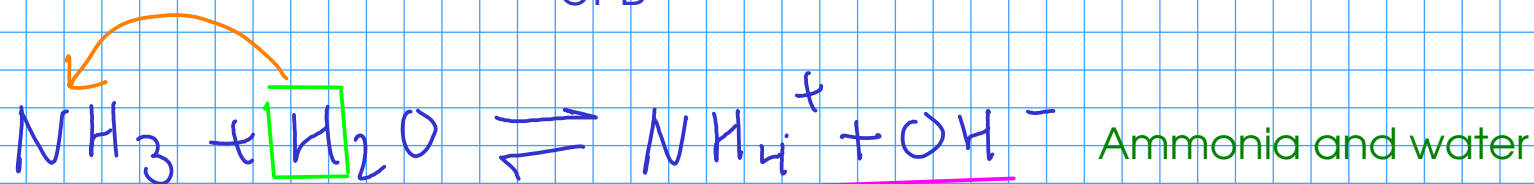
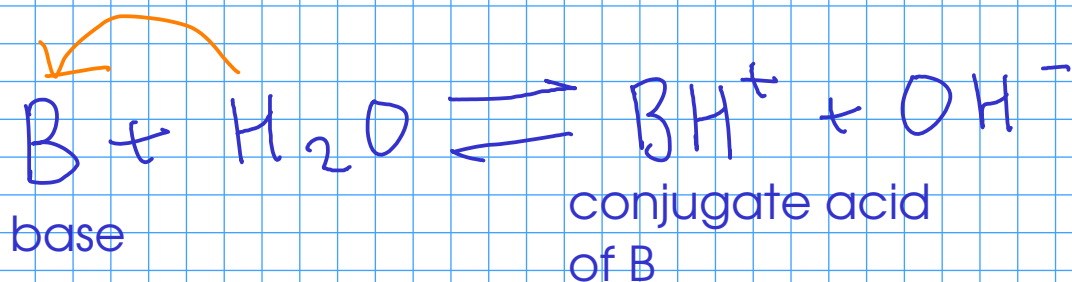
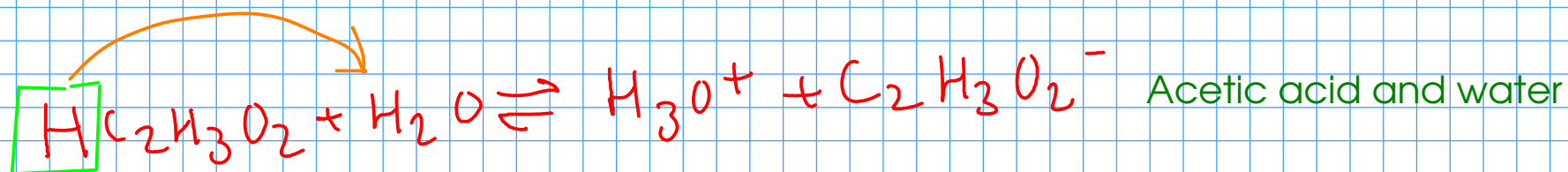
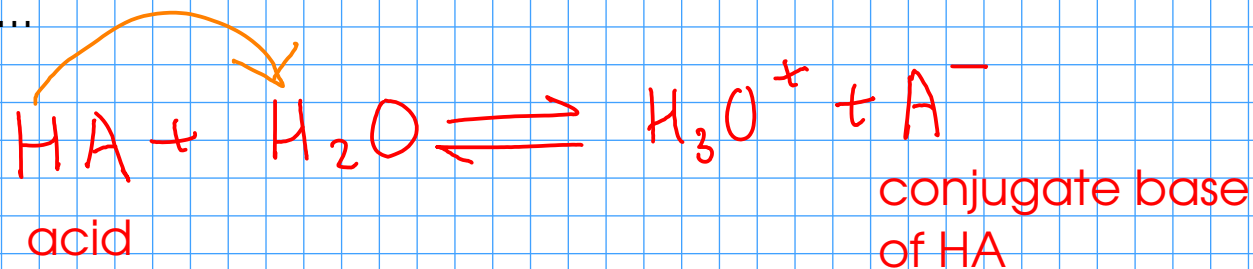


BRONSTED-LOWRY THEORY

IN WATER...



This is why we often call an ammonia/water solution "ammonium hydroxide"!

In the red reactions, water functions as a base. In the blue reactions, water functions as an acid!

LEWIS THEORY

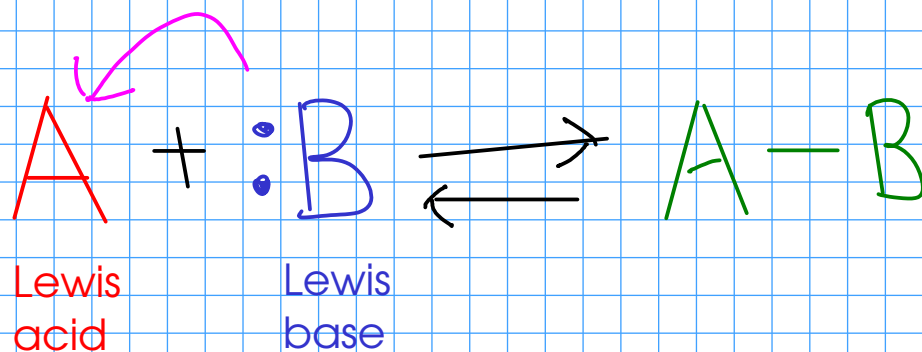
- Lewis theory treats acid-base chemistry as ELECTRON-TRANSFER chemistry involving pairs of electrons
- Lewis acid-base reactions form new covalent bonds (of interest to organic chemists!)

ACIDS are ACCEPTORS of electron pairs

... this is why some METAL IONS, even though they contain no hydrogen ions, can exhibit ACIDIC character. Many metal ions can accept a pair of electrons form a COMPLEX with a Lewis base!

BASES are DONORS of electron pairs.

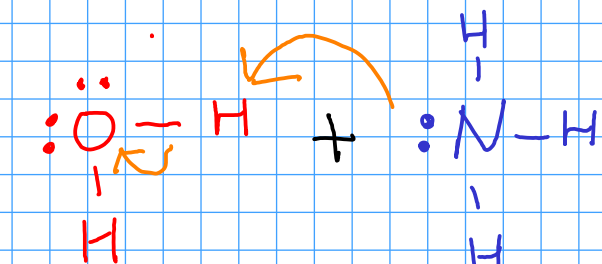
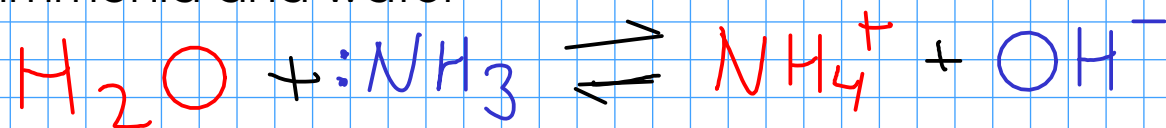
... so, Lewis bases have LONE PAIRS OF ELECTRONS in their Lewis structures



... In a Lewis acid-base reaction, electrons are donated from the Lewis base to the Lewis acid. This forms a new COVALENT BOND between the acid and the base.

LEWIS THEORY

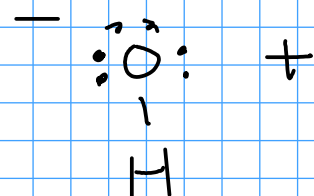
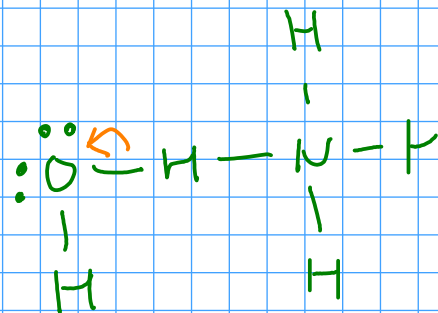
Example: ammonia and water



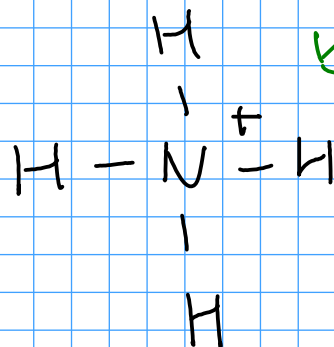
Ammonia donates a pair of electrons to a hydrogen on the water molecule, forming a new covalent bond (and breaking another)

water
(Lewis acid)

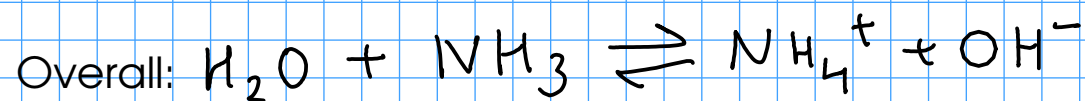
ammonia
(Lewis base)



hydroxide ion



ammonium ion



COMPARING THE THEORIES

- From Arrhenius to Lewis, the definitions get broader as you go along. In other words, the later definitions include MORE SUBSTANCES under the acid/base umbrella.

If something is an Arrhenius acid, it is also an acid in the Bronsted or Lewis picture.

If something is an Arrhenius base, it is also a base in the Bronsted or Lewis picture.

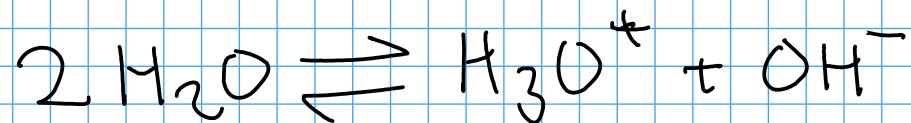
All Bronsted acids are Lewis acids, and all Bronsted bases are Lewis bases.

... but not all Lewis acids/bases (like the metal ions) are Bronsted or Arrhenius acids/bases.

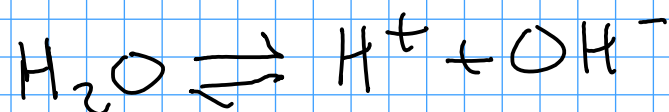
... We will primarily use the BRONSTED-LOWRY theory from this point in the course!

WATER CHEMISTRY

- Water self-ionizes!



or



This is an equilibrium reaction!

$$K = \frac{[\text{H}_3\text{O}^+][\text{OH}^-]}{[\text{H}_2\text{O}]^2}$$

(X) = molar concentration of "X"

In aqueous solution, (H_2O) is essentially constant, so we roll that into K.

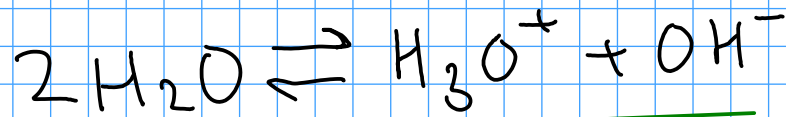
$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

WATER CHEMISTRY

- The self-ionization of water has a small equilibrium constant. What does this imply?

THE CONCENTRATION OF HYDROXIDE AND HYDRONIUM ION
IN PURE WATER IS VERY SMALL!

How small?



In pure water, the concentration of hydroxide and hydronium must be equal, since they are formed at the same time and at the same ratio from the ionization reaction of water.

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1 \times 10^{-14}$$

Solve... Let " z " = $[\text{H}_3\text{O}^+] = [\text{OH}^-]$

$$(z)(z) = 1 \times 10^{-14}$$

$$z^2 = 1 \times 10^{-14}$$

$$z = 1 \times 10^{-7} \text{ M} = [\text{H}_3\text{O}^+] = [\text{OH}^-]$$

"p" NOTATION

- "p" notation helps us deal with the very small numbers we encounter when working with acids, bases, and water.

- based on log base 10

"p" means $-\log_{10}$

log

~~ln~~

So,

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+]$$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

$$\text{pOH} = -\log_{10} [\text{OH}^-]$$

$$[\text{OH}^-] = 10^{-\text{pOH}}$$

"p" NOTATION

- Apply "p" notation to the water self-ionization reaction!

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14}$$

becomes ...

$$pK_w = pH + pOH = 14.00$$

Taking the "p" (negative log base ten) of the equilibrium constant is often used for BUFFER SOLUTIONS, which we'll discuss later!