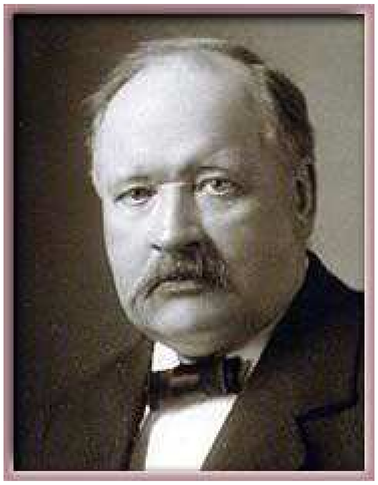
Acid Base Diary

At one time acids and bases were just solutions, with no chemical understanding. In the late 1880's a man in Sweden named Svante Arrhenius developed his Arrhenius Theory of Acids and Bases. It covers nearly 99% of all acids and bases in chemistry. He won the Nobel Prize in Chemistry in 1903 for this. That's him at right.



He stated that an acid was a substance that had excess hydrogen ions in aqueous solution, and that a base had excess hydroxide ions in solution. Further, that acids plus bases will be able to neutralize each other into salty water.

A list of acids you need to be familiar with is on Table K in your reference table. They are listed in strength from top to bottom.

$HCl_{(AQ)}$ hydrochloric acid	dissociates very well	LOTS of H^+	The stronger the acid, the more H^+ ions (and more other ions as well). The more ions, the better it conducts electricity. Strong acids are good electrolytes. Weak acids are poor electrolytes. Acetic acid is written "normally" and as an ORGANIC acid as well. Both are the same thing, but with different naming styles.
$HNO_{2(AQ)}$ nitrous acid	dissociates very well	LOTS of H^+	
$HNO_{3(AQ)}$ nitric acid	dissociates very well	LOTS of H^+	
$H_2SO_{3(AQ)}$ sulfurous acid	dissociates very well	Lots of H^+	
$H_2SO_{4(AQ)}$ sulfuric acid	dissociates well	many H^+	
$H_3PO_{4(AQ)}$ phosphoric acid	dissociates less well	less H^+	
$H_2CO_{3(AQ)}$ or $CO_{2(AQ)}$ carbonic acid	dissociates poorly	few H^+	
$HC_2H_3O_{2(AQ)}$ or $CH_3COOH_{(AQ)}$ acetic acid or ethanoic acid	dissociates quite poorly	very few H^+	

The concentration of hydrogen ions can be written this way: $[H^+]$
The higher the concentration, the more acidic the solution is.

The "opposite" of acids are called bases. A base solution (according to Arrhenius Theory) is that any substance that produces excess hydroxide ions in solution is called a base.

The list of bases in table L in your reference charts is expanded just below.

NaOH sodium hydroxide	dissociates very well	LOTS of OH^{-1}	The stronger the base, the more OH^{-1} ions (and more other ions as well). The more ions, the better it conducts electricity. Strong bases are good electrolytes. Weak bases are poor electrolytes. Ammonia is a weak base & a common household chemical. The BrØnsted-Lowry Theory can explain it.
KOH potassium hydroxide	dissociates very well	LOTS of OH^{-1}	
$Ca(OH)_2$ calcium hydroxide	dissociates somewhat	Less OH^{-1}	
NH_3 ammonia	an exception to the Arrhenius Theory, read more below.	There are no apparent OH^{-1}	

Arrhenius Theory also explains how and why when acids and bases combine that they can "neutralize" each other.

Acids and bases combine to ALWAYS form water + a salt.
Salts are chemical substances that are IONIC COMPOUNDS:

IONIC COMPOUNDS are: metal cation + nonmetal anion.

The neutralization reaction is summarized as:

acid + base yields water + a salt

This is a type of double replacement reaction, as you will see below, but because it also neutralizes the acid and base, that is the appropriate name.

example reaction:

$$HCl_{(AQ)} + NaOH_{(AQ)} \text{ ---> } HOH_{(L)} + NaCl_{(AQ)}$$

The hydrogen ions from the acid combine to the hydroxide ions in the base, and they make water. The other ions, the cation from the base (here that is the Na) plus the anion from the acid (here that is the Cl), combine to form a salt (an ionic compound, here that is table salt).

Another example could be:

$$H_2SO_{4(AQ)} + 2KOH_{(AQ)} \text{ ---> } 2HOH_{(L)} + K_2SO_{4(AQ)}$$

Acids have excess [H^{+1}] while bases have excess [OH^{-1}]

When a solution has equal numbers of hydrogen and hydroxide ions, they combine PERFECTLY into water, and the combined solution is no longer acid or base, it is said to be **NEUTRAL**. The acid is neutralized by the base. The base is neutralized by the acid. There are still many ions in solution (the anions of the acid and the cations of the base), so the resulting neutral solution is still a good electrolyte.

The Exceptional Situation of the base Ammonia: no apparent OH^{-1} ions.

In chemistry there are many exceptions to many "rules", including the Arrhenius Theory of Acids and Bases. Ammonia is a common substance used for cleaning houses (bathrooms mostly) and also used in fertilizer production. We use the reversible reaction of ammonia synthesis when examining LeChatelier's principle, and the economics of ammonia cannot be understated. It's important stuff. It is also an exception to the Arrhenius Theory of bases.

There are NO hydroxide ions in ammonia. In fact, there are NO IONS at all. How then can it act like a base? An alternate theory, called the BrØnsted-Lowry Theory describes acids and bases a totally different way than Arrhenius did. This theory is actually much more comprehensive, but it is more complicated.

We only have to know this: an alternative theory explains how ammonia is a base.
The BrØnsted-Lowry (B/L) theory is this:
B/L bases will be hydrogen ion "acceptors"
B/L acids will be hydrogen ion "donors"

Examine this equation, and see if you can spot how ammonia "accepts" a H^+ ion and how water "donates" it to the ammonia.

$$NH_{3(AQ)} + H_2O_{(L)} \text{ ---> } NH_4{}^{+1}{}_{(AQ)} + OH^{-1}{}_{(AQ)}$$

Imagine this way:
NH_3 plus H^{+1} equals $NH_4{}^{+1}$
H_2O minus H^{+1} equals OH^{-1}

NH_3 becomes $NH_4{}^{+1}$ in solution (RED),
and <u>that</u> hydrogen ion that it gains comes from the water.

The end result is the key, there are hydroxides in solution to provide the base properties. Technically, by definition, ammonia is the base and the hydroxide it the conjugate. This is all outside of our class chemistry, but the actual theory.
There are more theories about this as well, some concerning pairs of electrons, but they are all for NEXT YEAR, or college chemistry. Arrhenius theory explains all of our acids & bases except ammonia.

Measuring concentration of acid and base with the pH scale.

How concentrated are those hydrogen ions (or hydroxide ions)? The measure of the concentration of hydrogen ions is called the pH. The math for this is outside the scope of our class, but,

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

which means that the pH is equal to the negative logarithm of the concentration of hydrogen ions in solution.

This scale runs from zero (very high H^+ concentration, very strong acid) up to 14 (very low H^+ concentration, very strong base).

At exactly pH 7.0 the concentration of hydrogen ions equals the concentration of hydroxide ions, and they all turn to water.

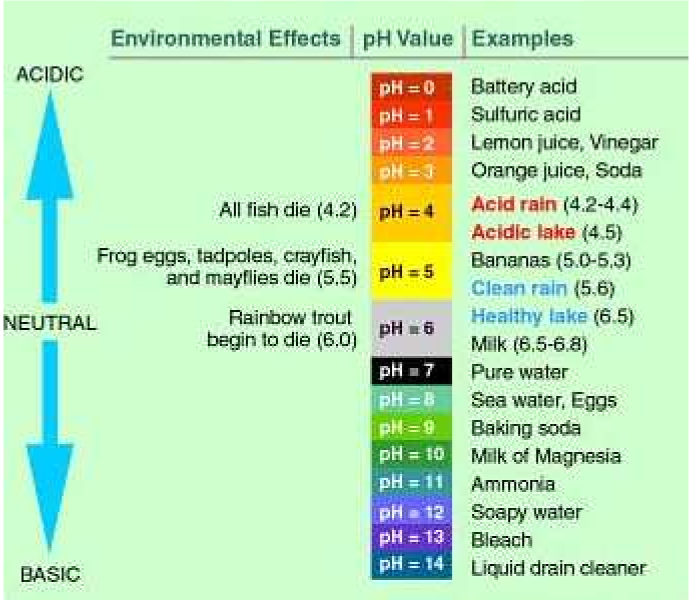
$$[H^+] = [OH^-] \text{ at NEUTRAL, neither acid or base}$$

At a pH of 5.0, there are more H^+ ions than OH^- ions in solution.

At a pH of 12.8, there are more OH^- ions than H^+ in solution.

Many solutions have equal numbers of both ions, water for example is one. Any aqueous solution with molecules (sugars or alcohols) are also neutral.

A substance such as pure alcohol, without water, acts unusual. What pH would it have? No pH? or a neutral pH (but there are not equal numbers of hydrogen ions and hydroxide ions, there are NO ions. I hope to find out the answer to this conundrum soon!

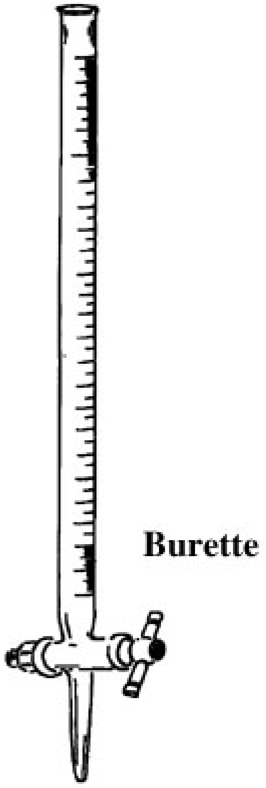


Environmental Effects	pH Value	Examples
	pH = 0	Battery acid
	pH = 1	Sulfuric acid
	pH = 2	Lemon juice, Vinegar
	pH = 3	Orange juice, Soda
All fish die (4.2)	pH = 4	Acid rain (4.2-4.4) Acidic lake (4.5)
Frog eggs, tadpoles, crayfish, and mayflies die (5.5)	pH = 5	Bananas (5.0-5.3) Clean rain (5.6)
Rainbow trout begin to die (6.0)	pH = 6	Healthy lake (6.5) Milk (6.5-6.8)
	pH = 7	Pure water
	pH = 8	Sea water, Eggs
	pH = 9	Baking soda
	pH = 10	Milk of Magnesia
	pH = 11	Ammonia
	pH = 12	Soapy water
	pH = 13	Bleach
	pH = 14	Liquid drain cleaner

So,

Acids have $[H^+] > [OH^-]$ bases have $[H^+] < [OH^-]$

Acid Base Neutralization in Lab



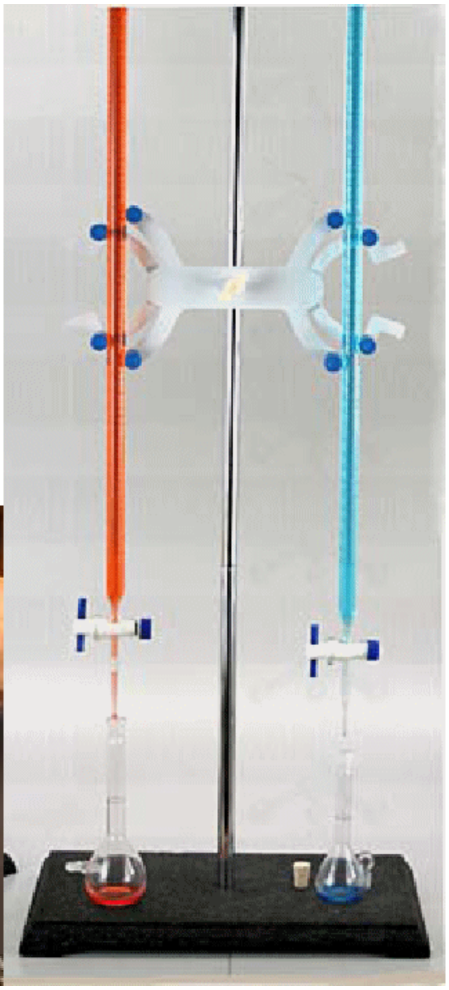
To neutralize an acid with a base, or do the reverse, to neutralize a base with an acid, you have to EXACTLY match up the hydrogen ions with the hydroxide ions to balance them and end up with no excess of either. In lab we will use a buret, which is a long glass tube with MANY lines. Burets are expensive and very accurate, so be careful handling them and reading them.

We use a buret of acid to neutralize a beaker of base.

We use a buret of base to neutralize a beaker of acid.

In Vestal, the Acid is ALWAYS on the left, for safety. It does not matter at all which side of the set up the acid is, but if we all do the same, less errors will happen.

In that photo below, the acid is RED and the base is BLUE



We read a buret by the accurately drawn lines on it. When it is "FULL", it reads 0.0 mL (none missing). When you open the valve at the bottom, solution flows out.

At the end you can read the final reading in the buret, which is always higher than when you started.

The difference between initial and final readings is how many mL of solution you ran out of the buret.

At some point you will have put in enough acid to the base, or base to the acid to perfectly neutralize the solutions. How will you know when they are at neutral?

If you add an acid base indicator into the beaker there will be a color change at neutral to show you when the hydrogen ions equal the hydroxide ions (or be very close to equal).

Acid Base Titration Math

Titration is the process to measure how acidic or basic a solution of unknown concentration is. We use a known molar strength acid and drip in the unknown base until we reach the neutral point. This point is a color change due to the presence of an acid base indicator (we'll use phenolphthalein in lab).

The formula, which is on the back of your reference table is:

$$(M_A)(V_A) = (M_B)(V_B)$$

And stands for the molarity of the acid multiplied by the volume of the acid is equal to the molarity of the base multiplied by the volume of the base. Since we'll always know both volumes and the molarity of either the acid or base, we can solve for the missing molarity.

At neutral (or as close as we can get in lab) when the acid and base are neutralized, we can use our chemistry math to calculate unknown molarity of the base.
Volume units can be any units, as long as they are the same on both sides of the equal sign.

An example problem:
If it takes 35.6 mL of acid of 3.40 Molarity to completely neutralize 88.5 mL of KOH, what is the molarity of the base?

$$(M_A)(V_A) = (M_B)(V_B)$$

$$(3.40)(35.6) = (M_B)(88.5)$$

Solve for molarity of base

$$13.676... = (M_B)$$

$$13.7\text{ M} = (M_B)$$ with 3 significant figures

Phenolphthalein as you can look up on table M, changes color at pH between 8.2 and 10.0 from colorless to pink. Since we add base to acid, drop by drop, making that acid less acidic and more neutral, when we get to neutral we won't notice since the solution will be colorless before we get to pH 7 and stay colorless until pH 8.2 or so. This is clearly worth explaining since it makes NO SENSE at first glance.

Because we're working with relatively small volumes, one drop should be enough base to move from colorless to clear, or in reverse, one drop of acid should make the pink basic solution return to colorless. We can get to within a drop of neutral in our lab, but not perfectly to a pH of 7.0 with this indicator.

Enjoy high school, it's not perfect but it's way better than anything else. Electronic pH meters exist, but they're difficult to maintain, expensive, and we'd only need them for a couple of days a year, which means phenolphthalein is our indicator of choice.

Unusual Acid Base Neutralization Situations

When we neutralize HCl and NaOH using the "normal" titration math formula of $M_AV_A=M_BV_B$, often we overlook the exact formula meaning. M_A means the molarity of the acid, or more importantly, the molarity of the H^+ ions.

1.0 moles of HCl have 1.0 moles of H^+, and 0.50 moles of HCl have 0.50 moles of H^+.

But what if it's a "double acid" or "triple acid"? H_2SO_4 has 2 H^+ ions per molecule, or two moles of H^+ ions per mole of molecules. Phosphoric acid, H_3PO_4 has 3 H^+ ions per molecule, or three moles of H^+ per mole of molecules.

The M_A in these problems needs to be looked at closely, because with H_2SO_4 and NaOH, there is NOT a 1:1 ratio of H^+ hydrogen ions to OH^{-1} hydroxide ions. This changes the math somewhat, and you need to watch for this.

One mole of hydrochloric acid and one mole of sodium hydroxide neutralize to one mole of sodium chloride and one mole of water

$$HCl_{(AQ)} + NaOH_{(AQ)} \longrightarrow NaCl_{(AQ)} + HOH_{(L)}$$

Here it's clear it takes 2 moles of sodium hydroxide to neutralize one mole of the sulfuric acid, because the acid is a "double H^+" acid rather than a single like HCl above.

$$H_2SO_{4(AQ)} + 2NaOH_{(AQ)} \longrightarrow Na_2SO_{4(AQ)} + 2HOH_{(L)}$$

Here, it takes three moles of sodium hydroxide to neutralize one mole of phosphoric acid, because the acid is a "triple H^+" acid and the base is a "single OH^{-1}" base.

$$H_3PO_{4(AQ)} + 3NaOH_{(AQ)} \longrightarrow Na_3PO_{4(AQ)} + 3HOH_{(L)}$$

The same works in “reverse” when the acid is single, but the base is doubled.

$$2HCl_{(AQ)} + Ca(OH)_{2(AQ)} \qquad CaCl_{2(AQ)} + 2HOH_{(L)}$$

It takes 2 moles of the “single H^+” acid to neutralize the “double OH^{-1}” base.

The M_A or the M_B stands for the molarity of the H^+ ions, or the molarity of the OH^{-1} ions.

They do not stand for the molarity of the molecules of acid, or formula units of base.

1 mole HCl = 1 mole H^+ ions (single acid)
1 mole H_2SO_4 = 2 moles H^+ ions (double acid)
1 mole H_3PO_4 = 3 moles H^+ ions (triple acid)

1 mole NaOH = 1 mole OH^{-1} ions (single base)
1 mole $Ca(OH)_2$ = 2 moles OH^{-1} ions (double base)

Make sure you don’t always “assume” that the $H^+:OH^{-1}$ ratio is 1:1, because it’s not and you need to adjust for that.

Use of an Indicator for Neutralization

In lab we will use an indicator in our beakers to show us with color when we have neutralized (or come very close to) our acid base combination.

There are several indicators on table M that we could use, litmus and bromthymol blue both have some color shift through the pH value of 7.0 which is neutral.

Both would give us a fairly qualitative measure and would not be 100% accurate.

If we use phenolphthalein instead, which we do use, there is the clear to pink color change which is much easier to see. Of course that change happens at 8.2 rather than 7.0 pH.

This "problem" is solved by this thinking: with the volumes of acid and base we are using, the difference between 7.0 and 8.2 in pH is quite literally ONE DROP. The mistake we are making in pH exactitude is nearly outside of our ability to measure accurately. We will take this inaccuracy on the chin (so to speak) and realize that using either of the other two indicators would introduce as much error, or more, since the change is between colors, our exactitude would be as inaccurate as accepting the pH values provided by the phenolphthalein.

We won't be "exactly" neutralizing in lab, but we'll be close enough for a high school chemistry lab. In order to be closer to true neutral we'd need to use an expensive pH meter, which we do not have (or need). Color indicators are pretty accurate, considering that they are qualitative and not quantitative.

One odd extra point: the cation called the **hydronium ion** from table E (polyatomic ions, page 1 of reference table) fits into acid and base. A different way to write and to "understand" hydrogen ions in water has these ions attached to the water.

Water + hydrogen ions = hydronium ions $H_2O + H^{+1} \longrightarrow \mathbf{H_3O^{+1}}$

This ion shows up on the Regents exam from time to time, but it really indicates acid ions in solution. Hydronium ions are also known as hydrogen ions in water.

Acid Base Indicators of Table M

on Table M there are six chemical indicators we can use to give us a handle on the pH values of a solution. They will show, by color, a range of pH values that could be for a solution. They never can give an EXACT pH reading, but they can show a qualitative measure of concentration easily.

Methyl Orange is the first indicator. The table shows us that for any solution with a pH value of 3.2 or less, the solution would be red with this indicator. If the pH of the solution changed to 4.4 or higher, then the color of the solution would be yellow. Only in that 3.2 to 4.4 range of pH would the colors be a blur between red an yellow.

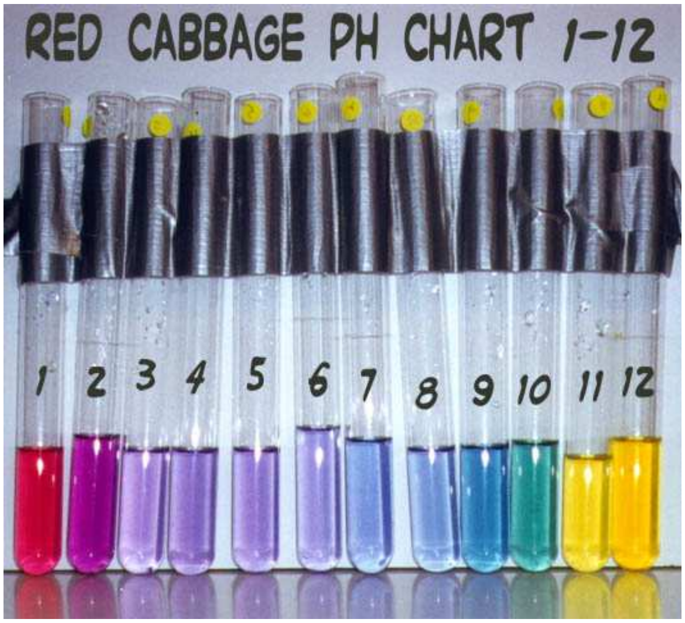
By using this indicator you could be able to accurately say that the solution has a pH of less than 3.2 (if it were red), or over 4.4 (if it were yellow), or between those values (if it were "orange-y" in color). This is not a quantitative measure of concentration, but it is quick and the colors are fun.

Thymol blue is the last indicator in table M. That shows as yellow if the pH of the solution is less than 8.0, or blue if it's over 9.6. Between those values the color is a blurry greenish color.

Each indicator has its own range and will be helpful at different times. There are many other acid base indicators (methyl red, indigo carmine, etc.) that we could use, but the reference table lists just six.

In lab we used cabbage juice, which is called a "natural" acid base indicator since it comes from a plant and not a bottle in the chemistry store room.

We showed that cabbage juice can change 9 different colors relatively easy, and it will change at least 14 different colors under mildly careful conditions (concentration of cabbage juice, proper measuring of acid & base, plus really good eyes to note these differences!).



Note how good this dog's eyes really are! Acid on his right, base on his left.

(that was a bad joke!)

How do acid base indicators actually work?

Most indicators are weak acids or bases. The example I will use is phenolphthalein indicator. It has a chemical formula of: $C_{20}H_{14}O_4$ but as you will learn, these types of formulas for ORGANIC MOLECULES like this can hide what the compound really is, which is a weak acid.

The phenolphthalein molecule is colorless. When it ionizes in water it will turn pink, which is what we see. When a base is added to the solution (pH rises above 8.2 on table M), LeChatelier's predicts correctly that there will be a shift forward as these hydroxide ions combine with the loose hydrogen ions, letting more and more phenolphthalein ions in solution.

In this "equation" phenolphthalein is written as H-Ph. The H stands for the hydrogen that gets dissociated from the rest of the "Ph" phenolphthalein molecule. This puts H^{+1} + Ph^{-1} ions into solution. Since it is a weak acid, not many molecules dissociate (like vinegar or acetic acid). This molecule is colorless and there are too few "Ph^{-1}" anions to pink up the solution.

$$\text{H-Ph}_{(S)} + H_2O_{(L)} \;\; <===> \;\; H^{+1}{}_{(AQ)} + Ph^{-1}{}_{(AQ)}$$

"phenolphthalein" plus water in dynamic equilibrium with hydrogen ions plus "phenolphthalein ions"

add OH^{-1} shifts equilibrium forward as the hydroxides + hydrogen ions = WATER, putting more and more of the PINK "phenolphthalein ions" into the solution.

Adding more acid, or H+ ions shifts the equilibrium to the reverse, which lowers the amount of pink ions on the loose, making the solution colorless at LOWER pH.

As acids are added, the excess H^{+1} ions actually force the Ph^{-1} ions into reverse. Less Ph^{-1} ions means even less pink (too few to see becomes less).

When a base is added to the solution, these OH^{-1} ions immediately attach to the loose hydrogen ions, which forces the reversible reaction to shift it's equilibrium forward, putting more and more of these pink Ph^{-1} ions into solution. That gives the solution a pink color we see (at pH above 8.2, which has more and more OH^{-1} ions!).

Table M practice

Using table M from the reference tables, determine what color each solution should be with the specific indicator in it. Answers on the NEXT PAGE (don't peek)

pH	indicator	Color???
1.5	litmus	
2.3	thymol blue	
3.1	methyl orange	
4.7	bromthymol blue	
5.9	bromcresol green	
6.6	methyl orange	
7.0	thymol blue	
7.2	bromcresol green	
8.4	phenolphthalein	
9.6	litmus	
10.2	bromthymol blue	
11.8	phenolphthalein	
12.0	thymol blue	
14.0	phenolphthalein	

Table M practice **ANSWERS**

Using table M from the reference tables, determine what color each solution should be with the specific indicator in it. Answers on the NEXT PAGE (don't peek)

pH	indicator	Color???
1.5	litmus	red
2.3	thymol blue	yellow
3.1	methyl orange	red
4.7	bromthymol blue	yellow
5.9	bromcresol green	blue
6.6	methyl orange	yellow
7.0	thymol blue	yellow
7.2	bromcresol green	blue
8.4	phenolphthalein	hint of pink
9.6	litmus	blue
10.2	bromthymol blue	blue
11.8	phenolphthalein	pink
12.0	thymol blue	blue
14.0	phenolphthalein	pink