Extraction

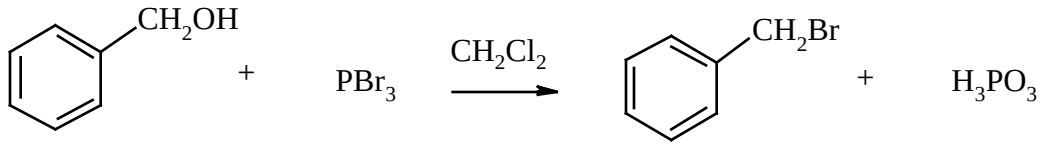
W. H. Bunnelle, L. A. Meyer, R. E. Glaser

Introduction

Chances are, everyone in this class has done an extraction, probably several of them. Ever make a cup of tea or a pot of coffee? Anything but instant, and you've done extraction. Simply put, extraction involves mixing of mutually insoluble materials, where a component of one of the phases moves into the other. To make a cup of tea, hot water is mixed with dried tea leaves. The leaves are not soluble in the water, so there are two immiscible phases, the solid and the water. However, the water dissolves some of the compounds originally present in the tea leaves - these materials are extracted into the water. On the other hand, the preparation of a cup of instant coffee is not an extraction at all. Instant coffee powder is completely soluble in the water, and so there is no second phase. This is simply a process of solution.

Extractions are often classified according to the nature of the phases involved. The water-tea leaves system is an example of liquid-solid extraction, and one can think of many such examples where components migrate from a solid phase to a liquid phase and *vice versa*. In this experiment, we will be concerned with liquid-liquid extraction, a very common laboratory operation which can be used as a separation or purification technique.

Consider the following reaction:



This is a fairly standard organic reaction, involving the conversion of an alcohol to an alkyl halide. For now, imagine that we have the crude reaction mixture, which contains benzyl bromide (the organic product) and phosphorous acid in dichloromethane solvent. How do we separate these to get the pure product? It turns out that a very simple way to separate the phosphorous acid from the organic material is to add water. The dichloromethane solvent is immiscible with water, and separates as a second layer (the organic phase). Now the molecules of benzyl bromide can stay in the organic phase, or migrate to the aqueous phase. What property will determine where these molecules go? The tendency for a compound to reside in a solution is, of course, measured by its solubility. Benzyl bromide has a very much greater solubility in dichloromethane than in water, and so essentially all of this compound remains in the organic layer. For the phosphorous acid, on the other hand, the situation is just the opposite. As a very polar, ionized material, H_3PO_4, is much more soluble in water than in dichloromethane. Consequently, the phosphorous acid molecules will migrate out of the dichloromethane phase into the aqueous layer, much like those water-soluble constituents of tea leaves. Now, if the dichloromethane layer can be removed and the aqueous layer left behind, we will achieve, in a very simple way, a separation of the products of the reaction. Quite obviously, extraction will only work for separation of materials which have very different solubilities. It is a technique used most often for rudimentary purification of crude reaction mixtures.

The device used to carry out this separation is called a separatory funnel (see Fig. 1). It does resemble a funnel, if you use your imagination a bit. The glass body has a conical shape like a funnel, but this one closes at the top to a ground-glass joint, which can be shut with a stopper. The stem of the separatory funnel incorporates a stopcock, so that flow out of the bottom of the funnel can be interrupted. A separatory funnel is well-designed for its function. The larger volume at the top of the funnel is helpful for mixing the phases when the funnel is inverted, and the tapering lower portion of the funnel permits clean separation of the phases.

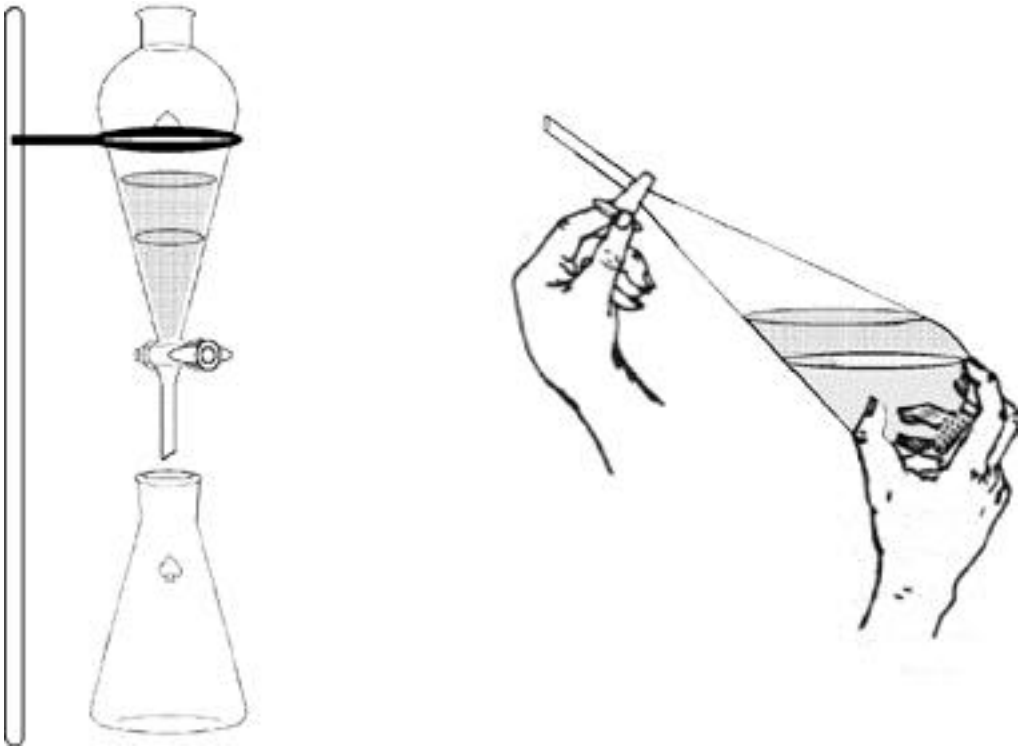


Figure 1. The separatory funnel in the ring stand (left), and venting the funnel (right).

How to Use a Separatory Funnel

Operation of a separatory funnel is not difficult, but there are several potential pitfalls which must be avoided. The first order of business is to get the two liquid phases in the funnel. Make sure the stopcock is closed! This may be obvious, but every year, without fail, some students forget or don't pay attention, and wind up with a mess on the benchtop, and precious little product. Just in case, double check. The funnel is suspended in an iron ring -- choose one that allows free passage of the stopcock, and catches the separatory funnel about 1/2 - 1 inch below its widest part. The ring should be clamped high enough on a ring stand so that a collection vessel (usually an Erlenmeyer flask) will fit beneath the lower tip of the funnel when it is in place in the ring.

Use a funnel to fill the sep funnel. It is important not to contaminate the ground joint at the top. If you get some grit or insoluble material on the joint, it will not seal tightly when you stopper it, and will leak all over your hands when you invert the funnel for mixing. By the way, you should make sure that your stopper and stopcock don't leak anyway. Before you try a 'real' extraction, take some water into the sep funnel, shake it around, and look carefully for leaks. If it does, check with your TA for help - you may have a misfit stopper or stopcock.

Now, with both phases together, stopper the funnel and remove from the iron ring. With one hand cupped over the stopper end of the funnel (and holding the stopper in place between the fingers), and the other hand cradling the stopcock, gently invert the separatory funnel so that the contents run back to the fat end of the funnel, leaving the stopcock end empty. Now, with the sep funnel held at about a 45° incline, and with the stopcock end facing away from you and away from your lab neighbors, open the stopcock. You will hear a pfffft, as excess pressure vents from the sep funnel. Where does this pressure come from? There are many possibilities, depending on the extraction. In some cases, a gas (CO_2) is evolved, in others the warmth of your hands is enough to vaporize some volatile solvent, causing a pressure increase. Often there is a small heat of mixing which causes the same thing. This can go both ways, in fact, sometimes a vacuum is created, and the pfffft goes inward (tffffp). Regardless, it is very

important to vent the separatory funnel repeatedly during use. Since the pressure increase can be substantial, especially when CO_2 is generated in the mixing, you should always begin cautiously - one gentle inversion of the sep funnel and vent. Next, close the stopcock, and gently rock the funnel back and forth one or two times. Vent again. Continue this process for several cycles, until you are sure that the phases have mixed thoroughly. Vent frequently, for pressure buildup in a sep funnel can be very dangerous. The excess pressure will force liquid out through the stopper onto your hand, or, even worse, may cause the funnel to burst. Of course, you must make certain that the narrow neck of the funnel is completely drained before opening the stopcock, the pressure will squirt any remaining liquid out through the stopcock.

The object is to ensure good mixing of the layers, so that solutes can get from one phase to the other. Always start the mixing very gently, with frequent venting. The mixing may gradually become more vigorous, from rocking to sloshing to shaking, but never for more than a few seconds at a time, and always with frequent and repeated venting. It is a good idea, before moving to the next level of mixing, to replace the sep funnel in the ring stand (close the stopcock!), and watch for the separation of layers. If the phases separate rapidly into two layers with a distinct boundary, it's OK to shake harder. Too-vigorous mixing may cause an emulsion to form. If this happens, and you get a cloudy or milky looking mixture which does not separate, you've got trouble. The best way to deal with emulsions is not to get them in the first place. They can usually be avoided if proper care is taken during sep funnel mixing, and close watch is kept to detect the first small amounts of emulsion. If, despite these precautions, an emulsion still forms, don't despair - there are some tricks which may help. See breaking emulsions in "Other Aspects of Extraction".

When the mixing is completed, the sep funnel should be vented one last time, the stopcock closed, and the funnel returned to the ring stand. The stopper is then removed (REMEMBER!), and the funnel allowed to stand until the layers separate. This means a clear, distinct separation of phases, with a visible meniscus between the layers. There will sometimes be some insoluble material in the vicinity of the dividing line between the phases, and as long as it doesn't amount to much, this can be ignored. The separation is completed by draining the lower layer out through the stopcock, into an appropriately sized Erlenmeyer flask. This should be done slowly. In no case should a vortex form. If it does, the funnel is being drained far too rapidly, and material from the upper layer will be drawn through the stopcock along with the lower layer. Draining should be continued until the interface is 5 -10 mm above the stopcock barrel, and the separatory funnel allowed to stand for a few more minutes. During this time, the lower phase which adhered to the walls of the separatory funnel will slide down and collect at the narrow base of the funnel. A gentle swirling of the sep funnel will facilitate this fall-out. The accumulated material can then be drained carefully, so that the stopcock is closed just as the interface reaches it. If there is a small amount of emulsion or insoluble material at the interface, this is best taken with whichever phase will be extracted further; often the emulsion will break with further extraction and the insoluble material can be removed by filtration later on.

The lower phase has now been removed, and the upper phase remains in the sep funnel. If further extraction of this phase is necessary, a fresh portion of the lower phase can be added, and the mixing and draining process repeated. Eventually, the upper phase should be removed from the sep funnel by pouring it out through the upper neck. The upper phase should not be drained through the stopcock - this would contaminate it with material from the lower phase which may adhere to this part of the glassware. Obviously, if further extraction of the lower phase is required, it can be added to the empty sep funnel, fresh upper phase added, and the extraction operation carried out.

Other Aspects of Extraction

Now, what kinds of solvents do we use? Almost without exception, one of the phases is water, or some aqueous solution. The other is a water-immiscible solvent, for example ethyl ether, dichloromethane, hexane, etc. The important thing is that the solvent is not miscible with water. If it is, you will get only one layer, and then there is nothing to separate. Solvents which are more dense (e.g, dichloromethane) or less dense (e.g., ether) than water can be used; thus, the aqueous phase may turn out to be the upper or the lower layer, depending on the circumstances. Since you will usually want one of the layers, because it contains your product, and will want to discard the other, since it contains impurities, it is critical that you

know which layer is which. Now, if you know the densities of the solvents you are using (something to look up for your lab notebook), you will be able to predict which layer is on top, and which is on the bottom. You should be aware, however, that the densities listed in references like the CRC Handbook of Chemistry and Physics (printed or online) are for the pure solvents. We are dealing here with solvents which contain varying amounts of solutes, and these can change the density of the solution. Therefore, to make certain that you know the identity of a particular layer (organic phase or aqueous phase), the following simple test should be carried out: one or two drops of the suspected organic phase is added to 1 ml of water in a clean test tube. If a homogeneous solution is obtained on mixing, then this was probably the aqueous layer. To make sure, carry out the same test on the other layer - the organic phase should separate from water (or at least form a cloudy mixture). Finally, you can avert potential disaster by always saving both layers until you have finished the experiment and know that you have the right stuff. In that case, even if you make a mistake, you will have the other extraction layers, and will be able to recover your product.

Now, what about emulsions? An emulsion consists of tiny droplets of one phase suspended in another. If the droplets are small enough, they can remain suspended for a long time, although eventually the mixture should separate into two layers. Therefore, one remedy for emulsions is to wait a little while. Lab time is limited, though, so if substantial separation hasn't occurred within 5 minutes, try something else. A trick which occasionally works is to add some saturated sodium chloride solution, and/or additional organic solvent. These things are easy to try, and if they work, great. But don't expect too much, and don't waste too much time if these don't seem to work. In my experience, the most sure-fire method for dealing with emulsions (if such a thing exists) is to filter the whole mixture through a bed of crushed diatomaceous earth (trade name: Celite®). Vacuum filtration is used - the mixture is poured onto a thin (~1/2") layer of Celite® in a Buchner funnel. Quite often, the filtrate will separate nicely after this treatment. The Celite® filtration removes microscopic, insoluble particles which sustain emulsions. When these are removed, the emulsion breaks down. If this trick doesn't work, you have my sympathy. Better start over, and this time, don't shake so hard!

The extraction operation is sometimes called 'washing.' This is really a matter of semantics, since both involve the same process. The distinction between extraction and washing is whether the immiscible solvent is added to remove the desired material (extraction) or to carry away impurities (washing). For example, the treatment of the reaction on page 1 is properly described as washing, since water was added to remove phosphorous acid (the impurity) from the dichloromethane solution containing benzyl bromide. If, however, our goal had been to isolate phosphorous acid, that would constitute an extraction.

Although dichloromethane appears to be totally immiscible with water, in fact each of these dissolves to a small extent in the other. This is likewise true for other organic solvents. The organic phase always contains a small amount of dissolved water. This should be removed so that it does not interfere with further purification of the product, for example by distillation or recrystallization. The water can be removed by addition of an appropriate 'drying agent' to the organic solution. These are anhydrous inorganic salts which readily absorb water into their crystal lattice to form hydrates. The driving force for hydrate formation is pretty substantial, and these salts will soak up nearly all of the water in the organic solution. Since the drying agent is not soluble in the solvent, it can then be removed simply by filtration. Typical drying agents are anhydrous calcium chloride ($CaCl_2$), anhydrous magnesium sulfate, and anhydrous sodium sulfate. There are others, but these are probably the most common. It is important to note that only the anhydrous forms of these salts are useful as drying agents. The hydrated forms already have their water - why would they pick up more?

The drying of organic solutions is done in an Erlenmeyer flask. A small amount of the drying agent (depends on the volume of solution, but a level teaspoon for 50 ml of solvent is a reasonable start) is added, and the mixture swirled gently. The initially free-flowing granular or powdered salt will begin to clump together - this is a sign of saturation with water. If some of the solid remains free-flowing, there is enough drying agent. Just let the mixture stand for about five minutes so that absorption of water is complete. If all of the solid clumps, additional portions of drying agent should be added, until unclumped solid remains. Large excesses of drying agent will absorb product, and should be avoided. You should not attempt to dry an organic solution if drops of water are visible. This will require too much drying agent. Instead, the mixture should be put back in the sep funnel, and the bulk water removed before

drying. It goes without saying that you cannot dry an aqueous solution. Make sure you are dealing with the right layer!

When enough drying agent has been added, and the drying has been allowed to proceed for ~5 min, the salt is removed by filtration. Normal gravity filtration is fine, but you will save time by using a fluted filter paper. Plan ahead - if the next step is to distill the solvent, use the proper round-bottom flask to collect the filtrate. Finally, in order to ensure complete transfer of product, the drying flask and filter should be rinsed with a small amount of solvent.

Theory of Extraction

Suppose we add a solute, compound **A** to a mixture of two mutually immiscible solvents. When **A** dissolves, where does it go? Depending on the nature of **A** and the respective solvents, **A** might be found mostly in one phase or the other, or perhaps in both. In general, a solute like **A** will be distributed, or partitioned between both available phases. An equilibrium distribution will be reached, and the equilibrium constant for this process is called the distribution coefficient, or sometimes the partition coefficient. If we assume that water is one of the solvents, and a generic organic solvent is the other, the distribution coefficient is given by:

$$K_{org/water} = [\mathbf{A}]_{org}/[\mathbf{A}]_{water}$$

The distribution coefficient is a measure of the tendency for the solute to reside in one phase versus the other and, to a very good approximation, is equal to the ratio of solubilities for **A** in the respective solvents. Thus, if **A** is much more soluble in water than in the organic solvent, the distribution coefficient will be small (<1), while a large distribution coefficient implies that **A** is much more soluble in the organic than in water. You should note that it is the concentration in each phase which is important, so that the quantity of **A** which is found in each phase will also depend on the volumes of solvent involved.

Those compounds with very large (say, >20) or very small (<0.05) distribution coefficients will be found nearly totally in one of the phases. In contrast, for compounds with intermediate values for the distribution coefficient, the distinction between phases is not so clear-cut. Significant amounts of material will be found in each phase. This raises the question of extraction efficiency. Suppose we begin with an aqueous solution of material we want to extract into ether. What fraction of the material can be extracted for a given volume of ether? Will more compound be extracted if all of the ether is used in one single extraction, or is it better to divide the ether into smaller portions and do multiple extractions of the aqueous phase? In fact, the latter course of action will provide significantly better recovery of the compound, as the following calculations will illustrate.

Let's assume, for **A**, that $K_{ether/water} = 4$. Now, if we have an aqueous solution of **A** (10 g) in 100 ml of water, how much **A** will be recovered by extraction with (a) 100 ml of ether in one extraction, or (b) 2 extractions of 50 ml each.

Begin with case (a):

$$K_{ether/water} = [\mathbf{A}]_{org}/[\mathbf{A}]_{water}$$

$$K_{ether/water} = \frac{\text{g } \mathbf{A} \text{ in ether / ml of ether}}{\text{g } \mathbf{A} \text{ in water / ml of water}} = 4$$

(Notice that the particular concentration units used are unimportant, as long as the same units are used in the numerator as in the denominator.) Now the volume of ether and the volume of water are both 100 ml, and substituting, we have:

$$\frac{\text{g } \mathbf{A} \text{ in ether}}{\text{g } \mathbf{A} \text{ in water}} = 4$$

Since: (g **A** in ether) + (g **A** in water) = 10 g, we have two equations and two unknowns, and this can be solved:

g **A** in ether = 8 g,

or 80% of the compound is extracted into the organic phase with a single 100 ml extraction.

For case (b), involving extraction of the same aqueous solution with two-50 ml portions of ether, we carry out the same kind of calculation for each extraction. For the first portion of ether:

$$\frac{\text{g } \mathbf{A} \text{ in ether / ml of ether}}{\text{g } \mathbf{A} \text{ in water / ml of water}} = 4$$

For this example, (ml of ether) = 50 ml, and (ml of water) = 100 ml, so:

$$\frac{\text{g } \mathbf{A} \text{ in ether}}{\text{g } \mathbf{A} \text{ in water}} = 2$$

Again, the total amount of **A** is 10 g, and solving the simultaneous equations yields:

g **A** in ether = 6.67 g

Now, if this much **A** was extracted into the ether phase, 3.33 g must remain in the aqueous phase. For another extraction with the 2nd 50 ml portion of ether, we have again:

$$\frac{\text{g } \mathbf{A} \text{ in ether}}{\text{g } \mathbf{A} \text{ in water}} = 2$$

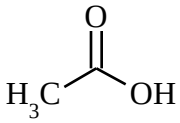
but for this second extraction the total amount of **A** is 3.33 g, and the equations can be solved to give:

g **A** in ether (2nd extract) = 2.22 g

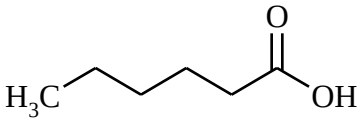
and the total extracted by the 2 - 50 ml portions of ether is 6.67 + 2.22 = 8.89 g, or 89 % of the available **A**. Comparison with the results obtained for case (a) clearly demonstrates that a more efficient extraction is obtained by the multiple extraction. In fact, one can carry out similar calculations for three 33.3 ml extractions, or four 25 ml extractions, and the efficiency would be still greater. Of course, some practical limit is reached beyond which the increased efficiency is not worth the extra effort required to carry out so many extractions. Somewhere between two and four extractions seems to be a reasonable compromise.

1-Octanol / Water Distribution Coefficients

The distribution of a solute between two immiscible phases is a very general phenomenon, and occurs in more places than a separatory funnel. This partitioning process is involved in a variety of biological and environmental processes including soil adsorption, bioconcentration, and lipophilic storage. The mobility of a molecule in a biological system depends on its ability to move through polar aqueous



acetic acid



hexanoic acid

media, as well as very non polar environments such as are found in cell walls. It is convenient to characterize materials as 'hydrophilic' or 'lipophilic,' depending on whether they are more soluble in aqueous or in non-polar environments, respectively. In order to develop models which will permit reliable predictions of the distribution of compounds in the environment or in a cell, a quantitative measure of lipophilicity is needed. One system which has been very useful is the 1-octanol / water distribution coefficient. 1-Octanol is not miscible with water, and serves as a mimic of the relatively non polar environments such as cell walls. The 1-octanol / water distribution coefficients have been determined for many compounds, and turn out to be a reasonably reliable indicator of lipophilicity for those compounds. We will determine the distribution coefficients for two carboxylic acids, acetic acid and hexanoic acid. The acids are chosen because their concentrations can be easily determined by titration. You will determine the quantity of acid in an aqueous solution before extraction with 1-octanol, and compare that to the amount of acid present in the aqueous phase after the extraction with 1-octanol. The difference represents the portion of acid present in the 1-octanol phase after the extraction, and this is all the information you need to determine the distribution coefficient. The amount of acid (in moles) present in an aliquot of stock solution is given by:

$$\text{moles(acid)} = \text{ml(NaOH)} \cdot \text{M(NaOH)}$$

where ml(NaOH) is the quantity of stock solution required to titrate the acid to the phenolphthalein endpoint. An identical aliquot is then extracted with 1-octanol, and the aqueous layer titrated as above. The amount of acid in the aqueous phase (in moles) is given similarly:

$$\text{moles(aq)} = \text{ml*(NaOH)} \cdot \text{M(NaOH)}$$

where ml* refers to this second titration. Then the quantity of acid in the octanol is given by:

$$\text{moles(oct)} = \text{moles(acid)} - \text{moles(aq)} = [\text{ml(NaOH)} - \text{ml*(NaOH)}] \cdot \text{M(NaOH)}$$

Finally, the distribution coefficient is given by the following expression:

$$K_{oct/aq} = \frac{\text{moles(oct)} / \text{vol(oct)}}{\text{moles(aq)} / \text{vol(aq)}} = \frac{[\text{ml(NaOH)} - \text{ml*(NaOH)}] / \text{vol(oct)}}{[\text{ml*(NaOH)}] / \text{vol(aq)}}$$

where vol(oct) and vol(aq) are the volumes of the octanol and aqueous phases, respectively, used in the extraction. Note that the NaOH concentration term drops out, and the measurement is independent of the concentration of base used for the titration.

The distribution coefficients for both acetic acid and for hexanoic acid will be determined, and comparison of the two will permit some conclusions regarding the influence of structure on lipophilicity. Can you make a prediction?

Extraction With Acids and Bases

Hexanoic acid is more soluble in 1-octanol than it is in water, and so most of the hexanoic acid will go into the organic phase. In contrast, if aqueous NaOH is used in place of water, very little of the hexanoic acid remains in the organic layer; most of it moves to the aqueous phase. These results imply that hexanoic acid is more soluble in aqueous base than it is in water alone. Why should this be? Actually, the answer is not that the acid is more soluble in base, but that it reacts with the base to form a more water-soluble material. As is typical of acid-base combinations, the reaction of hexanoic acid with NaOH produces a salt:

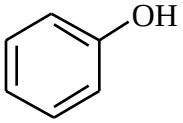
$$CH_3CH_2CH_2CH_2CH_2CO_2H + NaOH \rightarrow CH_3CH_2CH_2CH_2\,CH_2CO_2^-\,Na^+ + H_2O$$

The salt, being an ionic compound, is much more soluble in aqueous solution than it is in the relatively non-polar organic solvent. So you see, we can manipulate the partitioning of an acid between aqueous and organic phases by adjusting the pH of the aqueous phase.

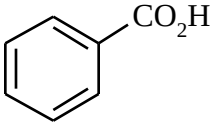
A similar situation holds for organic bases, the amines. These are derivatives of ammonia where one or more of the hydrogens of NH_3 have been replaced by alkyl groups. Like ammonia, amines are basic, and react with acids to form ionic salts:

$$CH_3CH_2CH_2CH_2NH_2 + HCl \rightarrow CH_3CH_2CH_2CH_2NH_3^+\,Cl^-$$

In this way, amines which are normally much more soluble in organic solvents than water can be easily extracted in aqueous acid solutions, because the amine is then converted to a water-soluble salt. Therefore, it is possible to selectively remove acidic or basic organic compounds by extraction with appropriate basic or acidic aqueous solution, and this is an easy method for the separation of such mixtures. Some fine-tuning of the technique is made possible by controlling the pH of the aqueous solution. For example, benzenols (phenols) are weaker acids than are carboxylic acids, and are only



phenol pKa = 10

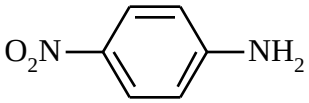


benzoic acid pKa = 5

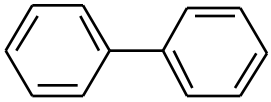
converted to salts with relatively strong bases (NaOH). Carboxylic acids, on the other hand, are converted to the salt even with weaker bases such as $NaHCO_3$. Thus, the benzoic acid can be selectively removed by extraction into an aqueous solution of bicarbonate; after the stronger acid has been removed, the phenol can be extracted by use of NaOH solution. Extractions with $NaHCO_3$ (and Na_2CO_3) are often used to remove acidic materials, but particular care must be taken to vent the separatory funnel early and often during the extraction. The reaction of acid with these bases generates CO_2 gas, and the pressure buildup in a closed separatory funnel can be dangerous.

$$\underset{\text{(acid)}}{H\text{-}A} + NaHCO_3 \rightarrow Na^+\,A^- + H_2O + CO_2\,(g)$$

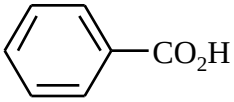
These points will be illustrated in the laboratory by the use of extraction to separate a mixture of three compounds. The solid mixture consists of varying amounts of p-nitroaniline, biphenyl, and benzoic acid. The first of these is an amine derivative, and therefore basic. Biphenyl is a typical organic hydrocarbon, and is neither acidic or basic, while benzoic acid, as already discussed, is an acid. The mixture will be separated according to the scheme outlined below.



p-nitroaniline

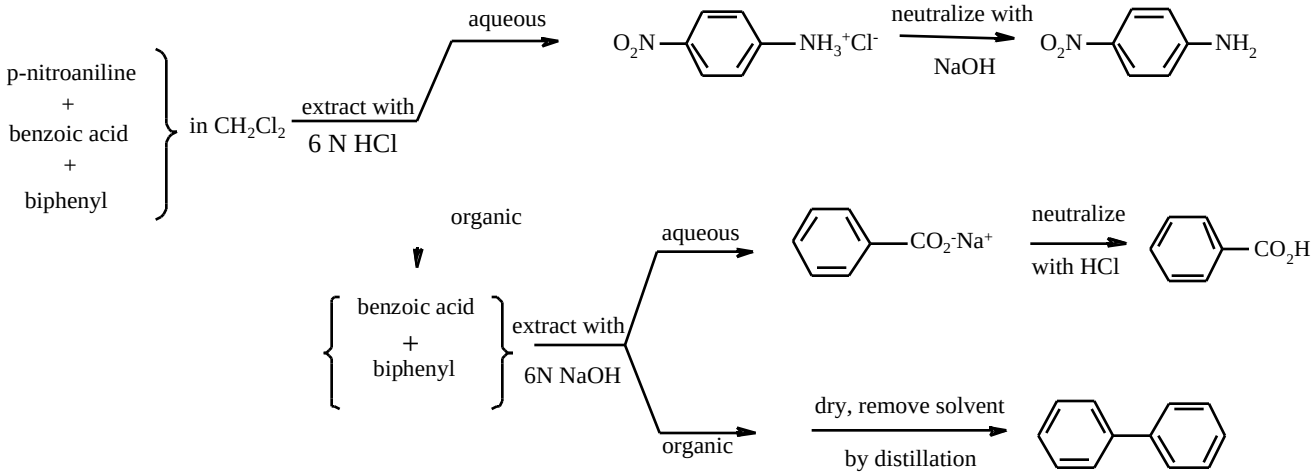


biphenyl



benzoic acid

The mixture is dissolved in an organic solvent, and extracted first with aqueous HCl. The HCl combines with p-nitroaniline to form the amine hydrochloride salt, which moves to the aqueous phase. Separation of the layers provides an aqueous solution of the amine salt, and an organic solution of benzoic acid and biphenyl. The p-nitroaniline can be recovered from the aqueous solution by neutralizing the acid. This converts the amine salt back to the amine, which is not water-soluble, and precipitates. The amine is then collected by vacuum filtration (see Expt. 1). Meanwhile, the next step of the separation is to extract the organic phase with aqueous NaOH. The benzoic acid is converted to its sodium salt, and moves to the aqueous layer. On separation of the phases, one obtains an aqueous solution of the sodium benzoate, and the organic solution still containing biphenyl. The benzoic acid is recovered by acidifying the aqueous solution, which causes the benzoic acid to precipitate, whereupon it is isolated by filtration. The biphenyl is isolated by drying the aqueous solution to remove dissolved water, and finally distillation of the solvent - the biphenyl will remain as a non-volatile residue. The whole operation is accomplished simply in a short period of time using these principles of extraction. The three components are isolated in nearly pure form; minor contaminants could be removed, if desired, by recrystallization.



Experiment 2

<u>Part 1</u>: The distribution coefficients for partitioning of acetic acid and hexanoic acid between 1-octanol and water will be determined. Note: Volume measurements for this part of the experiment should be made as accurately as possible.

Obtain a 25 ml aliquot of the stock acetic acid solution (~0.02 M) and place it in a 125 ml Erlenmeyer flask. Add two drops of phenolphthalein indicator and titrate to the phenolphthalein endpoint using the dilute NaOH solution provided (~0.02 M); record the titration data. Obtain a second 25 ml aliquot, placed it in a 125 ml separatory funnel, and extract with 25 ml of 1-octanol. The aqueous layer is then separated, and titrated to the phenolphthalein endpoint with the same NaOH solution used earlier. Note that this titration will require less NaOH solution than the first one. Go slowly and don't overshoot the endpoint, or you will have to repeat the extraction step. From the titration data, and the volumes of the phases used in the extraction, the distribution coefficient can be determined. Refer to the example calculation on p 8. The neutralized aqueous solutions (from the titration) can be disposed down the drain, and the octanol should be poured in the appropriately labeled waste bottle.

Repeat the procedure, this time using two aliquots of hexanoic acid solution. This one can be done before the acetic acid, it makes no difference at all. Do make certain, however, that the separatory funnel has been rinsed thoroughly with water, and allowed to drain as completely as possible before carrying out the extraction. Caution: the titration of the aqueous phase from the hexanoic acid extraction requires very little titrant. Be careful or you will miss the endpoint. From the data, determine the octanol/water distribution coefficient for hexanoic acid. Compare to that for acetic acid.

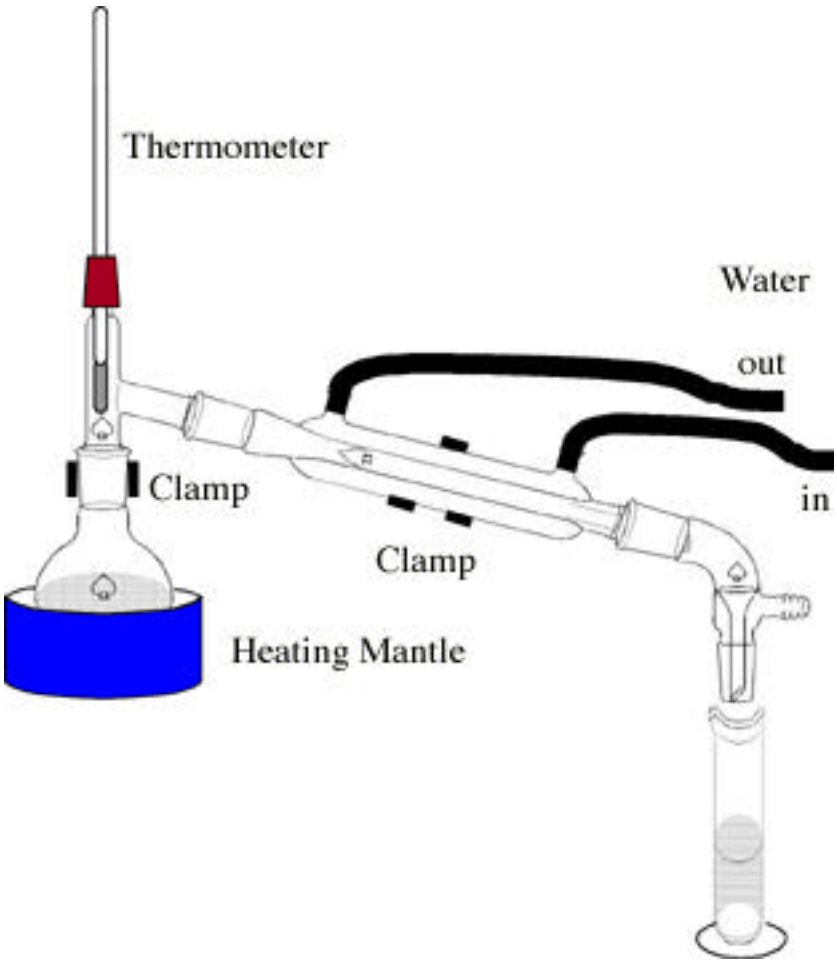


Figure 2. Simple Distillation Apparatus.

<u>Part 2</u>: A mixture of organic compounds will be separated by extraction, taking advantage of their different acid-base properties.

<u>Safety</u>

Relatively concentrated acids and bases are used in this experiment. Always be careful to point the tip of the funnel away from everyone when venting. A leaky separatory funnel will spill acids, bases and organics all over your hands. Make sure that yours does not leak! Extra care should be taken to avoid spilling these materials. A pair of rubber gloves will offer some protection. Dichloromethane is toxic, and has been categorized as a carcinogen. p-Nitroaniline is toxic, and can be readily absorbed through the skin. In case of exposure, wash your hands thoroughly with soap and water.

Weigh 3.0 g of the p-nitroaniline-benzoic acid-biphenyl mixture into a 125 ml Erlenmeyer flask, and dissolve in 40 ml of dichloromethane (this may require gentle warming on a hot plate). Pour the solution into a separatory funnel (use a funnel, make sure that the stopcock is closed), and rinse the flask with an additional 10 ml of dichloromethane. Add 15 ml of 3 N HCl, stopper the sep funnel, and mix the phases, venting the funnel frequently. Allow the layers to separate, and draw off the lower layer into the 125 ml Erlenmeyer. Determine which of these is the organic phase (see "Other Aspects of Extraction"). Set the aqueous phase aside in a labeled Erlenmeyer flask, and extract the organic phase one more time with 15 ml of 3 N HCl. Combine the aqueous layer with that from the first extraction. Repeat the process, extracting the organic phase twice with 15 ml portions of 3 N NaOH solution. The aqueous layers from the base extraction are combined, but kept separate from those from the acid extraction. The organic phase which remains after the extractions is dried by addition of anhydrous Na_2SO_4, one teaspoonful at a time, with swirling of the flask until some free-flowing (non-clumping) salt is observed. The organic solution is filtered (fluted filter paper) into a 100 ml round-bottom flask, and the flask and funnel rinsed with ~10 ml of dichloromethane. Set up for simple distillation (see Fig. 2). Make sure that all joints are fitted tightly and use a heating mantle as the heating source. Heating mantles are plugged into controllers and <u>never</u> plugged directly into the outlet. Distill the dichloromethane, collecting the solvent for disposal in the proper waste container. After all of the solvent has been removed, the boiling flask is allowed to cool, and the solid product (biphenyl) is removed and weighed.

Meanwhile, the HCl extracts should be cooled in an ice-water bath, and neutralized by addition of 3 N NaOH solution, until the mixture tests basic with litmus or pH paper (red litmus turns blue in base). The resulting mixture is chilled to complete precipitation, and the solid isolated by vacuum filtration (see setup on exp. 1). Wash the solid on the Buchner funnel with cold water, and dry by drawing air through the solid for 5 minutes, pressing with a clean cork to assist the removal of water. The dry solid is removed from the funnel and weighed.

Similarly, the NaOH extracts should be cooled and neutralized by addition of 3 N HCl, until the mixture is acidic (check with litmus or pH paper). The resulting precipitate is isolated by vacuum filtration (clean up the filtration apparatus from the first filtration before you carry out this one), washed with cold water, and suctioned as dry as possible. Weigh the dry solid, and calculate the percent recovery from the starting mixture for each of the three compounds.

The solids will be reused - please put these in the appropriate recovery jars.

Post-laboratory Questions

1. Why use two 15 ml portions instead of one 30 ml volume for the acid and base extractions in Part 2?

2. Compare the octanol/water distribution coefficients for acetic acid and hexanoic acid. Rationalize the difference.

3. Write equations for the acid-base reactions involved in Part 2 of this experiment.