ISUs FOR DUMMIES

Volume 2 :

Acids And Bases

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Hi. Since I was COMPLETELY BORED over the holidays I decided to make another ISU FOR DUMMIES BOOK.

For those of you who didn't read the other one (it was 19 pages long), the series of ISUS FOR DUMMIES basically takes away any of the problems u've been having with an ISU.

(I explain it like I'm explaining it to a dummy.)

Acids and Bases are A BIT harder than Gases, so YEAH.

Without further ado...

ACIDS AND BASES FOR DUMMIES!!!!

In the beginning we are introduced to two types of substances.

ACIDS. AND BASES.

ACIDS

Now those of you who actually read the other ISU'S FOR DUMMIES book (which must of been alot as the average was 89) You'll know my obsession with great flaming titles.

Continuing on...

An Acid is a substance that

- *Reacts with Hydrogen Metals such as Zinc or Magnesium to Liberate Hydrogen Gas*
- *Has a SOUR taste*
- *Conducts electricity*
- *Turns Litmus Paper Red*
- *Reacts with Metal Carbonates to Release Carbon Dioxide Gas*
- *Acids are DANGEROUS CHEMICALS*

BASES

A Base is a substance that

- *Has a Bitter Taste*
- *Feels slippery*
- *Conducts electricity*
- *Can cause severe burns and inflammation to the skin*
- *Turns Litmus Paper Blue*

- ***ACIDS react with acidic solutions to destroy or neutralize the acidic properties.***

(This statement will become really important when we start talking about Arrhenius)

Before we introduce Arrhenius, I want to get the term electrolyte across.

An Electrolyte is a substance which forms ions in an aqueous (water) solution.

Arrhenius's Theory

(a man with solution!!!)
In 1887 a guy named Svante Arrhenius presents the idea of "electrolytic dissociation"

Basically he says that

- *An acid is a substance that produces HYDROGEN IONS when dissolved in water*
- *A Base is a substance that produces Hydroxide Ions (OH-) when dissolved in water.*

Well for a time this definition sounded pretty cool, because it explained why Acids and Bases neutralize each other.
How?

Well think of it this way.

You have the reaction

HCl + NaOH --> NaCl + H2O

Or in other words, an acid plus a base --> a salt plus a water.

If you dissociate this equation you'll get:

$H^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl^- + H2O(l)$

Okay now let's get the net ionic equation...

$H^+ + OH^- \rightarrow H_2O$

Well there you go. A solution with H+ ions --> an acid.
A solution with OH- ions --> a base.
And they neutralize to form water.

So most ppl were pretty happy with that theory until they found the following **Problems**

- *THERE ARE SOME BASES WHICH DON'T HAVE HYDROXIDE IN THEIR FORMULAS.*
- *MOST HYDROGEN CONTAINING SUBSTANCES ARE NOT ACIDS.*
- *ARRHENIUS'S DEFINITION IS LIMITED TO REACTIONS IN WATER --> IT DOESN'T ALLOW FOR NON-AQUEOUS SOLVENTS.*

So obviously, people start thinking that Arrhenius might be wrong.

And in 1923, two guys, Dane Johannes Bronsted and Thomas Lowry made a new theory:

THE BRONSTED LOWRY DEFINITION OF ACIDS AND BASES.

- ***An Acid is Defined as a proton donor.***

- ***A Base is defined as a proton acceptor***

Now we can basically say that H+ is a proton. Why? Well a regular H atom has one electron and one proton. If you take away the electron (hence H+) then you get one proton. Not very hard.

So a Bronsted-Lowry Acid is a hydrogen-ion DONOR
And a Bronsted-Lowry Base is a hydrogen ion ACCEPTOR.

Okay, now we are going to think of two substances that are not defined as an Arrhenius Bases– BUT THEY ARE defined as a BRONSTED-LOWRY Bases

These are:

Cl- and NH3.

Neither dissociate to form OH-.
But they DO ACCEPT H+ (protons).

So, we can sum up this theory by saying the following statement

Acid $\rightleftarrows$ *Base +* H^+

So an Acid dissociates to form a Base and H^+*.*
And a Base + H^+ *makes an acid*

CONJUGATE BASES AND ACIDS:

We're going to take a different approach then Mrs. Pall did, and work backwards.

WATER can act as an Acid or a Base.

This is explained when you look at the following reactions.

$HCl + H_2O \rightleftarrows H_3O^+ + Cl^-$

Here, HCl *is your ACID,* H_2O *is your Base,*
AND

H_3O^+ *is what you call a CONJUGATE ACID*
Cl^- *is what you call a CONJUGATE BASE*

A conjugate acid is formed when a base receives a proton in a reaction.

A conjugate base is formed when an acid loses a proton in a reaction.

Hence, the conjugate acids and bases are ALWAYS PRODUCTS.

Arite, lets go on.

On page 4 of the review, she says that:

THE STRENGTH (degree of ionization-dissociation) of an ACID depends on the acid's CONJUAGATE BASE (ANION)

Underline this statement. Then highlight it. And then circle it.

It's a very important statement – you have to understand it

The Explanation:

Lets first define **Strong** *and* **Weak**

Take an acid that's going to lose its H (dissociate). It's important to realize that ALL of the acid DOES NOT dissociate all the time! Sometimes most of it does. Sometimes some of it does. And in some cases, all of it does.

So the more acid that's dissociated – the more STRONGER the acid.

And the smaller amount of dissociation – the WEAKER the acid.

Get it?
HCl is a VERY STRONG acid, and CH3COOH is a VERY WEAK ACID.
You can't tell if its a strong acid by simply looking at the equation.

Before looking at Lewis, I just want to clear something up:

Water can act as an acid or a base.

H_2O as a BASE:

$$HCl + H_2O \rightleftarrows H_3O^+ + Cl^-$$

<u>H_2O as an ACID:</u>

$$NH_3 + H_2O(l) \rightleftharpoons NH_4^+ + OH^-$$

Therefore Water is called <u>AMPHIPROTIC</u>.

Other terms include

<u>Monoprotic</u> (Donate ONE proton/mol)

<u>Polyprotic:</u> DONATES more than one proton per mole.

Lewis Acids and Bases

To be honest, I really don't like Lewis's Acids and Bases.

You can always tell if something's an acid or a base using Bronsted Lowry. Just see if it has an H in front of it – then its an acid.

But, you see Bronsted-Lowry did not consider substances like Boron Trifluoride (BF3) or Aluminum Chloride (AlCl3) Acids.

(because they didn't have an H)

Lewis Did.

He defined his theory as:

- ***An Acid is a Molecule or Ion that can accept an electron pair.***

- ***A Base is a Molecule or Ion that can donate an Electron Pair.***

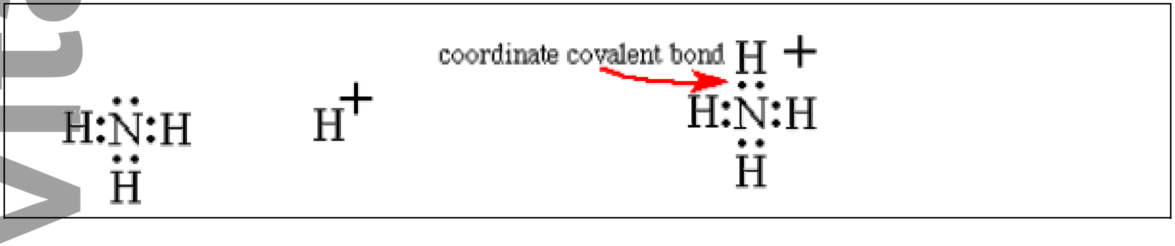
Bronsted Lowry said that ACIDS DONATED PROTONS. Lewis says that ACIDS ACCEPT An ELECTRON Pair. Note that they're both saying the same thing, except with different terms – protons and electrons.

Okay, this is where most people get confused. Electron Pair? Why an electron pair? Why not an electron? How do the examples work?

And this is the explanation.

First, we'll go over a definition. You know what an ionic bond is. You know what a covalent bond is. But what about a co-ordinate covalent bond?

A co-ordinate covalent bond is a bond in which both electrons are provided by ONE atom. For example:



So the NH3 basically provides the electrons for the H+

- *So the H+ accepts an electron pair, and the NH3 donates an electron pair.*

So lets look at one of HER examples.

$BF_3 + :NH_3 \longrightarrow BF_3NH_3$

So in this case, BF3 needs another electron pair, so it gets it from NH3. So BF3 is the acid, NH3 is the base.

Strength of Acids

We've already gone over strength. But now we're going to bring a few constants in.

As a review:

We have to always remember that when an acid dissociates to form a base and H+ ions, **ALL OF IT DOES NOT DISSCIATE.**

For example if we had the acid: CH3COOH, we dissociate it to form H3O+ and CH3COO-.

But not all of it dissociates. When we look at a solution with this acetic acid, we can tell that there are **only a few H3O+ ions and CH3COO- ions in the solution. The solution mostly contains most of the original molecules of CH3COOH.**

Therefore the reaction is said to be on the "LEFT" side – and said to be WEAK.

$$CH3COOH \; + \; H2O(l) \; <==> \; H3O^{+} \; + \; CH3COO^{-}$$

On the other hand, if we look at a solution with HCl,

We can tell that almost all of it dissociates into H+ and Cl- ions. The acid is therefore STRONG. (and is said to be on the RIGHT side)

Now, we're going to introduce a new constant. This is called:

$$\mathbf{K_a}$$

This is called the acid dissociation constant. The STRENGTH of the acid is determined by the following equation

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

(What Mrs. Pall wrote)

If you look at a SIMPLIFIED version of this equation, you can look at:

$$K_a = \frac{[A^-][H^+]}{[HA]}$$

This one is easier to understand.

Basically its saying that the Ka constant is the multiplication of the number of H+ ions and "base" ions (such as CH3COO-) – or in this case "A-". This is the divided by the number of original acid molecules.

SO, if you have a LOT of A- ions and a LOT of H+ ions, and you don't have much HA left, that means you have a very STRONG acid, because most of it DISSOCIATED, and there are very few ORIGINAL ACID MOLECULES LEFT.

Therefore if you have a HIGH Ka, your dissociation is totally to the RIGHT, and you have a STRONG ACID.

If you have a SMALL Ka, you didn't have a lot of dissociation, so its completely to the LEFT, and you have a WEAK ACID.

Get it?

Before we go on, I just want to clear up a couple of stuff

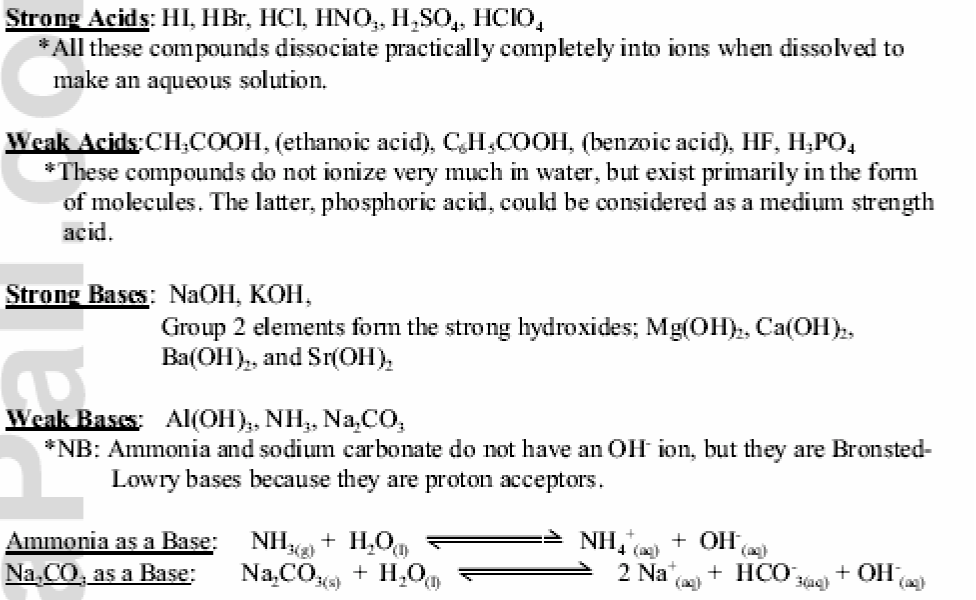
1. *A strong acid yields a weak conjugate base.*

2. *NEVER confuse the strength of an acid with its concentration.*

They are two different things. You can have a concentrated weak acid, or a diluted strong acid.

The strength and the concentration both affect the effeciveness of the acid, but you must remember that they ARE TWO SEPARATE FACTORS.

EXAMPLES OF ACIDS AND BASES



Strong Acids: HI, HBr, HCl, HNO_3, H_2SO_4, $HClO_4$
*All these compounds dissociate practically completely into ions when dissolved to make an aqueous solution.

Weak Acids: CH_3COOH, (ethanoic acid), C_6H_5COOH, (benzoic acid), HF, H_3PO_4
*These compounds do not ionize very much in water, but exist primarily in the form of molecules. The latter, phosphoric acid, could be considered as a medium strength acid.

Strong Bases: NaOH, KOH,
Group 2 elements form the strong hydroxides; $Mg(OH)_2$, $Ca(OH)_2$, $Ba(OH)_2$, and $Sr(OH)_2$

Weak Bases: $Al(OH)_3$, NH_3, Na_2CO_3
*NB: Ammonia and sodium carbonate do not have an OH^- ion, but they are Bronsted-Lowry bases because they are proton acceptors.

Ammonia as a Base: $NH_{3(g)} + H_2O_{(l)} \rightleftharpoons NH_4{}^+{}_{(aq)} + OH^-{}_{(aq)}$
Na_2CO_3 as a Base: $Na_2CO_{3(s)} + H_2O_{(l)} \rightleftharpoons 2\,Na^+{}_{(aq)} + HCO^-{}_{3(aq)} + OH^-{}_{(aq)}$

THE AUTO-IONIZATION OF WATER.

Important note: all constants refered to: K_c, K_w, K_a, and K_b are temperature-dependent. All discussions are assumed to be at 25 °C, i.e. standard temperature.

The following equation describes the reaction of water with itself (called autoprotolysis):

$H_2O + H_2O <==> H_3O^+ + OH^-$

The equilibrium constant for this reaction is written as follows:
$K_c = ([H_3O^+] [OH^-]) / ([H_2O] [H_2O])$

However, in pure liquid water, $[H_2O]$ is a constant value. To demonstrate this, consider 1000 mL of water with a density of 1.00 g/mL. This 1.00 liter (1000 mL) would weigh 1000 grams. This mass divided by the molecular weight of water (18.0152 g/mol) gives 55.5 moles. The "molarity" of this water would then be 55.5 mol / 1.00 liter or 55.5 M.

This is considered to be constant. Cross-multiplying the above equation gives:

$\mathbf{K_c [H_2O] [H_2O] = [H_3O^+] [OH^-]}$

Since the term $K_c [H_2O] [H_2O]$ is a constant, let it be symbolized by K_w, giving:

$\mathbf{K_w = [H_3O^+] [OH^-]}$

At 25 degrees Celsius, (standard temperature), the value of Kw is ALWAYS 1.0×10^{-14} mol^2dm^{-3}

The concentration of hydrogen ion (H+) in *pure* water is very small being:
1.0×10^{-7} dm^{-3}. The Hydroxide concentration in pure water is also
1.0×10^{-7} dm^{-3}

This means that in one L of water, one out of 10^7 molecules is a H_3O+ ion. Also, in one L of water, another one out of 10^7 molecules is a OH- ion!

pKa.... pKw.... (what's with the p?)

WHAT IS pKa? Or pKw?

Adding a ‘p’ in front of a Ka or Kw just makes life simpler.

I mean, if you had to write Kw = 1.0 X 10^{-14} all the time, it would be a lot of writing. And scientists are lazy.

So they just add a ‘p’ in front of the Kw.

$$pKw = -\log_{10} Kw.$$

$$pKw = -\log10\,(1.0*10^{-14})$$

$$pKw = 14.$$

So it’s much easier just to write pKw --> and say 14 instead of 1.0 * 10^{-14}

Because we’re dealing with logarithms here, we have to realize that:

THE SMALLER THE pKa, the LARGER the Ka --> The STRONGER the ACID.

The concept of pOH and pH

The pOH of a solution is the negative logarithm, to the base ten, of the hydroxide ion concentration.

$$pOH = -\log10[OH]$$

so pOH can range from 1 to 14. Don’t get it confused with pH, that’s something else.

PH is the concentration of HYDROGEN IONS in a solution,
pOH is the concentration of HYDROXIDE ions.

So let’s say you looked at a dissociation, and you’re like –

well the number of H_3O+ ions is 10^-5.

And the number of OH ions is 10^-9.

(look at page 14 for a better explanation of ionic concentration)

Just in case you didn’t get that, if you have 10^5 molecules, one of them is an H_3O+ ion. And one out of 10^9 molecules is an OH ion.

So that means you have way less OH ions than you do H_3O+ rite?

So your pOH is 9, and your pH is 5.

A pH of 5 is acidic. (you have more H3O+ ions)

Which makes complete sense. If you have MORE H+ ions (10^-5 is biger than 10^-9), you’ll have a lower pH – and you have an acidic solution.

(Oh by the way, treat H3O+ and H+ as the same thing, H3O+ just has to do with water, its more ‘proper’ to use it.)

Now always if you look at the two numbers 5 and 9, you’ll notice that they add to 14.

SO:

$$\mathrm{pH} + \mathrm{pOH} = \mathrm{p}K_w = 14.00 \qquad \text{at } 25^{\circ}\mathrm{C}$$

PH and pOH might change, but pKw always stays the same – 14.

After this, she just has stoichometry, which is exactly the same thing as we’ve always done – so....

CONGRATULATIONS! YOU’RE DONE THE ISU!

Now just look over the ISU package, just in case you missed anything.
By the way if you have any questions my email is:
white_shining_angel@hotmail.com