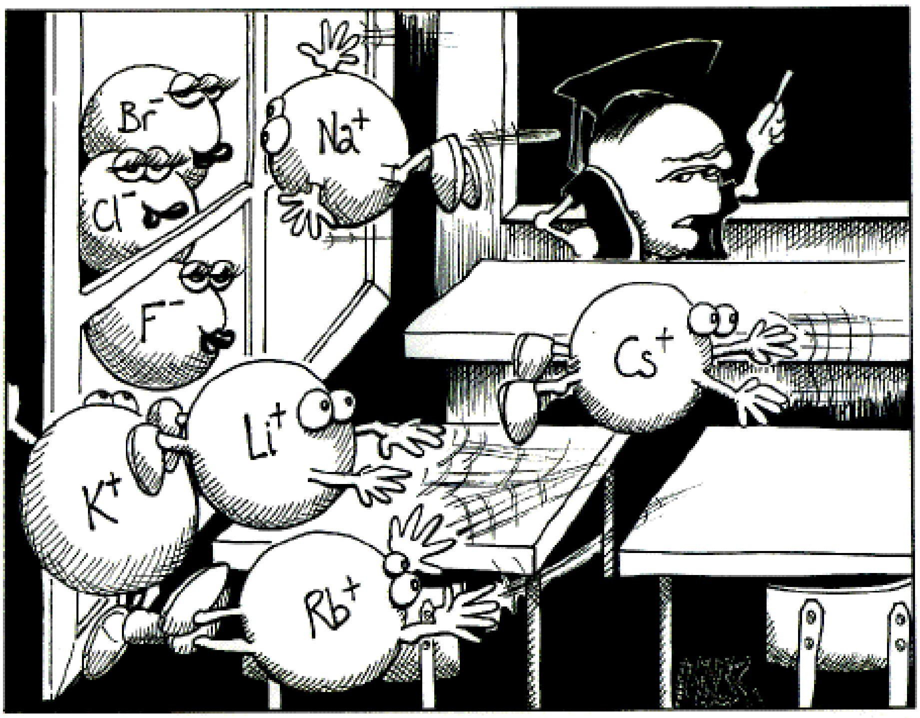
Acids & Bases

Acids and bases are something you've certainly heard of before. Most of you will already know a thing or two about acids.

You may know something about the pH scale used to measure how strong an acid it. You've likely all heard of acid rain. Many of you might associate acids as being a little dangerous and know that you wouldn't want to spill any on yourself. Indeed some acids, and bases as well, are definitely very dangerous. However many of you will likely have already eaten or drank some acid today and splashed bases over your face and hands. And you all have a stomach full of acid right now.



"Perhaps one of you gentlemen would mind telling me just what it is outside the window that you find so attractive...?"

1.1 What are Acids & Bases?

Here's a partial list of some common acids and bases, along with some chemical formulas:

Some Common Acids	Some Common Bases
hydrochloric acid, HCl (stomach acid is HCl)	sodium hydroxide, NaOH (lye or caustic soda)
sulfuric acid, H_2SO_4	potassium hydroxide, KOH (lye or caustic potash)
nitric acid, HNO_3	magnesium hydroxide, $Mg(OH)_2$ (milk of magnesia)
acetic acid, $HC_2H_3O_2$ (vinegar)	calcium hydroxide, $Ca(OH)_2$ (slaked lime)
carbonic acid, H_2CO_3	ammonia, NH_3
formic acid, HCOOH	
citric acid, $C_6H_8O_7$	
acetylsalicylic acid, commonly known as aspirin $C_6H_4(OCOCH_3)CO_2H$	

Acids and bases have characteristic properties. The table below highlights some of these key characteristics. You'll want to learn this list.

Key Characteristics of Acids	Key Characteristics of Bases
▪ sour taste (e.g. lemons; grapefruit; vinegar; sour milk)	▪ bitter taste
▪ react with active metals such as zinc and magnesium to produce hydrogen gas	▪ generally no noticeable reaction with active metals
▪ form electrolytic solutions (conduct electricity) because they produce ions	▪ form electrolytic solutions (conduct electricity) because they produce ions
▪ cause certain dyes to change color; litmus paper turns red, for example	▪ cause certain dyes to change color; litmus paper turns blue for example
	▪ slippery feel (e.g. soapy feel)
▪ neutralized by bases (neutralized means that the substance no longer has acidic, or basic, properties)	▪ neutralized by acids

By now you may already made a good guess as to the key "components" of acids and bases. If you look carefully at the formulas listed on the last page (Acids and Bases Section 1.1) you'll see that the acids all have one thing in common, and most of the bases another.

Although the properties of acids and bases had been recognized for a long time, it was Svante Arrhenius in the 1880's who determined that:

- the properties of acids were due to the presence of hydrogen ions, H^+, and
- the properties of bases were due to the presence of hydroxide ions, OH^-.

ACIDS

hydrochloric acid, HCl

sulfuric acid, H_2SO_4

nitric acid, HNO_3

acetic acid, $HC_2H_3O_2$ (vinegar)

carbonic acid, H_2CO_3

formic acid, HCOOH

acetylsalicylic acid, $C_6H_4(OCOCH_3)CO_2H$

BASES

sodium hydroxide, NaOH

potassium hydroxide, KOH

magnesium hydroxide, $Mg(OH)_2$

calcium hydroxide, $Ca(OH)_2$

ammonia, NH_3 - oops! Where's the OH^-?

This became known as the **Arrhenius Theory of Acids and Bases**

1.3 Ionization & Dissociation

You should recall from our unit on solutions that electrolytic solutions are those that conduct electricity because the substance dissolves in water to produce ions.

- Solutions that conduct electricity well are **strong electrolytes** - they are good conductors because they break down well and produce many ions in solution.
- **Weak electrolytes** do not conduct electricity as well because fewer ions are produced in solution.

Dissociation

When ionic compounds dissolve to produce ions the process is typically called **dissociation**.

Dissociation of ionic compounds occurs when water molecules "pull apart" the ionic crystal. This occurs due to strong attractions between the polar ends of the water molecule and the positive and negative ions within the crystal. Water molecules then surround the positive cations and negative anions; this is called hydration.

Examples of dissociation equations:

$NaCl_{(s)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$

$Na_2SO_{4\,(s)} \rightarrow 2\ Na^{+}(aq) + SO_4^{2-}{}_{(aq)}$

$(NH_4)_3PO_{4\,(s)} \rightarrow 3\ NH_4^{+}(aq) + PO_4^{3-}{}_{(aq)}$

$NaOH_{(s)} \rightarrow$

$KOH_{(s)} \rightarrow$

$Mg(OH)_{2\,(s)} \rightarrow$

Pull

What do you notice about the last three reactions shown above? All three of them produce hydroxide ions, OH^- . These three compounds - NaOH, KOH, and $Mg(OH)_2$ are all Arrhenius bases since they produce hydroxide ions in solution.

Ionization

Ionization is a process very similar to dissociation, however it involves acids as a rule. Here some ionization reactions:

$HCl_{(g)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$

$H_2SO_{4\,(g)} \rightarrow 2\,H^{+}_{(aq)} + SO_4^{2-}{}_{(aq)}$

$HC_2H_3O_{2\,(l)} \rightleftarrows$

(The reason a double arrow is used for acetic acid will be discussed later)

HNO_3 --->

The above examples are really the simplified versions of what really happens. Acids a re only acids when in solution with water. Check this out!

$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$

$H_2SO_{4\,(g)} + H_2O_{(l)} \rightarrow 2\,H_3O^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

What's happening? Can you see it? It appears some donating and accepting of H^+ is going on!

$HC_2H_3O_{2\,(l)} + H_2O_{(l)} \rightleftarrows$

$HNO_3 + H_2O$ -->

You should notice a H_3O^+, it's called **hydronium**

$$H_3O^+_{(aq)} = H^+_{(aq)}$$

1.4 Brønsted-Lowry Theory of Acids & Bases

That lack of an OH^- group in ammonia is a bit of a problem at first glance, but let's take a closer look at what happens when NH_3 is added to water:

ammonia, NH_3 - oops! Where's the OH^-?

$$NH_{3\,(g)} + H_2O_{(l)} \rightleftarrows NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

Voila! There's the hydroxide group we were looking for!

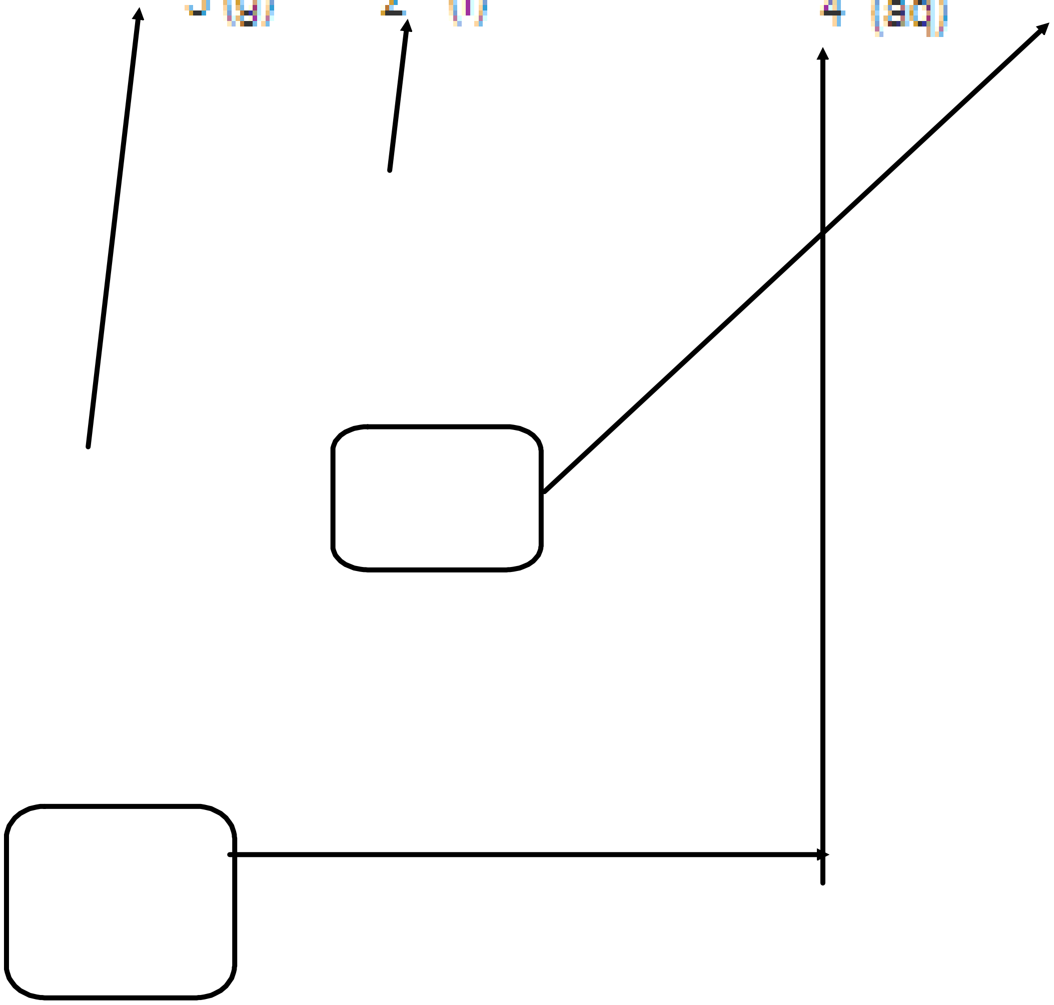
One of the current models of acids and bases was developed by Johannes Brønsted, a Danish chemist, and English chemist Thomas Lowry, each working independently. What is now known as the **Brønsted-Lowry theory of acids and bases** can be stated as follows:

Acids are substances that produce a hydrogen ion (or a proton donor)

Bases are substances that can accept a hydrogen ion (or a proton acceptor)

H^+ = proton

$$NH_{3\,(g)} + H_2O_{(l)} \rightleftarrows NH_4{}^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$$



Bronsted-Lowry Theory states:

-----> Acids are substances that produce a hydrogen ion, H^+ (proton) and are willing to donate this proton.

-----> Bases are substances that are willing to accept a hydrogen ion, H^+ (proton)

----> Acids are proton donors

----> Bases are proton acceptors

$$NH_{3\,(g)} + H_2O_{(l)} \rightleftarrows NH_4^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$$

Acceptor

Base

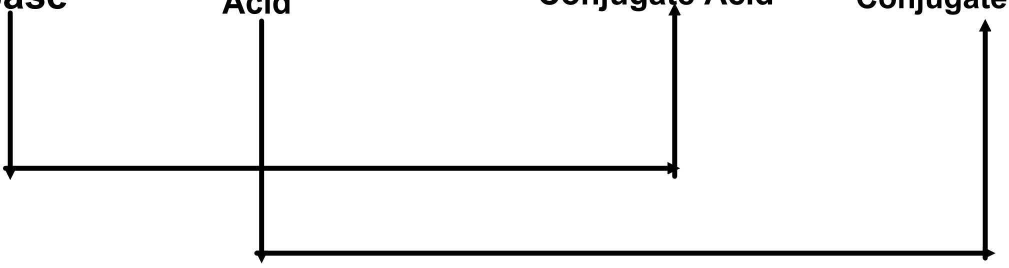
Donor

Acid

1.5 Conjugate Acid-Base Pairs

$$NH_{3\,(g)} + H_2O_{(l)} \rightleftarrows NH_4^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$$

Base **Acid** **Conjugate Acid** **Conjugate Base**



when H_2O donated its proton (H^+), it became OH^-

OH^- is now called a conjugate base

when NH_3 accepted the the proton (H^+), it became NH_4^+

NH_4^+ is now called a conjugate acid

****every acid has a conjugate base and the C.B. is what the acid became after proton donation.

****every base has a conjugate acid and the C.A. is what the base became after proton acceptance.

The conjugates will always be listed on the product side of the reaction.

Conjugate acid-base pairs differ from each other by the presence or absence of a single hydrogen ion (proton). Every acid has a conjugate base, and every base has a conjugate acid.

1.6 Amphoteric Substances

Were you surprised in the last section to see water being described as an acid? In the ammonia reaction, water acted as an acid because it donated a proton (hydrogen ion) to ammonia:

$$NH_{3\,(g)} + H_2O_{(l)} \leftrightarrow NH_4{}^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$$

Compare this to another reaction we looked at earlier when we saw how hydrochloric acid acted as an acid by donating a proton to water:

$$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^{+}{}_{(aq)} + Cl^{-}{}_{(aq)}$$

In this reaction, water is acting as a base because it accepts a proton from HCl.

Substances that can act as an acid is one reaction and as a base in another are called **amphoteric substances**

Just to confuse you, the term **amphiprotic** means the same thing

1. $HSO_4{}^{-} + H_3O^{+} \leftrightarrow H_2SO_4 + H_2O$ — $HSO_4{}^{-}$ accepts a proton from H_3O^{+}
2. $HSO_4{}^{-} + OH^{-} \leftrightarrow H_2O + SO_4{}^{2-}$ — $HSO_4{}^{-}$ gives (donates) a proton to OH^{-}

practice answers

Polyprotic Acids

Acids that can donate more than one hydrogen ion are called **polyprotic**

We have seen examples of acids that contain more than one hydrogen ion that can be lost. Sulfuric acid, H_2SO_4, for example, has two hydrogen ions that it can give up. The first hydrogen ion is released as:

$$H_2SO_{4\,(aq)} \rightarrow H^+_{(aq)} + HSO_4^-{}_{(aq)}$$

$$\downarrow$$

$$HSO_4^-{}_{(aq)} \rightarrow H^+_{(aq)} + SO_4^{2-}{}_{(aq)}$$

How many H^+ can citric acid, $H_3C_6H_5O_7$, release?

Citric acid can release three H^+. The reactions would be:

1. $H_3C_6H_5O_{7\,(aq)} \rightarrow H^+_{(aq)} + H_2C_6H_5O_7^-{}_{(aq)}$
2. $H_2C_6H_5O_7^-{}_{(aq)} \rightarrow H^+_{(aq)} + HC_6H_5O_7^{2-}{}_{(aq)}$
3. $HC_6H_5O_7^{2-}{}_{(aq)} \rightarrow H^+_{(aq)} + C_6H_5O_7^{3-}{}_{(aq)}$

2.1 Strong & Weak Acids & Bases

You have undoubtedly heard of the pH scale before and know that it has something to do with indicating how strong or weak an acid is. In this part of the unit we will work towards defining acid and base strength in terms of pH, but there are several important steps along the way. It will be important that you understand each step.

pH means "potential of Hydrogen

- Strong electrolytes conduct electricity well because the compound produces many ions in solution
- Weak electrolytes conduct electricity poorly because they produce few ions in solution

Strong acids produce many H^+ ions
(or H_3O^+ ions)
weak acids produce few H^+ ions

The stronger the acid,
the more H^+ ions are produced

AND

Strong bases produce many OH^- ions
weak bases produce few OH^- ions

The stronger the base,
the more OH^- ions are produced.

Strong acids and bases are essentially one-way reactions - the acid or base breaks down completely to produce ions. At equilibrium there are very few reactants left (very low concentration); only products - the ions.

Strong Acids	Strong Bases
$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$	$NaOH_{(aq)} \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$
$H_2SO_{4\,(aq)} \rightarrow H^+_{(aq)} + HSO_4{}^{1-}{}_{(aq)}$	$Mg(OH)_{2\,(aq)} \rightarrow Mg^{2+}{}_{(aq)} + 2\ OH^-_{(aq)}$

Weak acids and bases, however, do not ionize completely. For weak electrolytes, equilibrium lies to the left side of the equation (the reactant side) and there will be few ions present. The double arrow is commonly used to indicate the partial ionization of the solution. Some examples:

Weak Acids	Weak Bases
$HC_2H_3O_{2\,(aq)} \leftrightarrow H^+_{(aq)} + C_2H_3O_2{}^-{}_{(aq)}$	$NH_{3\,(aq)} + H_2O_{(l)} \leftrightarrow NH_4{}^+{}_{(aq)} + OH^-_{(aq)}$
$HCHO_{2\,(aq)} \leftrightarrow H^+_{(aq)} + CHO_2{}^-{}_{(aq)}$	$NH_2CH_{3(aq)} + H_2O_{(l)} \leftrightarrow NH_3CH_3{}^+{}_{(aq)} + OH^-_{(aq)}$

2.2 Ionization Constants: K_a, K_b, and K_w

Because acid/base solutions are systems at equilibrium, we can write equilibrium constant expressions for these systems. We'll return to the equilibrium constant, a concept we first addressed in the Equilibrium unit.

K_a and K_b

Here are several examples, including some that show water as a reactant. Pay attention to the physical states and whether or not a particular substance is included in the equilibrium expression. Also notice that we now identify the equilibrium constant as K_a for acids and K_b for bases.

Table of Acid and Base Strengths gives K_a and K_b values

$$HCl_{(aq)} \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

$$K_a = \frac{[H^+]\ [Cl^-]}{[HCl]} = 1.3\times10^{6}$$

$$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

$$K_a = \frac{[H_3O^+]\ [Cl^-]}{[HCl]} = 1.3\times10^{6}$$

$$HCHO_{2\,(aq)} \leftrightarrow H^+_{(aq)} + CHO_2^-{}_{(aq)}$$

$$K_a = \frac{[H^+]\,[CHO_2^-]}{[HCHO_2]} = 1.8 \times 10^{-4}$$

$$Mg(OH)_{2\,(aq)} \rightarrow Mg^{2+}_{(aq)} + 2\,OH^-_{(aq)}$$

$$K_b = \frac{[Mg^{2+}]\,[OH^-]^2}{[Mg(OH)_2]}$$

Why a K_b expression now?

$$NH_{3\,(aq)} + H_2O_{(l)} \leftrightarrow NH_4{}^{+}{}_{(aq)} + OH^{-}{}_{(aq)}$$

Let's try this one ourselves

So we now have a numerical way to indicate acid and base strength - K_a and K_b

A large value of K_a
means there are many
H^+ ions in solution -
in other words, **a strong acid**

A large K_b indicates
many OH^- ions -
a strong base

K_w

We usually do not think of water as producing ions, but that isn't the case. Water does ionize, although not very well.

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

$$[H^+] = [OH^-] = 1.0 \times 10^{-7}$$

2.3 Calculating $[H^+]$ and $[OH^-]$

This is one of the most important sections of this unit mathematically, so be sure you understand every step.

How you calculate ion concentrations depends on whether you have a strong acid (or base) or a weak acid (or base).

Calculating Ion Concentrations for Strong Acids & Bases

For strong acids and bases, the concentration of the ions can be readily calculated from the balanced equation. Consider these examples carefully.

1. Calculate the hydrogen ion concentration in a 0.050 M solution of hydrochloric acid.

2. Calculate the hydroxide ion concentration in a 0.010 M solution of barium hydroxide, $Ba(OH)_2$. Barium hydroxide is a strong base.

Calculating Ion Concentrations for Weak Acids & Bases

Weak acids and bases require a much different approach to finding ion concentrations. Once you know you have a weak acid or base, follow these steps in finding ion concentrations:

1. Write a **balanced** equation for the reaction
2. You will need to know the value of K_a or K_b - if it is not given in the question, look it up in a Table of Acid and Base Strengths.
3. Set up the equilibrium constant expression. You will know the value of K_a (or K_b) and the concentration of the acid; you will be solving the equation for the concentration of the ions.

Follow along with these examples very carefully!

1. Calculate the hydrogen ion concentration in a 0.10 M acetic acid solution, $HC_2H_3O_2$. K_a for acetic acid, a weak acid, is 1.8×10^{-5}.

Begin by writing the **balanced** reaction:

Pull

The question gives us the concentration of the acid, $HC_2H_3O_2$ (0.10 M).

We need to find the concentration of H^+, which will also equal the concentration of $C_2H_3O_2^-$ (why?)

Because ionization is NOT complete because this is a **weak acid**, **[H^+] will NOT equal [$HC_2H_3O_2$]**. Instead we must calculate it using the equilibrium constant expression.

2. Calculate the hydroxide ion concentration, $[OH^-]$, in a 0.025 M solution of analine, $C_6H_5NH_2$, a weak base with $K_b = 4.3\times10^{-10}$

Pull

Finding $[OH^-]$ in Acids and $[H^+]$ in Bases

Remember K_w from the previous section? Now we learn why it is important.

sample calculation:

If

$[OH^-]$ = 0.001 3001

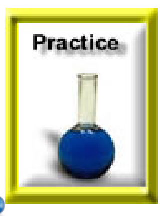
Use water's equilibrium constant to determine $[H^+]$:

K_w = $[H^+][OH^-]$

For any acid or base
you can calculate
both $[H^+]$ and $[OH^-]$

Acids
First determine $[H^+]$
then use K_w to calculate $[OH^-]$

Bases
First determine $[OH^-]$
then use K_w to calculate $[H^+]$



answers

2.4 The pH Scale

It is finally time to turn our attention to pH. pH is just another way to express $[H^+]$, the hydrogen ion concentration of an acidic or basic solution.

Recall:

pH means "potential of Hydrogen

Hydrogen acid concentrations are often small numbers, such as 1.3×10^{-3}. pH is a method of transforming this number into something that is a little easier to work with.

In math class you may have learned about logarithms - log for short. We'll leave the definitions of logs to math and just work with how to find them here.

Get your calculators out. Different calculators work in slightly different ways, and it will be **VERY IMPORTANT** for you to know how to use yours when working with logs.

Example: Find the log of 1.0×10^5

Pull

Number	Log
1×10^{-3}	
2.5×10^{12}	
3.5×10^{-9}	

What about pH?

pH is defined as the negative log of hydrogen ion concentration.

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

	$[H^+]$	pH
1.	1×10^{-3}	
2.	2.5×10^{-11}	
3.	4.7×10^{-9}	
4.	5.8×10^{-4}	
5.	1.0×10^{-7}	

Notice the last example.

1.0×10^{-7} is the $[H^+]$ in pure water.

Pure water therefore has a pH of 7.

Acids	Bases	Neutral solutions
pH < 7	pH > 7	pH = 7
The lower the pH, the stronger the acid	The higher the pH, the stronger the base	

Examples of Calculating pH

1. Calculate the pH of a 0.01M HNO_3 solution?

Begin finding pH by first finding $[H^+]$.

2. Find the pH of a 0.01 M solution of ammonia.

Pull

pOH

pOH means potential OH^- (Hydroxide)

$$pOH = -\log[OH^-]$$

$[OH^-] = 4.2 \times 10^{-4}$

pOH =

Pull

$pH + pOH = 14$

Does the number 14 ring a bell?

If pOH = 3.4, what's pH?

Finding $[H^+]$ when you know pH

To convert pH into $[H^+]$ involves taking the **antilog** of the negative value of pH .

$$[H^+] = \textbf{antilog } (-pH)$$

antilog is the 10^x button on your calculator

Example. We have a solution with a pH = 8.3. What is $[H^+]$?

You should get the answer **5.0×10^{-9}**

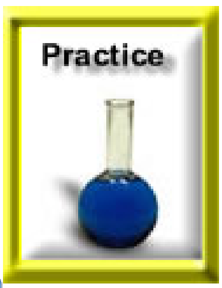
Here are two final examples. Think ***carefully*** and determine what, exactly, you are asked to find and the steps needed to get there.

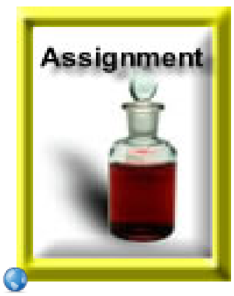
1. Find the hydronium ion concentration in a solution with a pH of 12.6. Is this solution an acid or a base? How do you know?

Pull

2. A 0.24M solution of the weak acid, H_2CO_3, has a pH of 3.49. Determine K_a for H_2CO_3 (carbonic acid).

Pull





2.5 Indicators

Indicators are dyes that change colour under varying conditions of acidity. Although not as accurate as instruments such as pH meters in determining acidity, indicators can be used to give less precise measure of acidity. You should recall in our introduction to acids and bases that we mentioned that litmus is red in acids and blue in bases. Litmus is an indictor that changes colour from red to blue in the pH range of 5.5 to 8.0. Other indicators and their colours are listed in the table of Acid - Base Indicators which you should print out and keep available.

Here are some questions to try. A short acid-base indicator table is given here:

Indicator	pH range	Colour change
methyl orange	3.2 - 4.4	red to yellow
litmus	5.8 - 8.0	red to blue
phenolphthalein	8.2 - 10.0	colourless to pink

1. A given solution turns methyl orange yellow, litmus blue, and phenolphthalein red. What is the approximate pH of the solution?

Pull

2. What color would methyl orange, litmus, and phenolphthalein turn when testing:

   a. vinegar (pH = 3)
   b. sea water (pH = 8)

	methyl orange	litmus	phenolphthalein
vinegar			
sea water			

Pull

3.1 Neutralization Reactions

What happens when an acid such as HCl is mixed with a base such as NaOH:

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

(this is a DDR)

When an acid and a base are combined, water and a salt are the products.

Double displacement reactions of this type are called **neutralization reactions.**

We can write an expanded version of this equation, with aqueous substances written in their longer form:

$$H^{+}_{(aq)} + Cl^{-}_{(aq)} + Na^{+}_{(aq)} + OH^{-}_{(aq)} \rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)} + H_2O_{(l)}$$

Removing the spectator ions

$$H^{+}_{(aq)} + OH^{-}_{(aq)} \rightarrow H_2O_{(l)}$$

3.2 Acid-Base Titrations

Acid-base titrations are lab procedures used to determine the concentration of a solution. We will examine it's use in determining the concentration of acid and base solutions. Titrations are important analytical tools in chemistry.

During an acid-base titration, an acid with a known concentration (a **standard solution**) is slowly added to a base with an unknown concentration (or vice versa). A few drops of indicator solution are added to the base.

The indicator will signal, by colour change, when the base has been neutralized (when $[H^+] = [OH^-]$).

At that point - called the **equivalence point** or **end point** - the titration is stopped. By knowing the volumes of acid and base used, and the concentration of the standard solution, calculations allow us to determine the concentration of the other solution.

Titration Calculations

To use titration information to calculate the concentration of the unknown solution, you must know the following information. Note the abbreviations that will be used in our calculations.

- The concentration of one of the solutions, the acid for example (M_A)
- The volume of acid used for the titration (V_A)
- The volume of base used for the titration (V_B)

What you will calculate:

- The concentration of the other solution, the base for example (M_B)

$$M_A V_A = M_B V_B$$

Example 1

During a titration 75.8 mL of a 0.100 standard solution of HCl is titrated to end point with 100.0 mL of a NaOH solution with an unknown concentration. What is the concentration of the NaOH solution?

Begin with a balanced equation for the reaction:

Pull

M_A = 0.100 M M_B = M_B

V_A = 75.8 mL V_B = 100.0 mL

$$M_A V_A = M_B V_B$$

Example 2

A 20.0 mL solution of strontium hydroxide, $Sr(OH)_2$, is placed in a flask and a drop of indicator is added. The solution turns colour after 25.0 mL of a standard 0.0500M HCl solution is added. What was the original concentration of the $Sr(OH)_2$ solution?

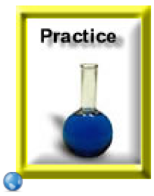
Write a **balanced equation** for the neutralization reaction:

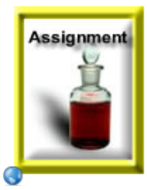
Our second example will involve a reaction that does **not** involve a 1:1 ratio between the acid and the base.

M_A	=	0.050 M	M_B	=	$\mathbf{M_B}$
V_A	=	25.0 mL	V_B	=	20.0 mL

Pull

$$M_A V_A = 2\, M_B V_B$$





answers