

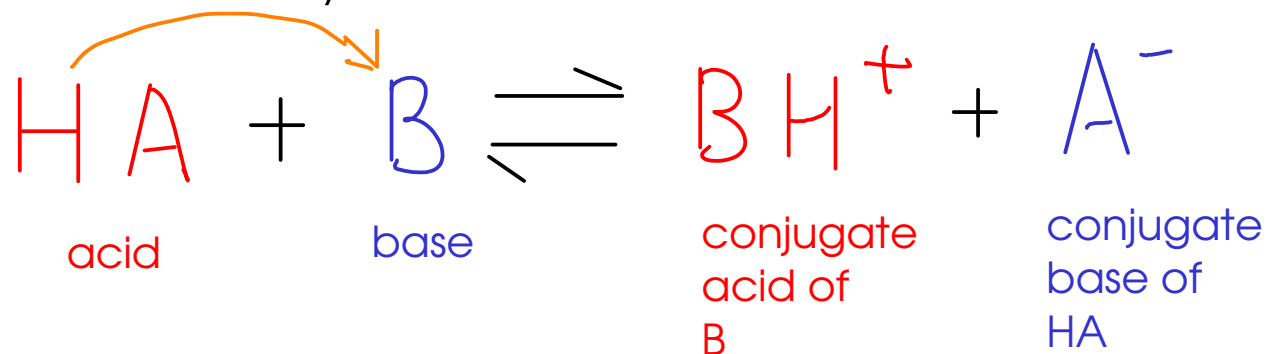
... a few examples of conjugate pairs:

Species	Conjugate
$\text{NH}_3$	$\text{NH}_4^+$
$\text{H}_2\text{O}$	$\text{OH}^-$
$\text{H}_2\text{O}$	$\text{H}_3\text{O}^+$
$\text{HC}_2\text{H}_3\text{O}_2$	$\text{C}_2\text{H}_3\text{O}_2^-$

RED for acid

BLUE for base

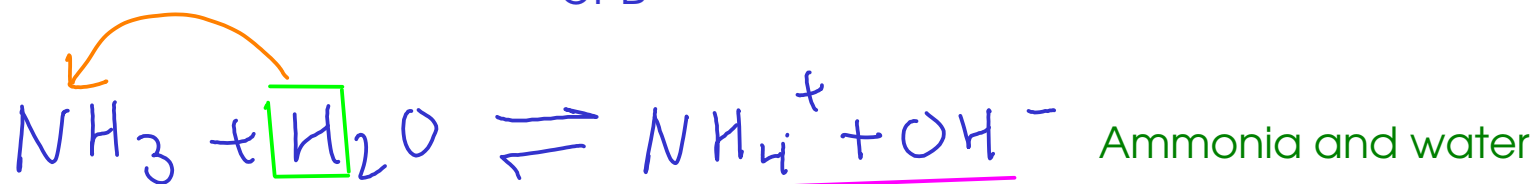
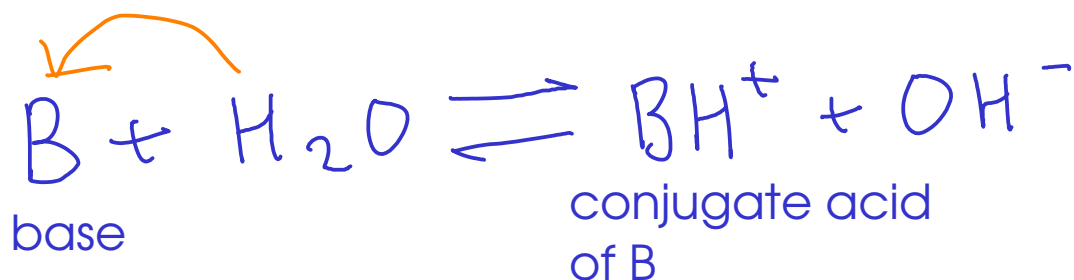
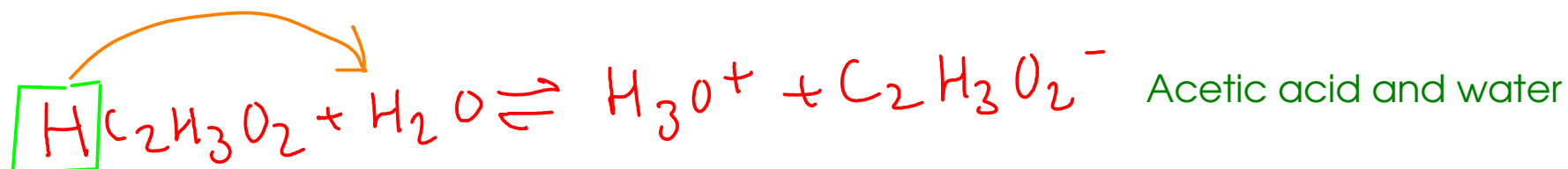
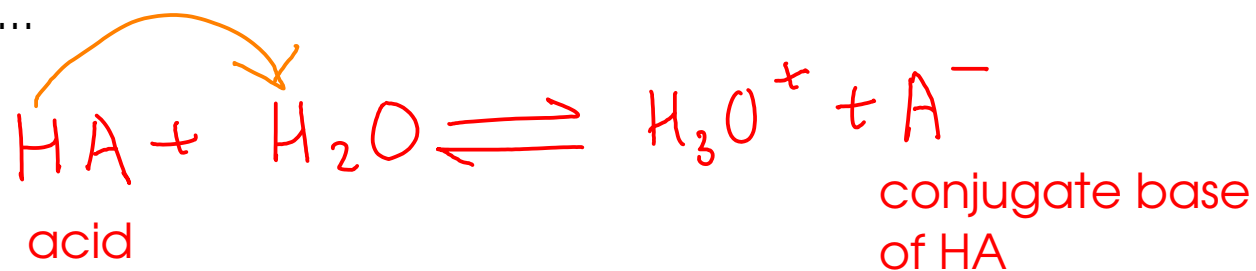
A generic Bronsted-Lowrey acid-base reaction:



... you should be able to write the products of a Bronsted-Lowry acid-base reaction, identifying the CONJUGATE PAIRS

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

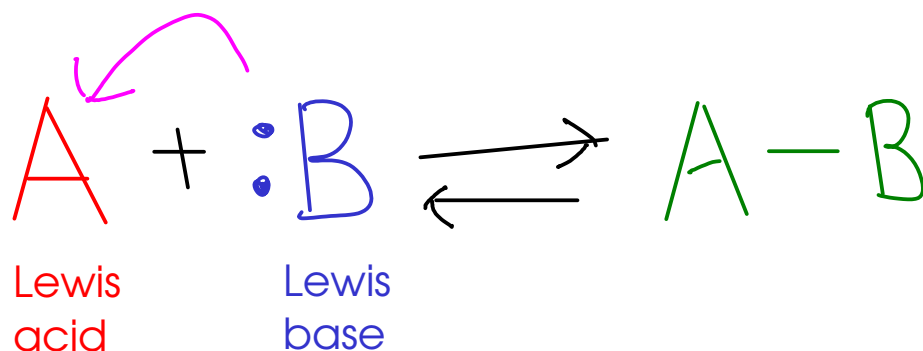
- Lewis theory treats acid-base chemistry as ELECTRON-TRANSFER chemistry involving pairs of electrons
- Lewis acid-base reactions form new covalent bonds (of interest to organic chemists!)

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 to form a COMPLEX with a Lewis base!

BASES are DONORS of electron pairs.

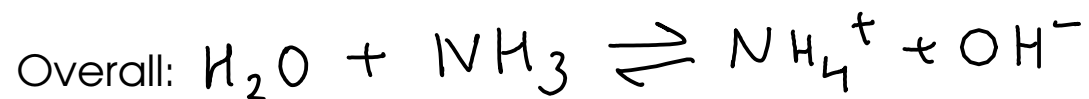
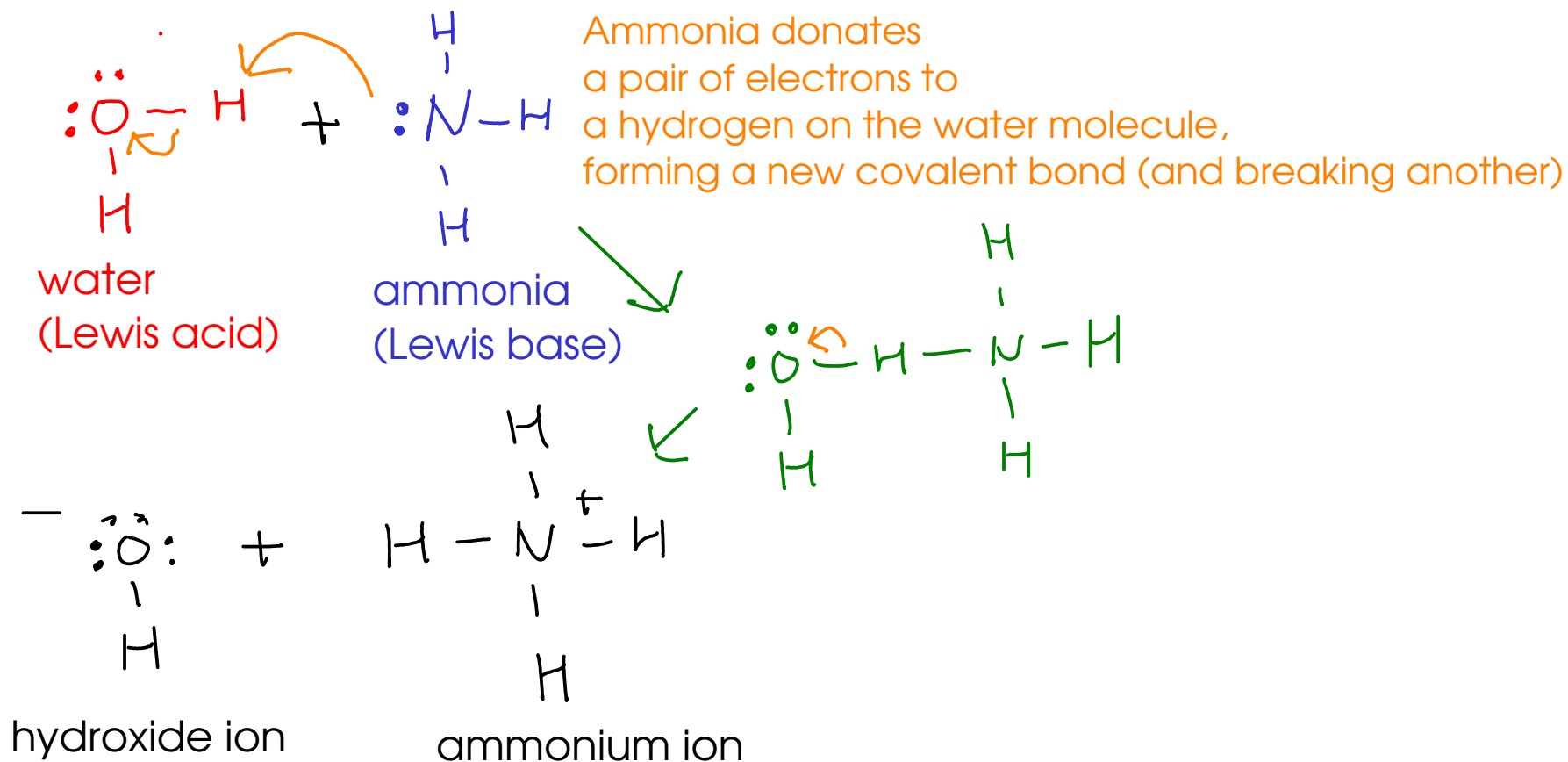
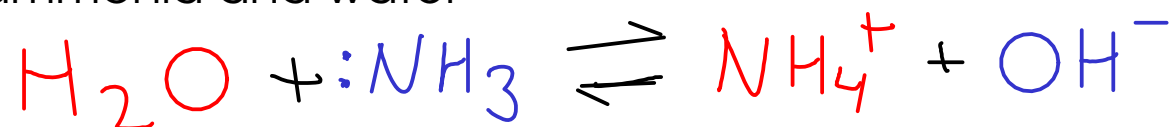
... so, Lewis bases have LONE PAIRS OF ELECTRONS in their Lewis structures



... 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



## 120 COMPARING THE THEORIES

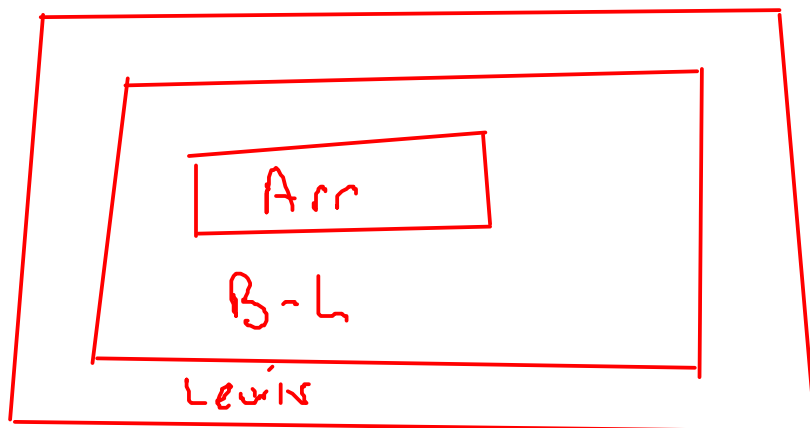
- From Arrhenius to Lewis, the definitions get broader as you go along. In other words, the later definitions include MORE SUBSTANCES under the acid/base umbrella.

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.

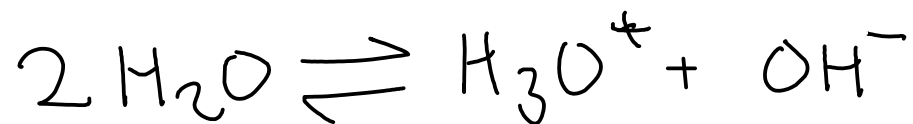
... but not all Lewis acids/bases (like the metal ions) are Bronsted or Arrhenius acids/bases.



... 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, ( $\text{H}_2\text{O}$ ) is essentially constant, so we roll that into K.

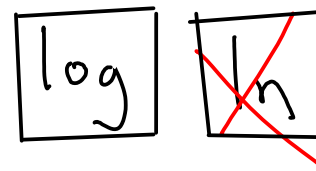
$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

## "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

On a calculator, use



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

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!



## ACIDITY AND ALKALINITY

- At  $\text{pH} = 7$ ,  $\text{pH} = \text{pOH}$ . The solution is considered NEUTRAL

↳ Also,  $[\text{H}^+] = [\text{OH}^-]$ !

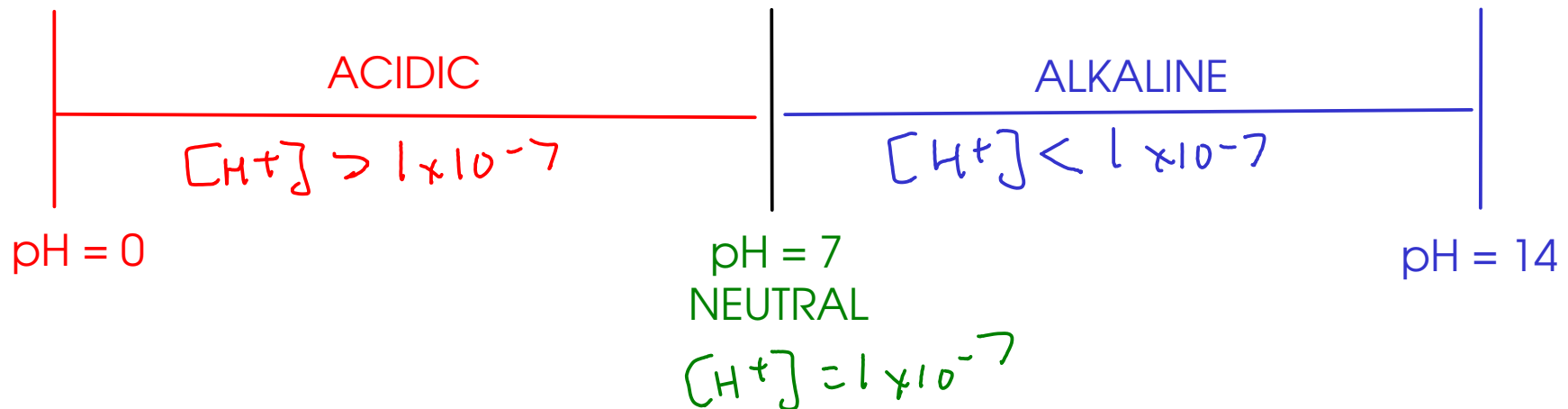
- At  $\text{pH} < 7$ ,  $\text{pH} < \text{pOH}$ . The solution is considered ACIDIC

↳ Also,  $[\text{H}^+] > [\text{OH}^-]$ !

- At  $\text{pH} > 7$ ,  $\text{pH} > \text{pOH}$ . The solution is considered ALKALINE (BASIC)

↳ Also,  $[\text{H}^+] < [\text{OH}^-]$ !

The pH scale...



## pH AND TEMPERATURE

$$pK_w = pH + pOH = 14.00$$

This equation is valid at room temperature, specifically 25°C.

Equilibrium constants depend on TEMPERATURE, and change with temperature.

So, the "neutral" pH (where the concentration of hydroxide and hydronium ions are equal) CHANGES with changing temperatures

This change is important at temperatures greatly different from 25°C.

As an example, consider average "normal" human body temperature: 37°C

$$\text{At } 37^\circ\text{C}, \quad pK_w = 13.60$$
$$pH \text{ of neutral solution} = \underline{\underline{6.8}}$$

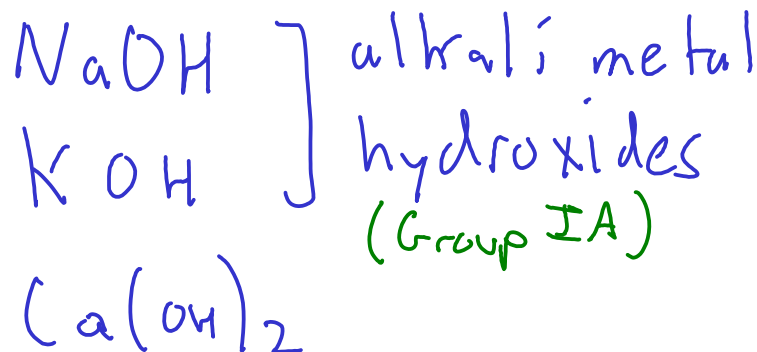
## ACID-BASE EQUILIBRIUM IN WATER

- Like other ELECTROLYTES, acids and bases IONIZE to some extent in water
- STRONG electrolytes ionize completely. Acids and bases that ionize completely in water are called STRONG ACIDS and STRONG BASES
- WEAK electrolytes ionize partially, remaining mostly non-ionized. Acids and bases that ionize only partially in solution are called WEAK ACIDS and WEAK BASES.
- Most acids and bases are WEAK!

### Common strong acids



### Common strong bases



## SIMPLE pH CALCULATIONS: STRONG ELECTROLYTES

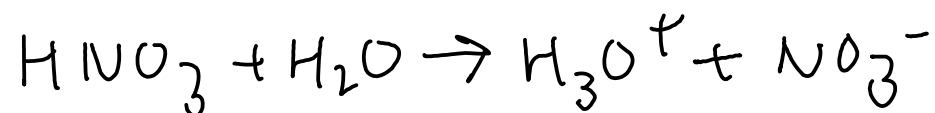
- With strong acids and bases, the acid or base completely ionizes in water. So, we only have to worry about the effect of the acid or base on the water equilibrium itself.

- Since the equilibrium constant for the self-ionization of water is so small, the strong acid or base will overpower the hydronium (for acids) or hydroxide (for bases) produced by the water.

Consider 0.025 M  $\text{HNO}_3$



Assume all  $\text{H}_3\text{O}^+$  comes from acid.!



So,  $[\text{H}_3\text{O}^+] \approx \text{nominal } [\text{HNO}_3] = 0.025 \text{ M } \text{H}_3\text{O}^+$

$$\text{pH} = -\log_{10}(0.025) = 1.60$$

Rounding: For logarithms, the places AFTER the decimal point are significant digits, while the numbers in front of the decimal are not (they're exponents)

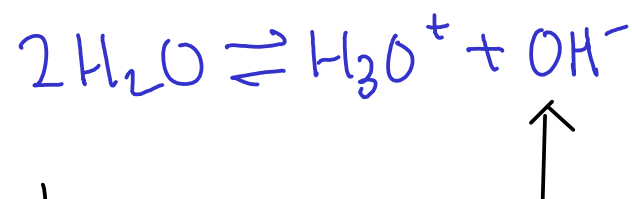
What would the pOH be?

$$\text{pH} + \text{pOH} = 14.00,$$

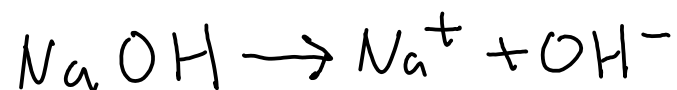
$$\text{pOH} = 12.40$$

... but we're usually not interested in pOH

Consider  $0.0125 \text{ M NaOH}$



Assume all  $\text{OH}^-$  comes from base



$$[\text{OH}^-] = \text{nominal } [\text{NaOH}] = 0.0125 \text{ M OH}^-$$

$$\text{pOH} = -\log_{10} [\text{OH}^-] = -\log_{10} (0.0125) = 1.90$$

... now change the pOH to pH so we can compare this to the acid problem we just worked.

$$\text{pH} + \text{pOH} = 14.00$$

$$\text{pH} + 1.90 = 14.00$$

$$\boxed{\text{pH} = 12.10}$$

Let's find the concentration of the hydronium ion, since that equals the concentration of HYDROXIDE PRODUCED BY THE WATER EQUILIBRIUM ITSELF. (We assumed that this number was much, much smaller than  $0.0125 \text{ M}$ !)

$$[\text{H}_3\text{O}^+] = 10^{-12.10} = 7.9 \times 10^{-13} \text{ M}$$