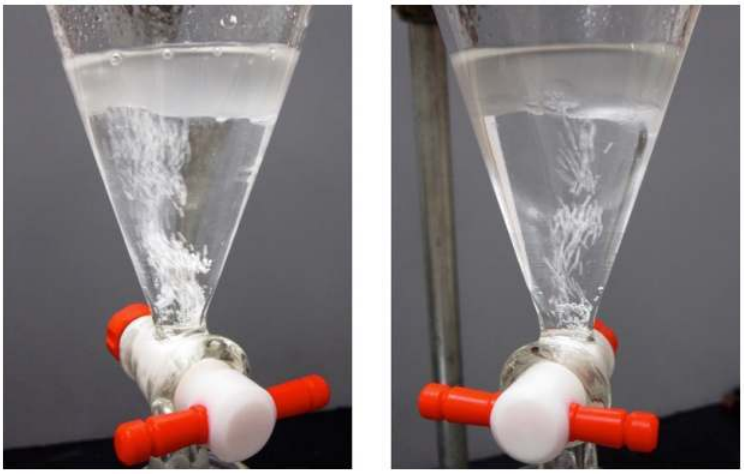
Sodium Bicarbonate and Sodium Carbonate

A normal part of many work-ups includes neutralization. It is important to neutralize any organic solvent that was exposed to an acidic or basic solution as trace acid or base may cause undesired reactions to occur when the solutions are concentrated.

Aqueous solutions of saturated sodium bicarbonate ($NaHCO_3$) and sodium carbonate (Na_2CO_3) are basic, and the purpose of these washes is to **neutralize an organic layer that may contain trace acidic components**. Even if an organic layer should not in theory dissolve very polar components such as acid, acid sometimes "hitches a ride" on polar components that may dissolve in an organic layer, such as small amounts of alcohol or water.

Neutralizing with sodium bicarbonate and sodium carbonate solutions produces carbonic acid (H_2CO_3), which is in equilibrium with water and carbon dioxide gas. This means that solutions often bubble during a neutralization wash in a separatory funnel.

Safety note: To prevent excess pressure from being generated by the release of carbon dioxide gas into a separatory funnel during neutralization, the layers should be gently swirled together before placement of the stopper. **They should be vented directly after inversion, and more frequently than usual**.



Dilute $NaHCO_3$ solution (bottom layer) vigorously bubbling during the wash of an acidic organic (top) layer.

Testing the pH After a Wash

To test whether a base wash with $NaHCO_3$ or Na_2CO_3 was effective at removing all the acid from an organic layer, it is helpful to test the pH. It is not possible to test the pH of an organic solution directly, however it is possible to test the pH of an aqueous solution that the organic solution has been in contact with using pH paper.

Brine (Saturated NaCl)

In some experiments, an organic layer may be washed with brine, which is a saturated solution of NaCl. The purpose of this wash is to **remove large amounts of water** that may be dissolved in the organic layer. Brine works to remove water from an organic layer because it is highly concentrated (since NaCl is so highly water soluble). A saturated NaCl solution causes water to draw into the solution from the organic layer.

Drying Agents

An organic layer is always treated with a drying agent after having been exposed to water in a separatory funnel. Drying agents are anhydrous inorganic materials that favorably form "hydrates", which incorporate water molecules into their solid lattice structure (for example, Na2SO4·7H2O). A drying agent is swirled with an organic solution to **remove trace amounts of water**. If drying agents are used to remove water, you might wonder "Why bother with brine; why not use lots of drying agent when the time comes?" The main reason to limit the amount of water present in an organic solution *before* the drying agent step is that the drying agent will often adsorb the compound along with water. Using as little as possible will maximize the yield.

Various drying agents: Anhydrous Na_2SO_4, anhydrous $MgSO_4$, Drierite $CaSO_4$, and anhydrous $CaCl_2$.

The most useful drying agents indicate when they have completely absorbed all of the water from the solution. Anhydrous magnesium sulfate ($MgSO_4$) is a fine, loose powder, but its hydrate is clumpy and often clings to the glass. A typical drying procedure is to add anhydrous $MgSO_4$ to an organic solution until it *stops* clumping and fine particles are seen, which indicate that there is no longer water available to form the clumpy hydrates.

Anhydrous calcium sulfate ($CaSO_4$), can be purchased containing a cobalt compound that is blue when dry and pink when wet (this is then sold under the name Drierite). In this way, blue Drierite can be used as a visual indicator for the presence of water.

a) Ethyl acetate solution containing fine particles of anhydrous $MgSO_4$, b) Ethyl acetate with clumpy $MgSO_4$ hydrate,c) Dry Drierite (blue), d) Wet Drierite (pink)

In some procedures Na_2SO4 or $CaCl_2$ are used if the solution is incompatible with $MgSO_4$. An advantage to these drying agents is that their granules are not easily dispersed, allowing for the solutions to be easily decanted. In many situations drying agents are interchangeable.

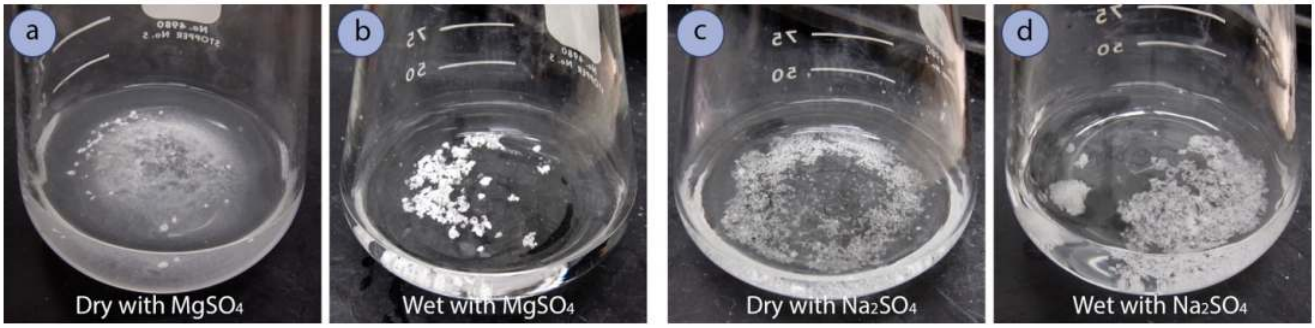
Survey of drying agents

Drying Agent	Hydrate formula(s)	Practical Comments	Other Comments
Magnesium sulfate	$MgSO_4 \cdot 7H_2O$	Quickly removes most water, and can hold a lot for its mass (0.15-0.75g water per g desiccant). Is a fine powder, so must be gravity filtered. Its high surface area means it will somewhat adsorb compound: be sure to rinse after filtering.	$Mg(H_2O)_2^{+4}$ is somewhat acidic, so is incompatible with highly acid-sensitive groups.
Sodium Sulfate	$Na_2SO_4 \cdot 7H_2O$ $Na_2SO_4 \cdot 10H_2O$	Removes water at a moderate rate, so the solution should be allowed to sit with the drying agent for some time. Can hold a lot of water for its mass (1.25g water per g desiccant), but may leave small amounts of water remaining. Solutions with Na_2SO_4 can usually be decanted.	Cannot dry diethyl ether well unless a brine wash was used.
Calcium chloride	$CaCl_2 \cdot 2H_2O$ $CaCl_2 \cdot 6H_2O$	Quickly removes water well, although larger quantities are needed than other drying agents (holds 0.30g water per g desiccant). If using a fine powder, the solution must be gravity filtered and drying agent rinsed. If using pellets, the solution should be allowed to sit for a few minutes, then decanted.	Absorbs water as well as methanol and ethanol.

Drying Agent	Hydrate formula(s)	Practical Comments	Other Comments
Calcium sulfate (Drierite)	$CaSO_4 \cdot 12H_2O$ $CaSO_4 \cdot 2H_2O$	Quickly removes water, but needs large quantities as it holds little water per gram. Are most often used in desiccators and drying tubes, not with solutions.	

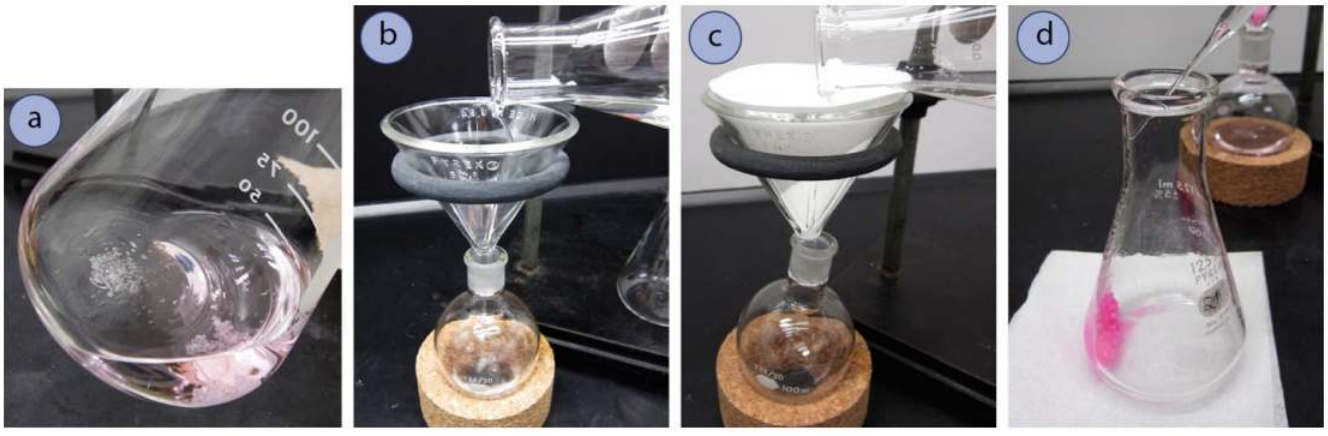
Drying Agents Procedure

1. The organic solution to be dried must be in an **Erlenmeyer flask**, as solutions can easily splash out of beakers when swirled.
2. Add a small portion of drying agent to the flask, the size of one pea and swirl the solution. Be sure to close the jar of drying agent when not in use, as the reagents are hygroscopic. After a short period of time, inspect the mixture closely.
   - If the entire drying agent clumps into pieces that are much larger than the original size, there is still water remaining in the flask. Add another portion of drying agent and swirl.
   - A solution is nearing dryness when fine particles are noticed that don't cling to other particles or to the glass when swirled. A wet organic solution can be cloudy, and a dry one is always clear.



Drying ethyl acetate with anhydrous $MgSO_4$ and Na_2SO_4: a) Dry $MgSO_4$ (very fine particles), b) Wet $MgSO_4$ (hydrate clumps), c) Dry Na_2SO_4 (small particles), d) Wet Na_2SO_4 (hydrate clumps)

4. When the solution is dry, separate the drying agent from the solution:
   - Carefully decant the solution into an appropriately sized round-bottomed flask, being sure to fill the flask no more than halfway.
   - If using $MgSO_4$, instead of decanting, gravity filter the solution into an appropriately sized round-bottomed flask.
   - With all drying agents, rinse the drying agent with a few mL of fresh organic solvent, and add the rinsing to the round-bottomed flask.



a) Wet hydrate of Na_2SO_4 can stick to the glass, b) Decanting a solution, c) Gravity filtering a solution, d) Rinsing the residual drying agent in the flask

Adapted from *Common Extraction Washes* by Lisa Nichols.

CHAPTER OVERVIEW

7: Rotary Evaporation

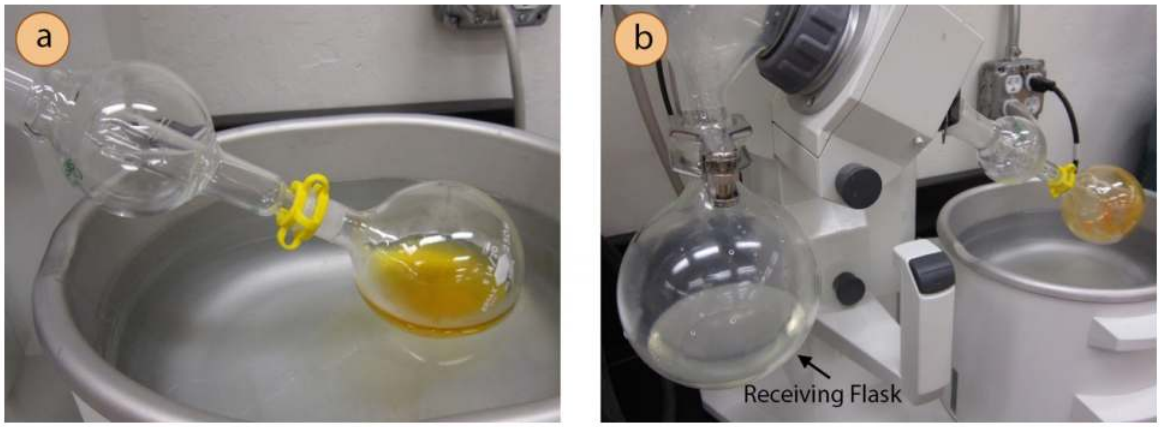
7.1: Overview of Rotary Evaporation

It is very common for a desired compound to be dissolved in a solvent. Solvents are used in separatory funnel extractions and column chromatography, and the solvent must be removed in order to isolate te desired compound. Solvents are regularly chosen that have lower boiling points than the compound of interest, so that there is some mechanism for their removal. In theory, a solution could simply be placed on a heat source to boil away the lower-boiling solvent, but this approach is not often used.

The preferred method for solvent removal in the laboratory is by use of a **rotary evaporator**, also known as a "rotovap". A solution in a round bottomed flask is placed in a water bath and rotated while the system is partially evacuated (by a water aspirator or vacuum pump). The reduced pressure in the apparatus causes the solvent to boil at a lower temperature than normal and rotating the flask increases the liquid's surface area and thus the rate of evaporation. The solvent vapor condenses when it comes into contact with a condenser and drips into a receiving flask. When the solvent is removed, the concentrated compound is left in the flask.

Rotary Evaporator.

Removal of solvent by a rotary evaporator is superior to evaporation under atmospheric pressure for many reasons. The process is much quicker (often takes less than 5 minutes), uses lower temperatures (so decomposition is unlikely), and uses less energy than boiling with a heat source.



a) Solution is placed on a rotary evaporator, b) After evaporation (solvent is in the receiving flask)

Adapted from *Overview of Rotary Evaporator* by Lisa Nichols.

7.2: Rotary Evaporation Procedure

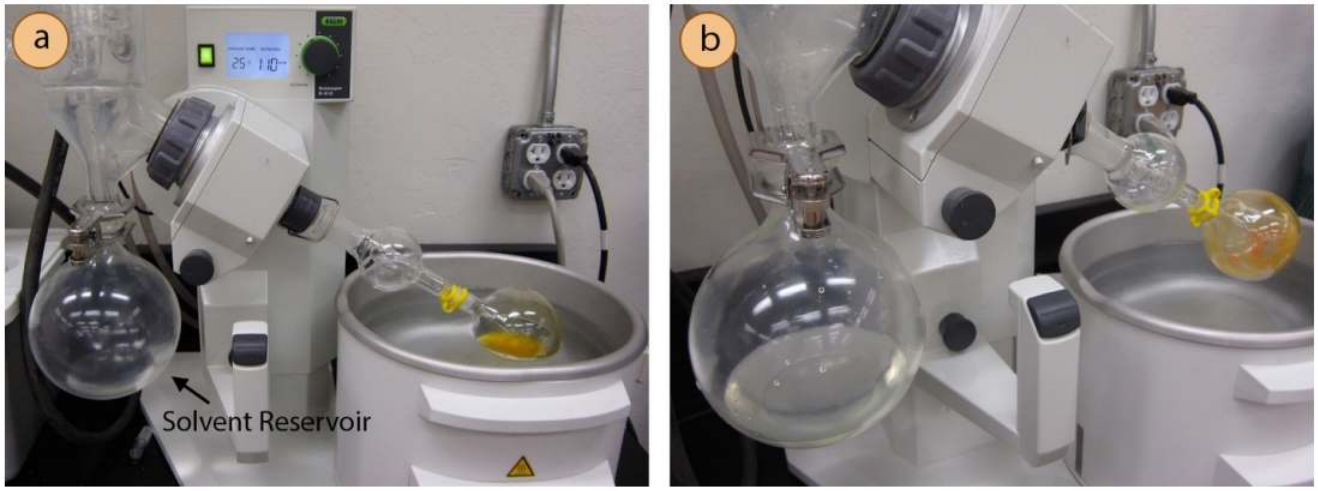
Rotary Evaporation Procedure

1. Pre-weigh a round bottomed flask, and fill it no more than half-full with the solution to be evaporated.
2. Connect the flask to the evaporator's "bump trap" using a plastic clip ("Keck clip"). The bump trap prevents foaming or splashing solutions ("bumping") from dirtying the condenser or collecting in the receiving flask where components cannot be recovered. A bump trap should be kept clean so that if solutions do collect there, they can be rinsed back into the original flask without worry about contamination from previously evaporated samples.
3. Begin slowly turning the round bottom flask. Use the joystick to lower the flask into the water so that the flask is partially submerged. Be sure the flask is not positioned so low that the joint with the plastic clip is in the water.
4. Turn on the vacuum source. A hissing sound should be heard from air being pulled through the stopcock. The partial vacuum will help hold the flask securely onto the bump trap.
5. Begin rotating the flask at a medium rate by adjusting the rotation notch (to roughly 110 rpm).
6. Close the stopcock on the evaporator by turning it perpendicular to the bleed valve. The hissing sound should stop, and the pressure inside the apparatus will decrease further.

a) Sample to be evaporated, b) Attachment of sample to bump trap on rotary evaporator, c) Lowering of sample into water bath.

7. Allow the solution to evaporate. There is no set amount of time required for complete evaporation.

- If the expected compound is a solid, keep evaporating until a solid or thin film appears.
- If the expected compound is a liquid, evaporate until the approximate expected volume is seen and it appears that the liquid level is no longer changing.
- If solvent bumps from the flask into the bump trap, raise the flask slightly out of the water bath, and/or allow small amounts of air into the apparatus by partially turning the stopcock. If large amounts of liquid have bumped, stop the evaporation and rinse the trap with solvent such that the rinsing is added back to the flask. Evaporate the solvent.



a) Evaporating a solution, b) Formation of solid in the flask.

To stop the evaporation, reverse all the previous steps: open the stopcock, stop the rotation, turn off the vacuum source, lift the flask from the water bath, and remove the flask.

Adapted from *Step-by-Step Procedure for Rotary Evaporation* by Lisa Nichols.

CHAPTER OVERVIEW

8: Reflux

8.1: Overview of Reflux

Overview of Reflux

A **reflux** setup allows for liquid to boil and condense, with the condensed liquid returning to the original flask. The liquid remains at the boiling point of the solvent (or solution) during active reflux. A reflux apparatus allows for easy heating of a solution, but without the loss of solvent that would result from heating in an open vessel. In a reflux setup, solvent vapors are trapped by the condenser, and the concentration of reactants remains constant throughout the process.

Many chemical reactions need to be heated in the organic laboratory to speed up the rate of the reaction. The main purpose of refluxing a solution is to heat a solution in a controlled manner at a constant temperature in order to conduct a chemical reaction.

Reflux apparatus

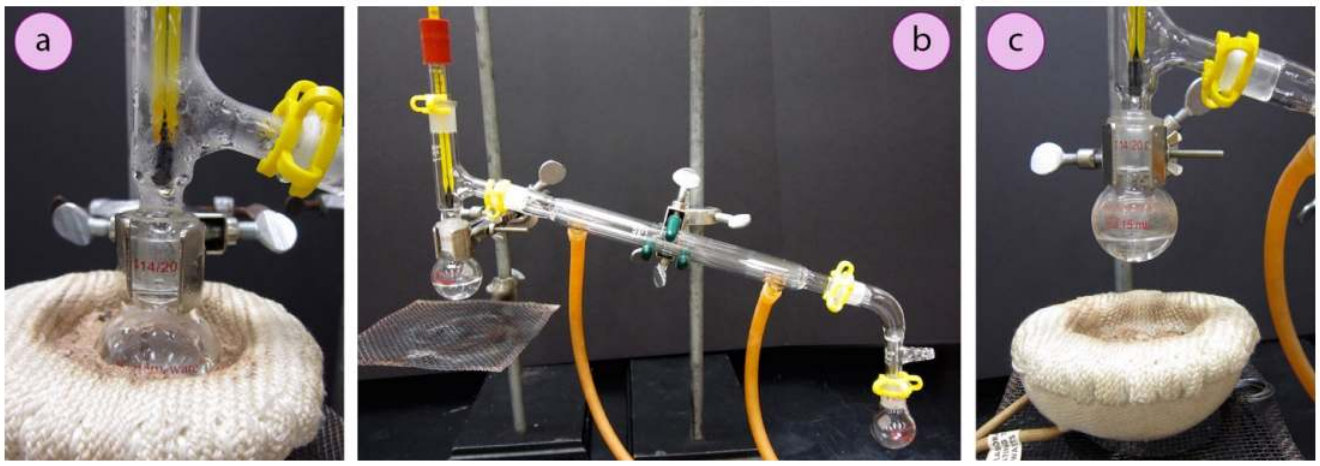
Adapted from *Reflux* by Lisa Nichols.

8.2: Heating Mantles

Heating mantles are a relatively safe way to heat flammable organic liquids in a round bottomed flask. The mantles are cup-shaped and designed for different sizes of round bottomed flask. The mantles should *never* be connected directly to the outlet, but first to a "**Variac**" which then connects to the outlet and delivers variable voltage to the mantle. A Variac set to "100" would be equivalent to plugging the mantle directly into the wall (100%), while a setting of "50" means the delivered voltage is halved (50%). By controlling the delivered voltage, Variacs are used to regulate the temperature of a heating mantle.

a) Three sizes of heating mantle, b) Heating mantle connected to a Variac.

Heating mantles take some time to warm up and cool down. The mantle will remain warm even after turning off the Variac, and therefore flasks have to be removed from the mantle in order to cool.



a+b) Distillation using a heating mantle, c) Cooling of a flask after heating.

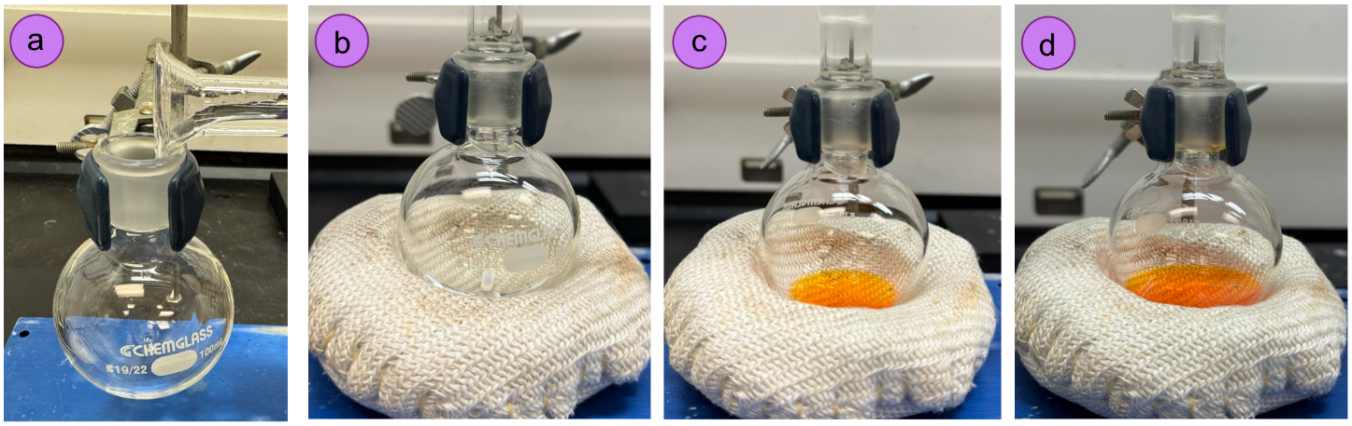
Safety note: the main hazard with heating mantles is that flammable organic liquids spilled on the surface of a hot mantle do have the possibility of ignition.

Adapted from *Heating Mantles* by Lisa Nichols.

8.3: Reflux Procedure

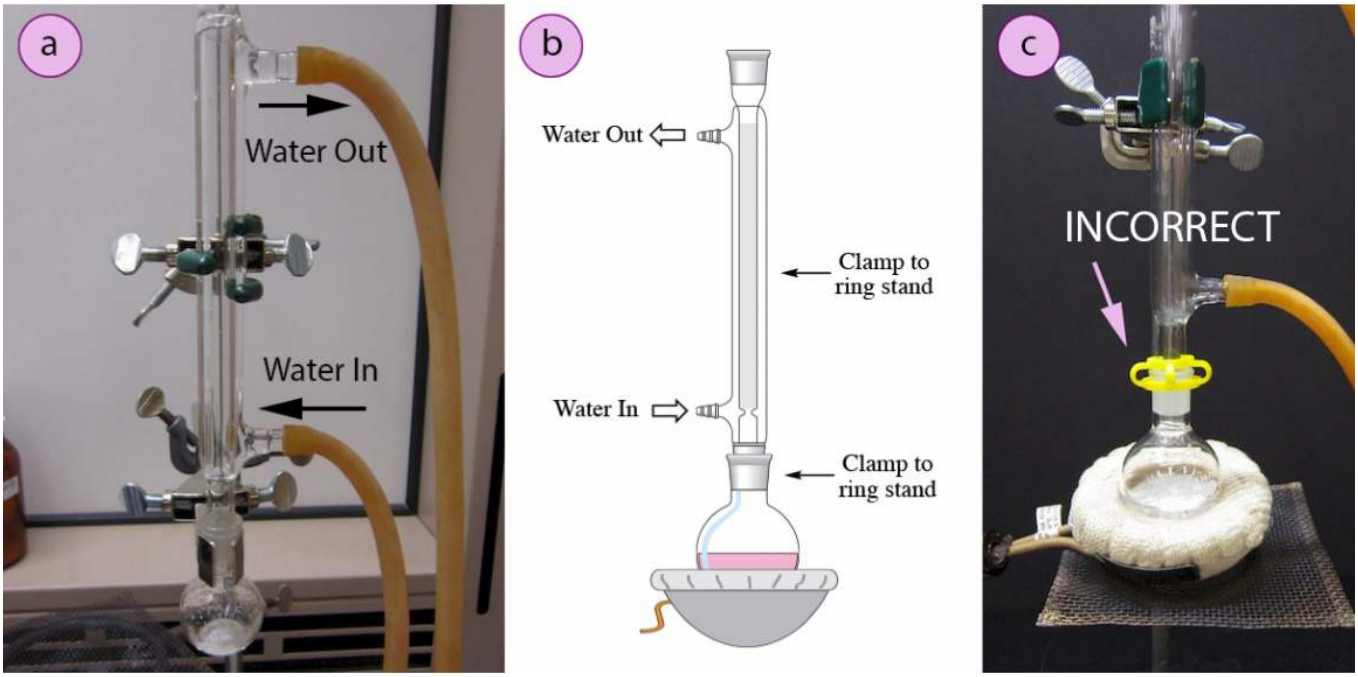
Step-by-Step Procedure

1. Pour the solution to be refluxed into a round bottomed flask, and clamp it to the ring stand or latticework with an extension clamp. The flask should be no more than half full.
2. Add a few boiling stones for **bump prevention**. Boiling stones should not be used when refluxing concentrated solutions of sulfuric or phosphoric acid, as they will colorize the solution.



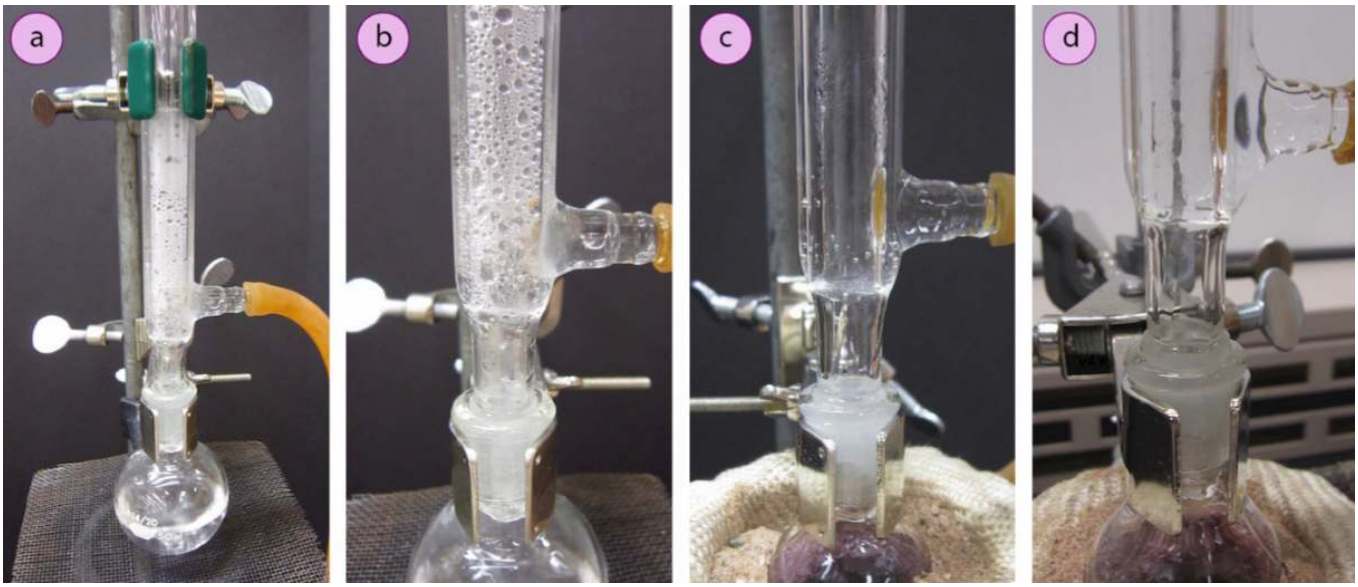
NEED CAPTION

3. Place rubber hoses on a condenser (wet the ends first to allow them to slide on), then attach the condenser vertically to the round bottomed flask. If using a tall condenser, clamp the condenser to the ring stand or latticework. Be sure the condenser fits snugly into the flask.
   Safety note: if the pieces are not properly connected and flammable vapors escape, they may be ignited by the heat source. Do not connect the round bottomed flask and condenser with a plastic clip.
4. Connect the hose on the lower arm of the condenser to the submersible pump and place it in a glass bowl full of ice water. Allow the hose on the upper arm to drain to the bowl. It is important that water goes in the bottom of the condenser and out the top (so water flows against gravity).



a) Reflux apparatus, with arrows indicating the direction of water flow, b) Reflux diagram, c) Incorrect clamping of a reflux apparatus

5. Hold the heating mantle in place beneath the round bottom flask with an adjustable lab jack. Turn on the heating mantle and begin circulating a steady stream of water through the hoses by turning on the submersible pump in bowl of ice water.
6. The appropriate heating rate occurs when the solution is vigorously boiling and a "**reflux ring**" is seen roughly one-third of the way up the condenser. A "reflux ring" is the upper limit of where hot vapors are actively condensing. Subtle movement may be seen in the condenser as liquid drips down the sides of the condenser.
7. If following a procedure in which you are to reflux for a certain time period (e.g. "reflux for one hour"), the time period should begin when the solution is not just boiling but actively refluxing in the bottom third of the condenser.
8. The heat should be turned down if the reflux ring climbs to half-way up the condenser or higher, or else vapors could escape the flask.
9. After the reflux is complete, turn off the heat source and remove the flask from the heat by dropping the heat source down using the lab jack.



a+b) Condensation seen in the condenser when refluxing water, c) Reflux ring of ethanol seen subtly in the bottom third of the condenser, d) Distortion of the ring stand in the condenser due to the refluxing ethanol solution.

Adapted from *Reflux* by Lisa Nichols.

CHAPTER OVERVIEW

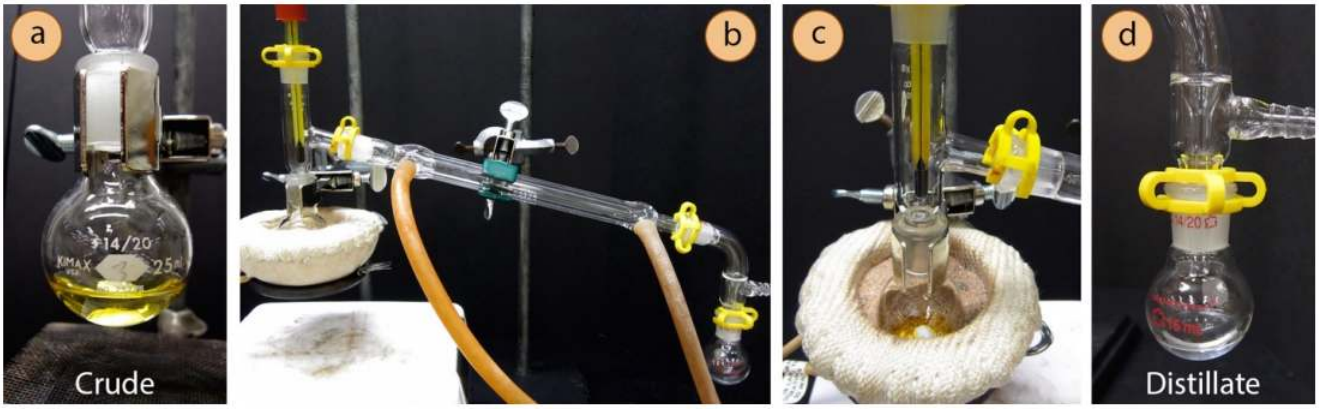
9: Distillation

9.1: Overview of Distillation

Distillation is a purification method for liquids, and can separate components of a mixture if they have significantly different boiling points. In a distillation, a liquid is boiled in the "distilling flask," then the vapors travel to another section of the apparatus where they come into contact with a cool surface. The vapors condense on this cool surface, and the condensed liquid (called the "distillate") drips into a reservoir separated from the original liquid. In the simplest terms, a distillation involves boiling a liquid, then condensing the gas and collecting the liquid elsewhere.

9.2: Simple Distillation

Simple distillation is an excellent purification tool for many liquids, and can be used to purify products from a chemical reaction. The series of photos below show the distillation of a crude sample of isoamyl acetate, formed through a Fischer esterification reaction. The crude sample was originally yellow, but the distillate was colorless, making obvious the removal of some materials through the process.



a) Crude yellow isoamyl acetate formed from a Fischer esterification, b+c) Simple distillation, d) Colorless distillate

A simple distillation works well to purify certain mixtures, specifically to separate a liquid from non-volatile impurities (e.g. solids or salts), or from small amounts of significantly higher or lower boiling impurities. A general guideline is that a simple distillation is capable of separating components if the **difference in boiling point of the components is greater than 100°C**. A simple distillation does not work well to purify a mixture that contains components with similar boiling points (when the difference in b.p. is <100°C).

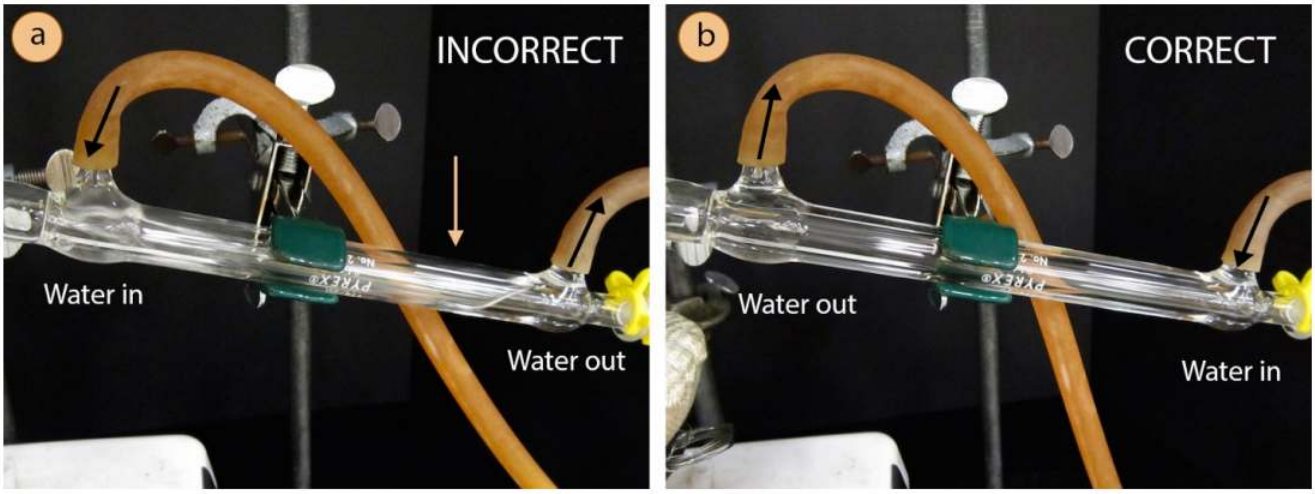
Adapted from *Uses of Simple Distillation* and *Separation Theory* by Lisa Nichols.

9.3: Simple Distillation Procedure

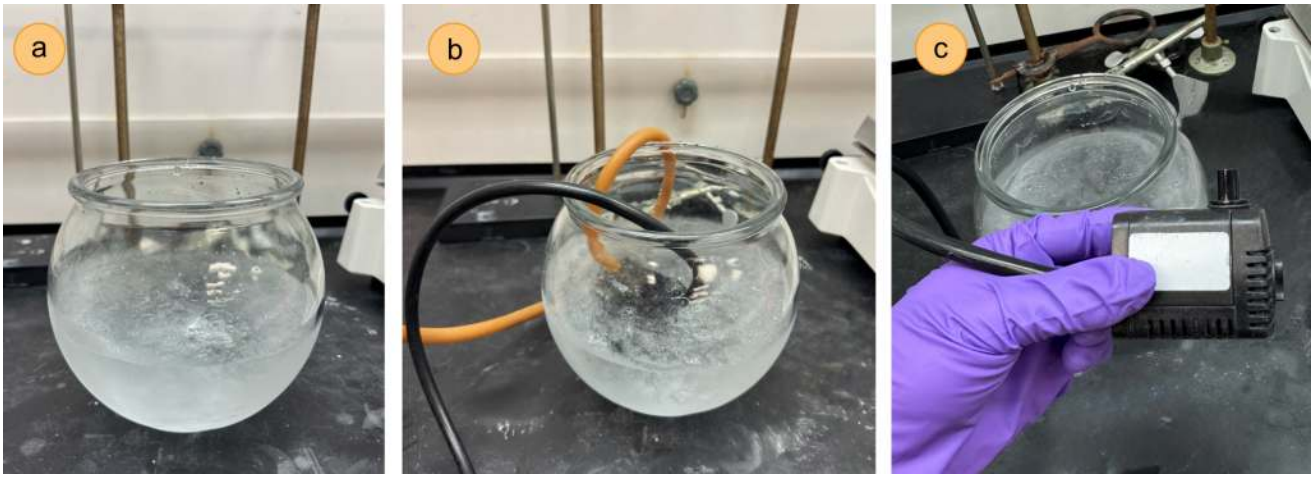
Condenser Hoses

The condenser is an intricate piece of glassware, and allows for cold water to circulate through the distillation apparatus. The circulating water does not mix with the sample to be purified, but instead passes through another jacket surrounding the hollow tube where the gaseous sample travels. It is important that the water jacket be *full* of cold water, to maximize the efficiency of condensing the gaseous sample. It is for this reason that the water hoses must be attached to the condenser in a certain way.

A hose should connect from the submersible pump to the *lower* arm of the condenser, forcing water to travel against gravity through the condenser. The hose connecting the *upper* arm of the condenser should then drain to the bowl of ice water. By forcing the water uphill, it will completely fill the condenser.



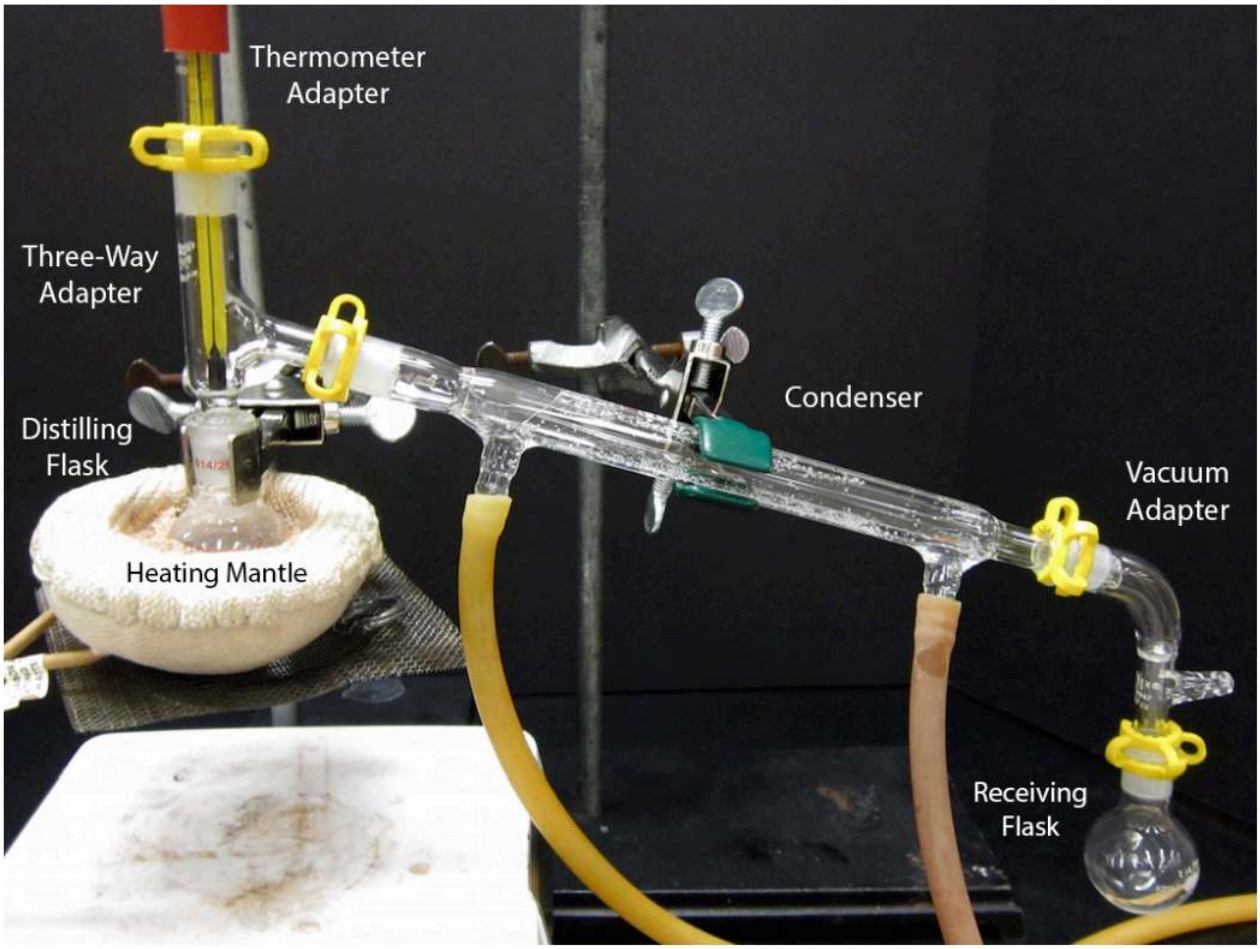
a) Condenser hoses connected incorrectly: notice that water does not fill the condenser's jacket (as indicated with the small arrow),
b) Condenser hoses connected correctly



a) Fishbowl with ice and water, b) Submersible pump, c) Hose connected to submersible pump and placed in fishbowl.

Simple Distillation Procedure

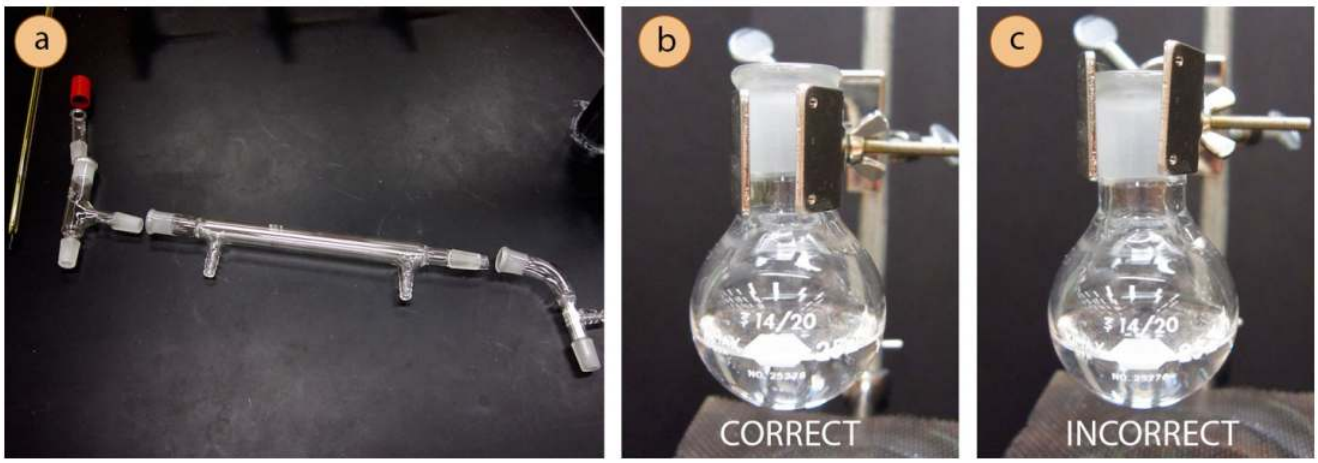
An assembled simple distillation apparatus is shown below. Assembly of this complicated apparatus is shown in this section piece by piece. The glassware used for this apparatus is quite expensive, and undoubtedly your instructor would appreciate care being taken when using this experiment.



Simple distillation apparatus

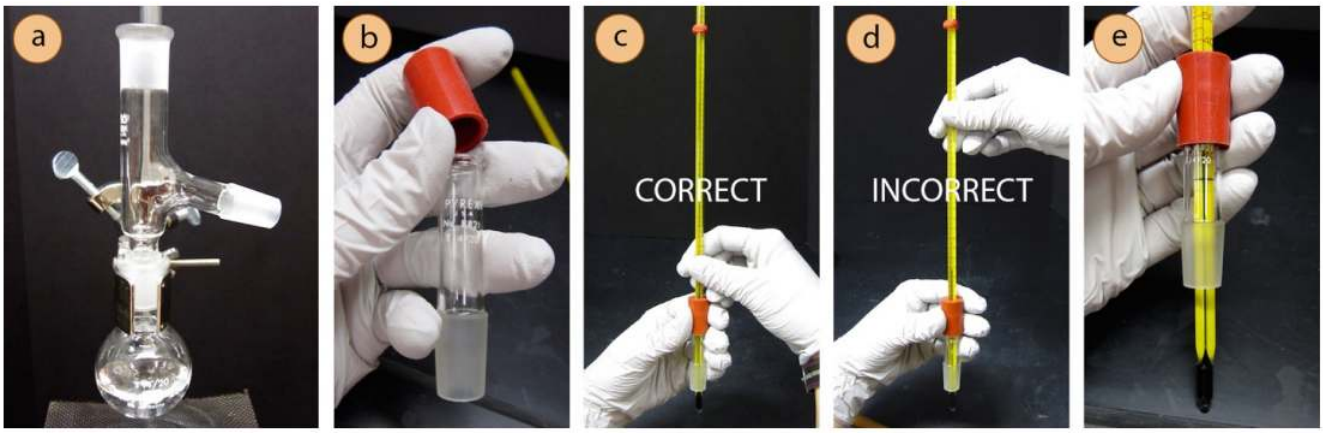
Assemble the Apparatus:

1. Pour the liquid to be distilled into a round bottomed flask, trying to avoid pouring liquid on the ground glass joint. The flask should ideally be between one-third to one-half full of the liquid to be distilled. If the flask is more than half full, it will be difficult to control the boil.
2. Add a few boiling stones to the solution to prevent bumping during heating.
3. Use a metal extension clamp to secure the round bottomed flask containing the sample to the ring stand or latticework. The clamp should securely hold the joint *below* the glass protrusion on the flask.



a) Distillation apparatus arranged on the benchtop, b) Correct clamping of the round bottomed flask, c) Incorrect clamping

4. Attach a three-way adapter (or "distilling adapter") to the round bottomed flask.
5. Attach a rubber fitting to the top of a thermometer adapter by stretching it over the glass. Then delicately insert a thermometer into the hole of the rubber fitting.
   Safety note: While inserting the thermometer, position your hands near the joint, not far from the joint or it may snap.

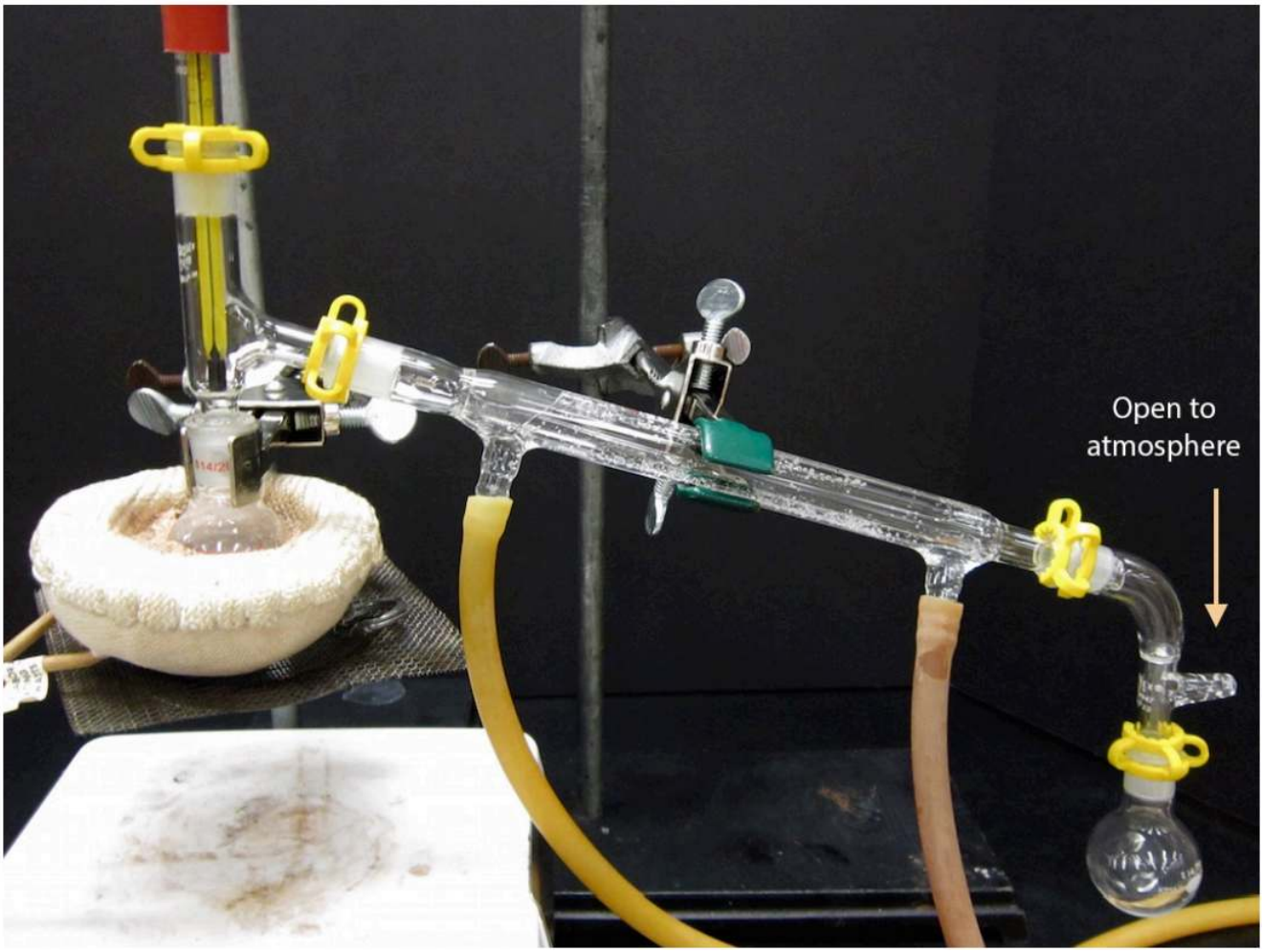


a) Flask and three-way adapter, b) Attachment of rubber fitting to thermometer adapter, c) Correct way to insert the thermometer, d) Incorrect way (thermometer will break), e) Assembled thermometer piece.

6. Connect the thermometer adapter to the three-way adapter, securing the joint with a plastic clip ("**Keck clip**"). The clip is directional, and if it doesn't easily snap on it is probably upside down. Check that the clip is not broken, and if it is, replace it.
7. Adjust the thermometer so the bulb is just below the arm of the adapter. If the bulb is positioned too high, it will not register the correct temperature of the vapors as they make their turn toward the condenser.
8. Wet the ends of two hoses using the faucet or by dipping into a beaker of water, then *twist* the wet ends of the two arms on the condenser. The hoses should fit onto the condenser arms higher than one centimeter or else water pressure may cause them to pop off.
9. Connect the condenser to the rest of the apparatus with another plastic clip. Use another clamp to secure the condenser to the ring stand or latticework, and position the condenser at a slight downward angle.

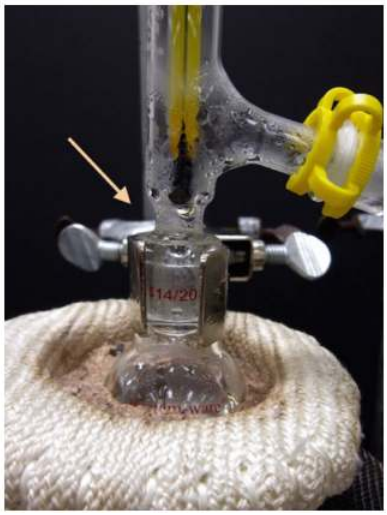
a) Connecting a plastic clip to secure a joint, b) Incorrect clamping: a plastic clip should not be used to connect the distilling flask to the three-way adapter (as indicated with the arrow), c) Correct positioning of the thermometer in the three-way adapter, d) Preparing the condenser

10. Connect the rubber hose attached to the *lower* arm of the condenser to the submersible pump and allow the hose attached to the *upper* arm of the condenser to drain to the bowl with ice water. The hoses may point either up or down.
11. Connect a vacuum adapter to the end of the condenser with a plastic clip. A receiving flask should be placed under the vacuum adapter. This can be another round bottom flask, graduated cylinder, or Erlenmeyer flask.
12. A completed distillation apparatus is shown below. No parts should be able to jiggle or they are not adequately clamped. Although it may appear to be a closed system (which would be dangerous to heat), the system is actually open to the atmosphere at the arm in the vacuum adapter.
13. Position the heating mantle beneath the round bottom flask using an adjustable lab jack that allows for some mechanism by which the heat can be lowered and removed at the end of the distillation.
14. When ready to begin the distillation, start circulating water in the condenser.



Completed simple distillation setup

15. Begin heating the flask. Depending on the size of the flask and the necessary temperature, the liquid may start boiling between 1-20 minutes. Condensation will eventually be seen on the sides of the flask, but the reading on the thermometer will not increase until the vapors immerse the thermometer bulb. Eventually condensation should be seen in the three-way adapter, and the thermometer temperature will rise.
16. Eventually a droplet may be seen traveling down the length of the condenser and liquid will begin to collect in the receiving flask. An appropriate distillation collects distillate at a **rate of 1 drop per second**. Distilling at too fast a rate prevents proper equilibration between liquid and gas phases, and results in poor separation. Although a slow distillation is necessary, heating the solution such that *nothing* is dripping into the receiving flask simply wastes time. There is no reason to heat so gradually that nothing distills.
17. Pay attention to the temperature on the thermometer throughout the entire time that liquid is distilling. Record the temperature when it has plateaued, or the highest temperature seen. If you know a mixture is distilling, record the temperature range over which liquid is actively dripping into the receiving flask. Only record temperatures that correspond to active distillation and full immersion of the thermometer with vapor.



Active distillation. Arrow indicates the joint that needs to be secure to prevent fire.

Stop Distilling

Cease the distillation when one of the following takes place:

1. If the liquid is nearly gone in the distilling flask.
   Safety note: It is unsafe to distill a flask to dryness as side reactions can occur when components are concentrated. This is especially dangerous with compounds that can form peroxides, as these can become explosive when reacting with concentrated solutions. Additionally, when the entire sample is in the gas phase, the system is not longer restricted to the boiling point of the liquid and may reach dangerous temperatures.
2. If the thermometer temperature was higher during the distillation but has dropped significantly. This normally corresponds to a lull between distilling two components, and means that one component has essentially completed distilling.
3. If the thermometer temperature has a dramatic spike. This normally corresponds to the beginning of distillation of a higher boiling component, and would contaminate the distillate if allowed to continue.
4. If anything surprising or unusual happens, such as thick smoke, darkening/thickening in the distilling flask, or uncontrollable bumping.

To stop the distillation, lower and remove the heat source from the round bottomed flask.

Glass wool

For high-boiling liquids, it may be difficult for vapors to reach the condenser as they too quickly are cooled by the glassware which is in contact with the air in the room. It may be helpful to insulate the distilling flask and three-way adapter to better retain heat and allow the sample to remain in the gas phase longer.

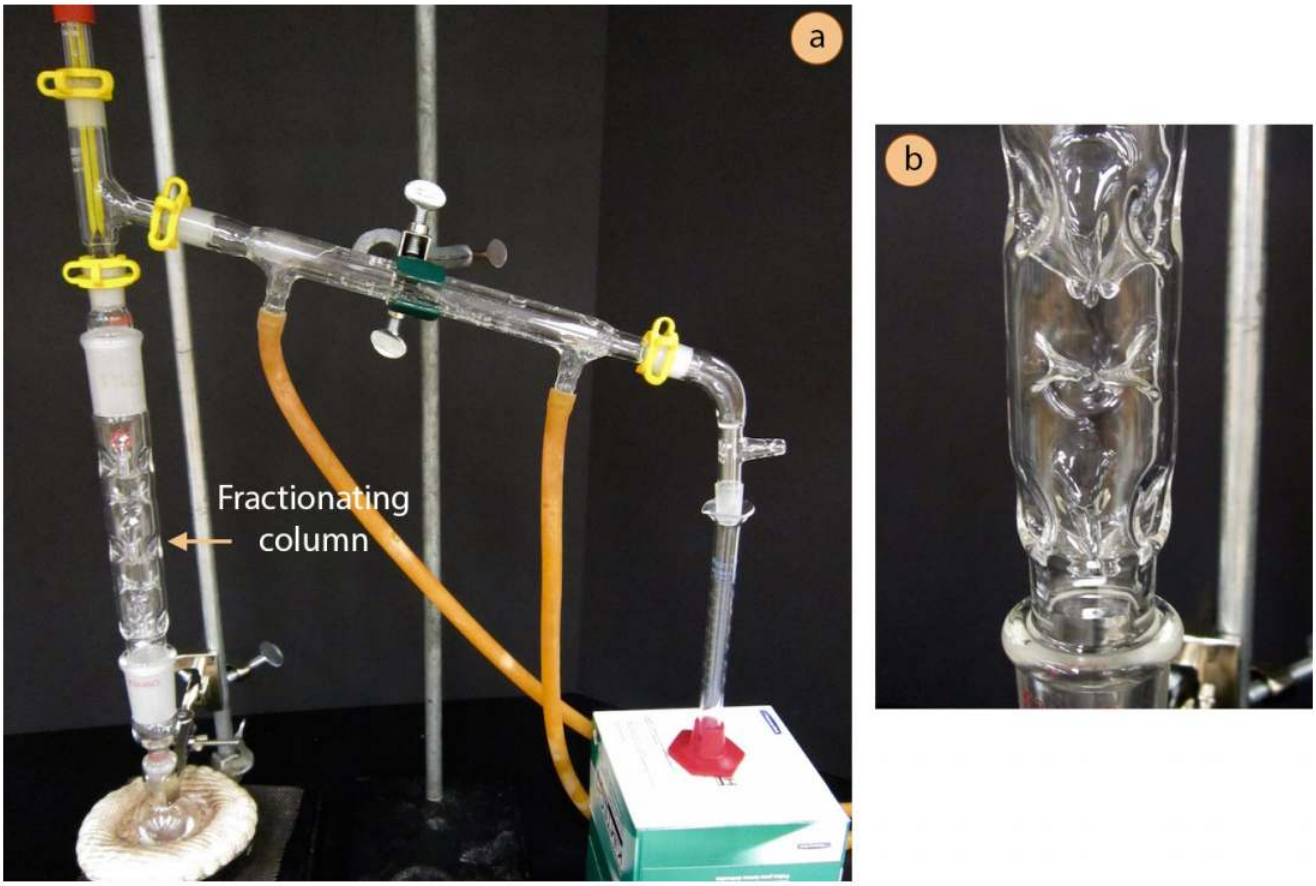
To insulate a portion of the distillation, wrap the parts prior to the condenser with **glass wool**. A small gap can be left in the insulation in order to "peek in" on activity inside the apparatus. Glass wool has an appearance similar to cotton, but unlike cotton is not flammable so is useful as an insulating material when an apparatus is to be heated.

a) Fibrous glass wool, b) Cotton-like glass wool, c) Wrapping flask and three-way adapter with glass wool, d) Foil over top of glass wool

Adapted from *Step-by-Step Procedures* by Lisa Nichols.

9.4: Fractional Distillation

A simple distillation is incapable of significant purification if the boiling points of the components are too close. When the difference in boiling points is less than 100^{o}C, a modification is necessary, namely insertion of a fractionating column between the distilling flask and three-way adapter.



a) Fractional distillation setup, b) Zoom in of fractionating column during distillation, with condensation dripping from the glass indentations

A fractionating column essentially allows for many successive distillations to take place at once, without dismantling the apparatus. A fractionating column contains indentations or a packing material with lots of surface area. The vapors temporarily condense on these surfaces and the heat of the distillation allows those pools of liquid to vaporize again. Every vaporization-condensation event (called a "theoretical plate") is similar to a simple distillation.

The choice of what fractionating column to use for which application depends in part on availability and the task at hand. Several columns are shown below: a) Vigreux column with glass indentations, b) Steel wool column made simply be loosely inserting steel wool into the cavity of a fractionating column, c) Glass beads filled in the cavity of a fractionating column.

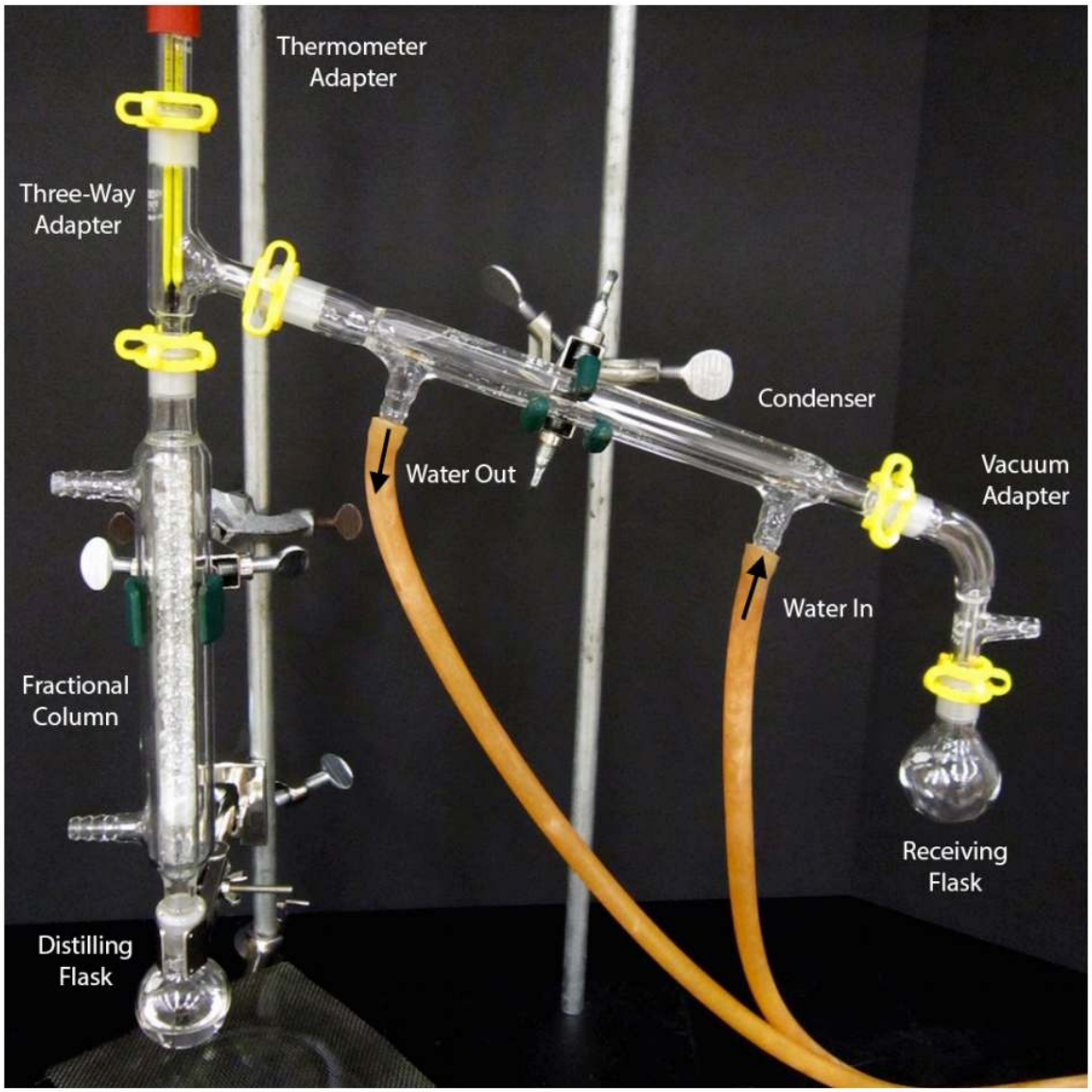
Different fractionating columns: a) Vigreux, b) Steel wool, c) Glass beads

These columns have different surface areas and numbers of theoretical plates, and thus differ in their ability to separate close-boiling components. A Vigreux column has the least surface area, making it the least capable of separating close-boiling components. And yet with the lowest surface area, it can have the highest recovery, making it the optimal choice if a separation is not particularly difficult. Glass beads have a high surface area, so are a good choice for separation of close-boiling components. And yet, beads will suffer the greatest loss of material. Steel wool columns have intermediate surface areas and their effectiveness can depend on how tightly the wool is packed in the column. Steel wool columns also cannot be used with corrosive vapors like those containing acid.

Adapted from *Theory of Fractional Distillation* and *Fractionating Columns* by Lisa Nichols.

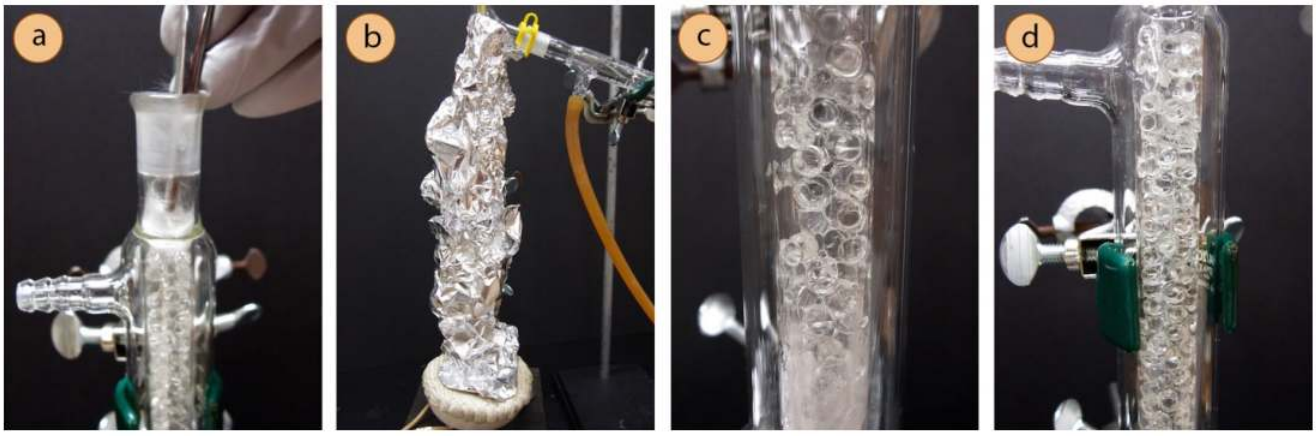
9.5: Fractional Distillation Procedure

An assembled fractional distillation apparatus is shown below, using glass beads in the fractionating column. Other columns may be substituted. It is assumed that readers have previously performed a simple distillation, so in this section are described differences between simple and fractional distillation.



Fractional distillation apparatus

1. The distilling pot will need to be heated much more vigorously than with a simple distillation, as there is a greater distance for the vapors to travel before reaching the condenser. The vapors will tend to reflux in the column (condense and drip back into the distilling pot) unless stronger heating is applied. If it is difficult to achieve more than a reflux, the column can be insulated by wrapping it with glass wool. This allows the column to maintain heat and the sample to remain in the gas phase longer. A small gap can be left in the glass wool to "peek in" on the activity in the column.
2. Ideally both liquid and gas should be seen in the fractionating column, as the sample needs to undergo many vaporization-condensation events. Droplets of liquid *should* be seen on the surfaces of the packing material, but there should never by a large pool of liquid. A "river" of liquid traveling up the column is called **flooding**. If a column floods, remove the heat until the liquid drains back into the distilling flask, then resume heating at a gentler rate.



a) Removal of glass wool plug on a beaded fractionating column, b) Insulating the column with foil, c+d) Condensation on the beads of a fractionating column

Flooding of a fractionating column

3. Cleaning of a fractionating column:

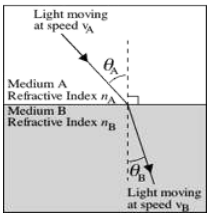
- Vigreux column: rinse with acetone. Don't use a scrub brush or the glass indentations may break.
- Steel wool column. rinse with large amounts of acetone. Don't rinse with water as wet steel will rust over time.
- Glass bead column: rinse with acetone, then place a small piece of glass wool wad to prevent the beads from pouring out when horizontal.

Adapted from *Step-by-Step Procedure for Fractional Distillation* by Lisa Nichols.

CHAPTER OVERVIEW

10: Refractive Index

The refractive index is a ratio of the speed of light in a medium relative to its speed in a vacuum. This change in speed from one medium to another is what causes light rays to bend. It is also important to note that light changes direction when it travels from one medium to another.

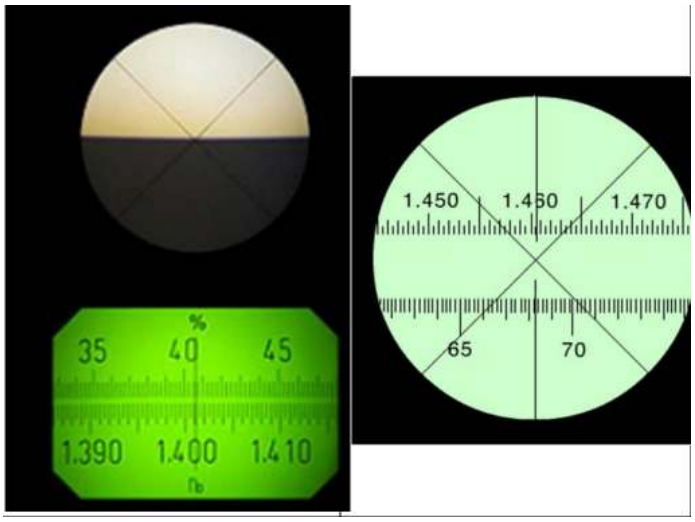


Refractive indices have many purposes and are used most frequently to differentiate between liquid samples. Therefore, this physical quantity characterizes liquids in the same way that melting points are used to characterize solids. This measurement can serve as a means of identification of a substance by comparing its refractive index to known literature values. Refractive indices can also be used to as an estimate of the purity of a compound by comparing the substance's refractive index to that of the pure compound.

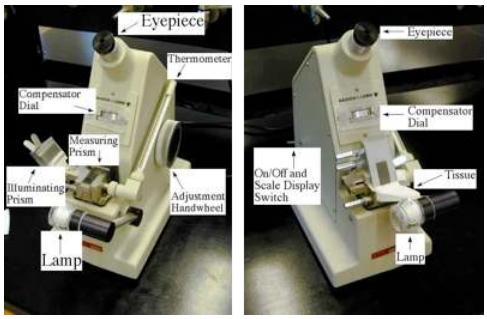
Two factors affect the value of the refractive index are:

1. **Temperature**: Refractive index values are usually determined at standard temperature.
2. **Wavelength of light**: The refractive index varies with wavelength. Only a specified wavelength should be used when measuring the refractive index.

A refractometer is used to measure the refractive index of a medium. A light source shines on the illuminating prism and light rays enter the sample moving in different directions. A detector at the back of the refracting prism produces the light and dark regions. Samples with different refractive indexes produce different angles of refraction which will cause a shift in the borderline between the light and dark regions. The borderline's position is then used to establish the refractive index of different samples by viewing an illuminated scale.



Refractive Index Lab Procedure



Refractometer

1. Open prism assembly and inspect the prism for cleanliness. If necessary, clean with acetone and a Kimwipe.
2. Place a few drops of the liquid sample on the lower prism using a pipette and close the prism.
3. Move the illuminator arm upwards.
4. Turn the adjustment control until the lower field appears dark and the upper field light.
5. Press contact switch at left side of instrument. Reading the top scale, estimating to the 4th decimal place.
6. Take a temperature reading from the attached thermometer.
7. Clean the prisms by wiping the sample with a lens tissue, followed by cleaning with acetone.

Temperature Correction

Since refractive indices vary with temperature, for each degree Celsius of temperature change, the average temperature correction has been found to be 0.00045 units for a wide range of compounds. The following equation can be used to determine the refractive index at 20.0ºC:

$$n_D^{20} = n_D^T + Y \quad \text{where} \quad \mathbf{Y = (T - 20°C) * 0.00045 / °C}$$

Ex. If the reading for the refractive index was 1.4370 at 18ºC, it would be corrected to 1.4370- 0.0009 = 1.4361 at 20ºC.

Refractive indices are reported in literature at 20.0ºC. In order for comparisons to be made between the literature and experimental refractive index, the experimental refractive index must be corrected to 20ºC.

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