

Chemistry 12

Tutorial 1

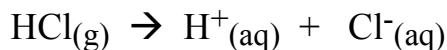
Bronsted Acids and Bases

Tutorial 1 will introduce the following:

1. **Hydronium ions** and how they are formed.
2. **Bronsted-Lowry definitions** of acids and bases.
3. **Equilibria** Involving Acids and Bases.
4. **Conjugate acid-base pairs**.
5. **Polyprotic acids** and **Amphiprotic Anions**.

Hydronium Ions and how they are formed

Acids dissociate in water to form hydrogen ions (H^+) and others. For example, when hydrogen chloride gas dissolves in water, you get....



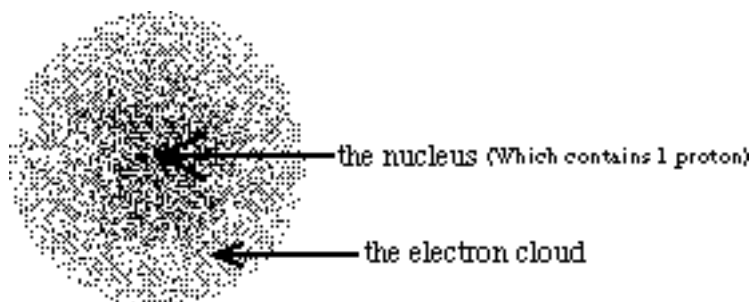
So you might picture some H^+ ions and Cl^- ions floating around between water molecules in the solution.

That's OK in Chemistry 11, but it's a bit oversimplified!

It IS true that H^+ ions are released from the HCl. But they don't just float around by themselves.

H^+ ions are hydrogen ions.

Let's talk about hydrogen atoms. Almost all hydrogen atoms consist of **one proton**, **no neutrons** and **1 electron**. The proton is deep in the center in the **nucleus** and the electron "buzzes" around the atom in what we call an electron "cloud".



When a hydrogen atom (H) forms a hydrogen ion (H^+), remember, it **loses** an electron.

When it does this, it also loses its electron cloud! So what's left? Just a very **very** tiny nucleus which contains **1 proton!**

For this reason, the H^+ ion is often called a **proton**. Because, that's exactly what it is! Reactions in which H^+ ions are transferred from one thing to another are called **proton transfers**.

The +1 charge on the H^+ ion or proton, is concentrated in a very small volume, much smaller than in any other ion. (All other ions have at least 1 electron, so they have an electron cloud, which makes them thousands of times bigger than H^+ , which has no electron cloud.)

Because this charge is **concentrated** in a very small volume, it acts like it is quite powerful and it is attracted strongly to anything even remotely negative!

Remember that in an acid solution, H^+ ions (or protons as we also call them) are surrounded by **water** molecules. Let's take a closer look at a water molecule.

Recall in Chemistry 11, you were introduced to "electron-dot" or "Lewis" diagrams of atoms. (These, again are an oversimplification but that's another story!)

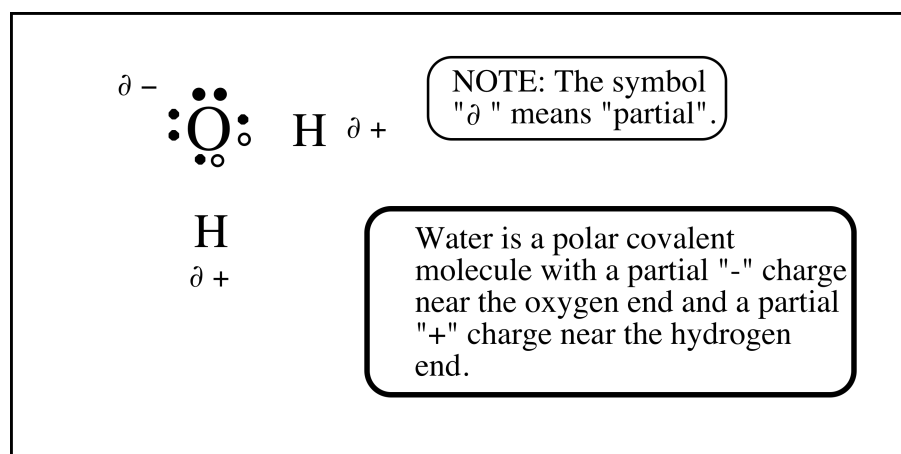
You might recall that an oxygen atom has 6 valence electrons (6 electrons in the outer energy level):



Hydrogen has one valence electron:

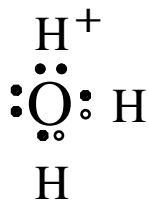


When hydrogen and oxygen combine to form water, they share their valence electrons. But, you may also recall that oxygen, being a non-metal has a **stronger pull** on electrons (electronegativity) than hydrogen, so the shared electrons are **closer to the oxygen atom**. This makes water a **polar covalent** molecule. Since there are more electrons close to the "oxygen end" of the water molecule, that end has a partial **negative charge**. The "hydrogen end" has less electrons around it, hence has a partial **positive charge**:



Now that partial negative charge on the oxygen end looks very attractive to our old friend the H^+ ion! And that's exactly where it goes. It "sits" on one of the "electron pairs" of the oxygen atom. Remember the H^+ has no electrons itself, so it doesn't bring any more electrons into the picture.

It has one proton, though. What that does is bring another + charge into the picture:



This thing (made up of a proton (H^+) added to a water molecule) is an **ion** because it has a charge.

Its formula is **H_3O^+** and it's called the **hydronium ion**.

The hydronium ion always forms when an acid dissolves in water. The H^+ from the acid always goes to the nearest water molecule and forms H_3O^+ .

Another way to look at the hydronium ion is to take the point of view of the proton (H^+).

Adding **water** to something is called **hydration**. (*Just like taking water away is called dehydration.*)

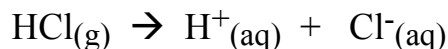
So if you were a proton, you would have a water molecule "added to you".

For this reason, a hydronium ion could be considered a **hydrated proton**.

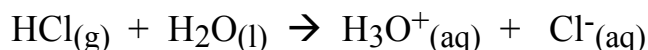
Whichever way you look at it, just remember that instead of thinking of an acid solution containing H^+ ions (as you did in Chem. 11), we now think of acid solutions containing H_3O^+ (hydronium) ions.

All acid solutions contain hydronium (H_3O^+) ions. It is the hydronium ion which gives all acids their properties (like sour taste, indicator colours, reactivity with metals etc.)

Now, recall that in Chemistry 11, when HCl gas dissolves in water, we wrote:

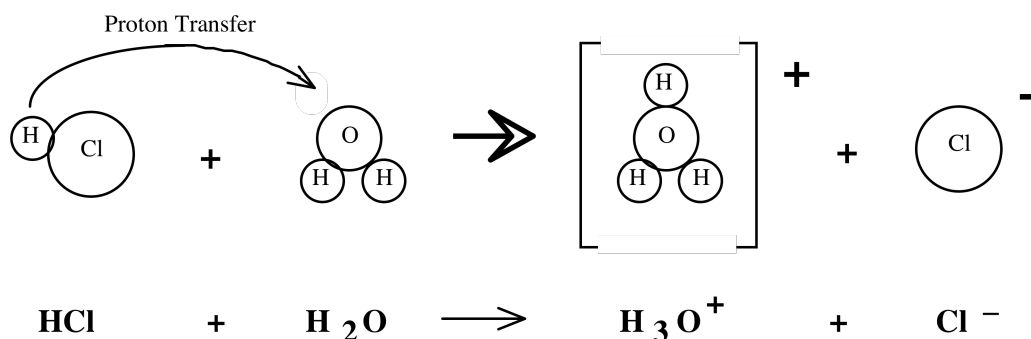


Now, in **Chemistry 12**, we write the following:



The proton (H^+) has been **transferred** from the HCl molecule to a water molecule, to form a hydronium (H_3O^+) ion and a Cl^- ion.

This type of reaction is called **ionization** (because ions are being formed) We can look at this using some models:



Make sure you study the diagram so you can visualize in your mind, what's going on when you see equations like this. In this diagram, you must realize that it is NOT an H atom that is moving. The H atom leaves its electron behind with the Cl, so it is **H⁺ (a proton)** that moves to the water molecule. The Cl⁻, now having the electron that H left behind, gains a negative charge.

All acids behave similarly in water; they donate (or give) a proton (H⁺) to the water, forming hydronium ion (H₃O⁺) and the negative ion of the acid. Another example might be the ionization of nitric acid (HNO₃):



There's a few of these for you to try on the next page....

Bronsted-Lowry Definition of Acids and Bases

You might recall that the definition of an "acid" according to Arrhenius was a substance that released H⁺ ions (protons) in water.

A couple of fellows called Bronsted and Lowry came up with a theory which is more useful when dealing with equilibrium and covers a wider range of substances.

Our apologies to Mr. Lowry, but from now on we will just refer to "Bronsted", when we actually mean both of them. It's just bad luck that his name came later in the alphabet!

According to **Bronsted** (and "What's his name?"):

An **acid** is any substance which **donates** (gives) a **proton** (H⁺) to another substance.

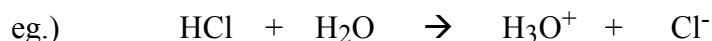
A **base** is any substance which **accepts** (takes) a **proton** from another substance.

Or we can also say:

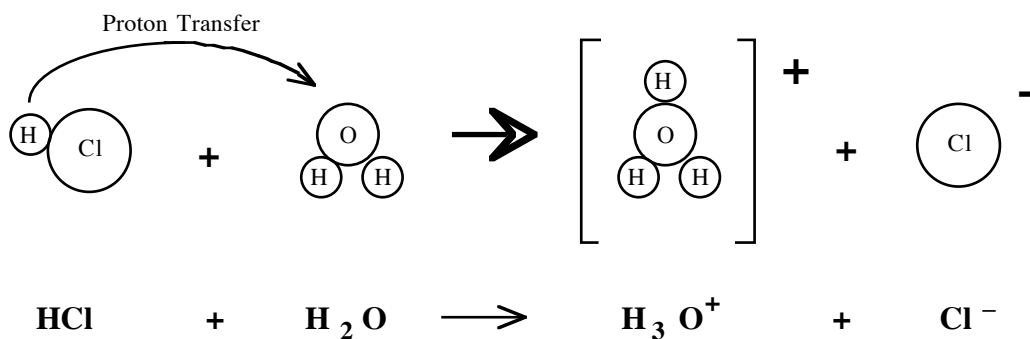
A **Bronsted Acid** is a **proton donor**

A **Bronsted Base** is a **proton acceptor**

Let's look at a couple of equations and see how we can identify the acids and the bases. (We will omit the subscripts (aq) etc. just for simplification here.)

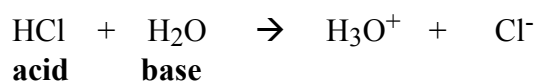


Looking at this diagram again:

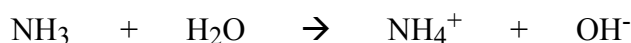


We see that the HCl is **donating** the proton and the water is **accepting** the proton.

Therefore **HCl is the Bronsted acid** and **H₂O is the Bronsted base**.



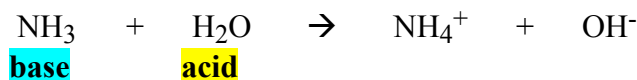
Let's look at another example:



Now, the NH₃ on the left has changed into NH₄⁺ on the right, that means it must have accepted (taken) a proton. (It has one more H and one more (+) charge.) Since it has accepted a proton it's called a **base**.

The H₂O, this time has donated (lost) a proton as it changed into OH⁻. (It has one less H and one less (+) charge --- one "less (+) charge" than "0" is (-1) or (-).) Since it has donated a proton it's called an **acid**.

So now we can label these:



Now, you may be a little confused! First we tell you that H₂O is a base (see the reaction near the top of this page), and then we go and tell you H₂O is an acid. What's going on, Bronsted?

Well, both of these statements are correct. Sometimes water acts like a base (takes a proton) and sometimes it acts like an acid (donates a proton). This is just like you. If you buy something (donate money) you are a buyer. If you sell something (accept money), you are a seller. I'm sure you have been both at various times. Animals that can live either in the water or on land are called *amphibians*.

For things that can be “either / or”, we can use the prefix “amphi”

A substance that can act as **either an acid or a base** is called **amphiprotic**.

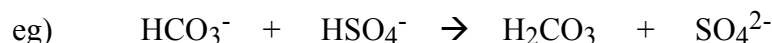
Water (H₂O) is an example of an **amphiprotic** substance. When it was with HCl, it acted like a **base**, but when it was with NH₃, it acted like an **acid**.

Not only molecules can lose or gain protons. Ions can too.

When something **loses a proton** (acts as an **acid**), it turns into something with one less H and one less (+) charge (which means the same as one more (-) charge.)

When something **accepts a proton** (acts as a **base**), it turns into something with one more H and one more (+) charge (which means the same as one less (-) charge.)

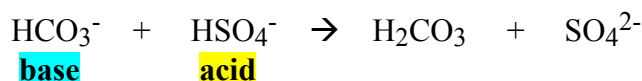
So what you have to do is look at the right side of the equation, as see whether the substance gained or lost a proton.



HCO₃⁻ must have accepted a proton (*1 H and 1 (+) charge*) to form H₂CO₃, so it must be the **base**.

HSO₄⁻ must have donated a proton (*1 H and 1 (+) charge*) to form SO₄²⁻, so it must be the **acid**.

so the answer is:



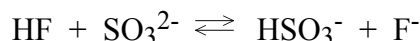
Read this example again, looking carefully at the charges and # of H atoms, and how they change from each reactant to it's product. By the way, there is no rule for which one comes first in the equation. Basically, each one has a 50/50 chance of coming first. You have to work it out by counting H's and charges.

Equilibria Involving Acid and Bases

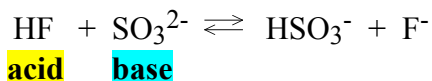
So far, we've been considering reactions which only go one way. In reality, most acid-base reactions go forward and in reverse. (They are at equilibrium)

If a proton is transferred during the **forward** reaction, we can also assume there will be a proton transfer in the **reverse** reaction.

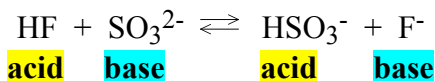
Here's an example:



If we consider the reaction going to the right, HF is donating a proton, and is therefore defined as the **acid**, while SO₃²⁻ is accepting a proton, and therefore acting as a **base**.

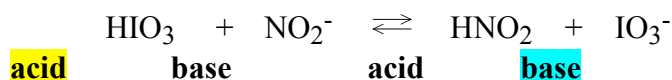


Now, when we look at the *reverse* reaction, in which HSO_3^- reacts with F^- to form HF and SO_3^{2-} , we see that HSO_3^- **donates** a proton and F^- **accepts** a proton. Thus, HSO_3^- acts as an **acid**, while F^- acts as a **base**. So in any acid, base reaction, we start out with an acid and a base on the left and we end up with another acid and base on the right.



Conjugate Acid-Base Pairs

Looking at this reaction:



Notice the HIO_3 on the left. We know that it must lose one proton (H^+) to become IO_3^- on the right. Also notice that HIO_3 is acting as an **acid** while IO_3^- is acting as a **base**.

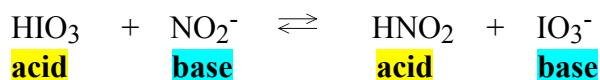
HIO_3 and IO_3^- form what is called a *conjugate acid-base pair*.

The only difference between these two is the IO_3^- has one less "H" and one more (-) charge than the HIO_3 . All conjugate acid-base pairs are like this.

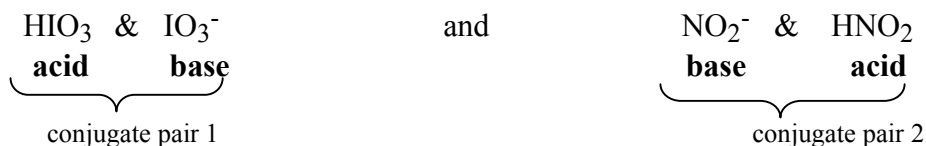
The form with one more H (eg. HIO_3) is called the **conjugate acid**.

The form with one less H (eg. IO_3^-) is called the **conjugate base**.

Out of every acid-base reaction, you always get 2 conjugate pairs. For example, in this reaction:



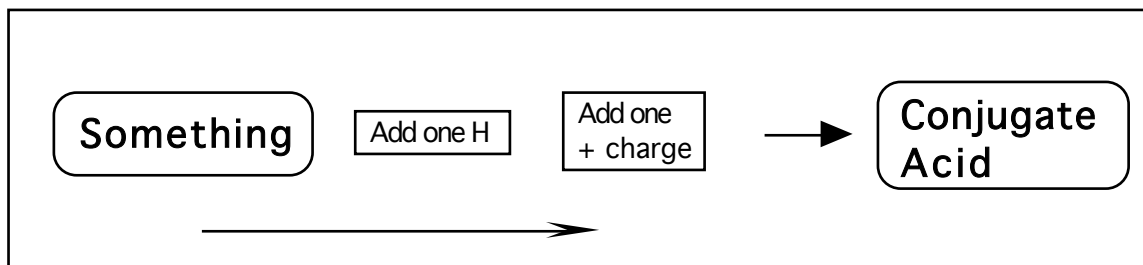
The two conjugate pairs are:



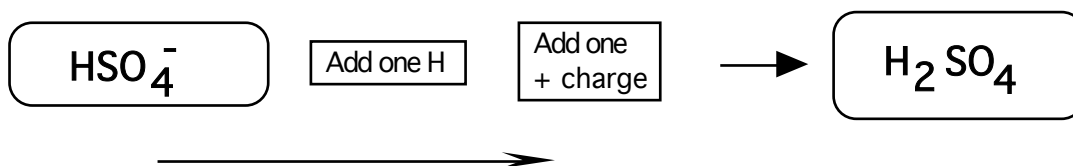
(NOTE: The "1" and the "2" in "conjugate pair 1" etc. has no special meaning. Pair 1 was just the one we happened to pick first. The NO_2^- & HNO_2 could just as well have been called "conjugate pair 1")

One of the things you'll be required to do is, **given an ion or molecule**, write the formula for the **conjugate acid** of it. Also, **given an ion or molecule**, write the formula for the **conjugate base** of it.

To find the conjugate acid of something:

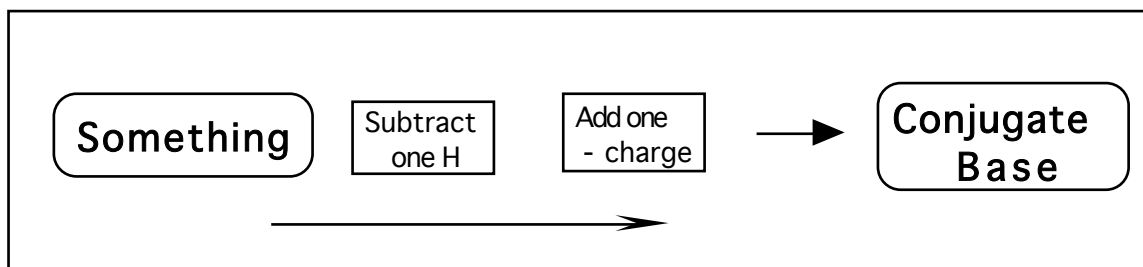


For example, let's say we want to find the conjugate acid of HSO_4^-

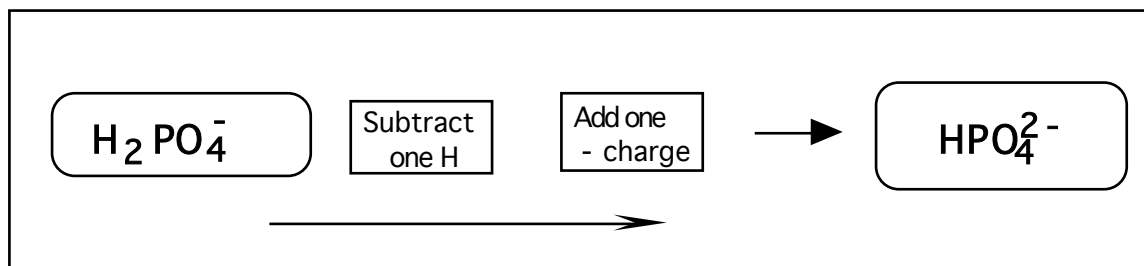


Remember, adding one (+) charge to something that has a (-) charge, brings the charge to "0".

Now it's time to find the **conjugate bases** of things. If you added an "H" and a (+) charge to get a conjugate acid, I think you can probably guess that to get a conjugate base, you subtract an "H" and add one (-) charge. (or subtract one (+) charge, which means the same thing!)



Let's say we want to find the conjugate base of the ion H_2PO_4^- . If we use this procedure:



Polyprotic Acids and Amphiprotic Anions

After looking at this rather fearsome title, read the following. It's not that bad!

So far, we've been looking at acids that only have one proton (H^+) to release. These are acids with one hydrogen in their formulas (eg. HCl , HNO_3 , $HClO$ etc.)

Acids that release only one proton are called monoprotic acids.

Believe it or not, acetic acid (CH_3COOH) is monoprotic. This is because only the "H" on the end of this acid (The "H" on the "COOH") comes off in solution. The other three "H"s are bonded directly and strongly to the Carbon atom in the "CH₃" and are not released.

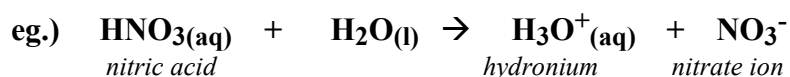
"H"s bonded directly to Carbon atoms (like in "CH₃", "CH₃CH₂" etc.) are NOT released in solution and are not considered as "acidic protons".

In another example, the acid $HCOOH$ and the acid $CH_3CH_2CH_2COOH$ are both monoprotic. (Only the "H" on the end of the "COOH" comes off.)

You will also notice that for these organic acids (or more precisely "carboxylic acids", the "acidic proton" is the "H" on the right end of the formula.

For inorganic acids (like HCl , HNO_3 , $HClO_3$ etc.), the acidic proton is always the "H" on the left side of the formula.

The ionization of a monoprotic acid is quite simple, as we've seen before:



Notice, that if the name of the acid ends in "ic", it's conjugate base is an ion that ends in "ate".

I would bet that you would be able to correctly guess the name that we give to acids that

release 2 protons (H^+ 's): "_____protic"

Acids that release two protons are called diprotic acids.

Some examples of diprotic acids are: H_2SO_4 , H_2CO_3 , H_2SO_3 etc.

Some acids go as far as releasing 3 protons:

Acids that release three protons are called triprotic acids.

Some examples of triprotic acids are: H_3PO_4 , H_3AsO_4 , H_3BO_3 etc.

To a chemists, anything more than one can be called “poly”

So here’s another term:

Acids that release more than one proton are called polyprotic acids.

Note here that “polyprotic acids” would *include* “diprotic” and “triprotic” acids.

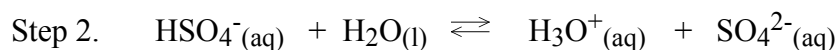
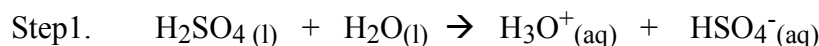
Personally, I don’t know of many acids that give off more than 3 protons. (One is “EDTA” or “ethylenediamine tetraacetic acid”)

Stepwise ionization of polyprotic acids

It is important to know that polyprotic acids, when mixed with water, do not release all their protons in one step.

They release protons one by one. Each step is a separate equilibrium.

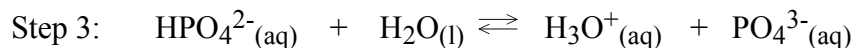
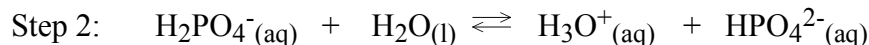
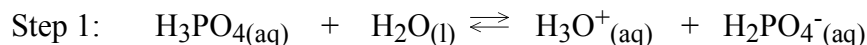
Let’s look at an example. Here is the 2 steps in the ionization of sulphuric acid (H₂SO₄)



You’ll probably wonder why there’s a single arrow in step 1 and a double arrow in step 2. Because H₂SO₄ is a **strong** acid, it releases it’s first proton 100%. That is the first step goes to completion. Every single H₂SO₄ molecule breaks up into H₃O⁺ and HSO₄⁻ ions.

But that doesn’t mean the **second** proton come off that easily! The HSO₄⁻ ion is **NOT** a strong acid, so it does **not** release all of it’s protons in water. That’s why there is a double arrow in the Step 2 equation.

Here are the equations for the step-wise ionization of a **triprotic acid** (phosphoric acid)



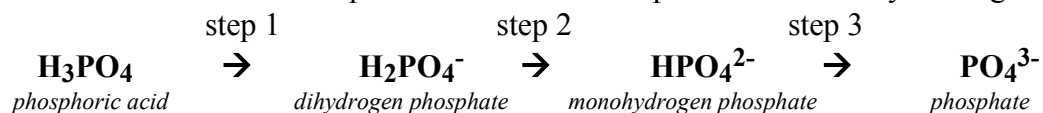
Each one of these is a separate equilibrium. Notice that one “H” and one (+) charge comes off of the starting acid each time.

Now take out your acid chart. (Yes, that means you! If you don’t, you probably won’t understand the next couple of statements and you will get frustrated etc.)

Notice the relative strengths of the acids H₃PO₄, H₂PO₄⁻ and HPO₄²⁻. You can find them all on the left side of the table.

We can interpret their relative positions by saying that each time a proton is removed from a polyprotic acid, it gets harder to remove the next one.

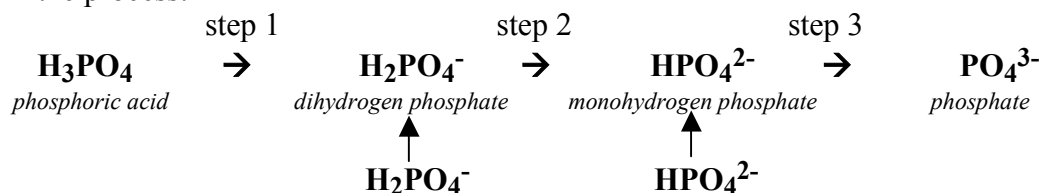
We can summarize the species formed in the stepwise ionization by leaving out the H_3O^+ and the H_2O :



Notice naming system for the ions as they form.

Amphiprotic Anions

Let's go back to our example of phosphoric acid and look at the two ions formed in the middle steps of the process:



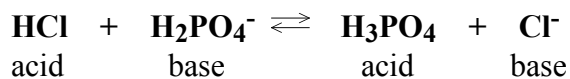
Notice that both of these ions have at least one "H" on the left and at least one (-) charge.

Anything with an "H" on the left can act as an **acid**, and many ions with a (-) charge can act as **bases**.

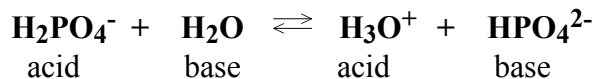
Because these ions H_2PO_4^- and HPO_4^{2-} can act as either acids or bases, they are called **amphiprotic**, and since they have at least one (-) charge, they are called **anions**.

Putting this all together, we call these **amphiprotic anions**.

The dihydrogen phosphate ion (H_2PO_4^-), when mixed with a strong acid (like HCl), will play the role of a base. eg.)



If the dihydrogen phosphate ion is mixed with a base, or even with water, it will play the role of an acid: eg.)



We will be looking more at amphiprotic anions later on in this unit, but for now, it's useful to remember that:

The ion(s) formed all but the last step of the ionization of a polyprotic acid are amphiprotic anions.