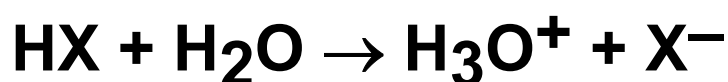


## ***Estimating pH***

- **Acids donate protons.**

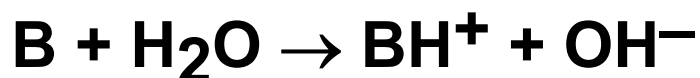
In H<sub>2</sub>O, the proton is donated to H<sub>2</sub>O, producing the hydronium ion.



Acidic solutions have high concentrations of H<sup>+</sup> (or H<sub>3</sub>O<sup>+</sup>) and low concentrations of OH<sup>-</sup>.

- **Bases accept protons.**

In H<sub>2</sub>O, the proton is donated by H<sub>2</sub>O, producing the hydroxide ion.



Basic solutions have low concentrations of H<sup>+</sup> (or H<sub>3</sub>O<sup>+</sup>) and high concentrations of OH<sup>-</sup>.

- $[H^+] \times [OH^-] = 1.0 \times 10^{-14}$  OR  
 $[H_3O^+] \times [OH^-] = 1.0 \times 10^{-14}$   
(Remember [ ] represent Molarities.)

- $p(\text{anything}) = -\log(\text{anything})$   
We generally use the pH scale.

$$pH = -\log(H^+) \quad \text{and} \quad pOH = -\log(OH^-)$$

$$pH + pOH = 14$$

- Acids have low pH values (0 – 6.9) and  
Bases have high pH values (7.1 – 14).
- An indicator is a compound which  
upon gaining or losing a  $H^+$  changes  
colors.

An indicator changes colors over a narrow pH range, such as phenolphthalein at pH's of 8.5-10.

A universal indicator has many compounds which together turn different colors over a wide pH range, such as red cabbage extract at pH's of 2-12.