Estimating pH

- ***Acids donate protons*.**

  In H_2O, the proton is donated to H_2O, producing the hydronium ion.

  $$HX + H_2O \rightarrow H_3O^+ + X^-$$

  Acidic solutions have high concentrations of H^+ (or H_3O^+) and low concentrations of OH^-.

- ***Bases accept protons*.**

  In H_2O, the proton is donated by H_2O, producing the hydroxide ion.

  $$B + H_2O \rightarrow BH^+ + OH^-$$

  Basic solutions have low concentrations of H^+ (or H_3O^+) and high concentrations of OH^-.

- **$[H^+] \times [OH^-] = 1.0 \times 10^{-14}$ OR**

  $[H_3O^+] \times [OH^-] = 1.0 \times 10^{-14}$

  (Remember [ ] represent Molarities.)

- **p(anything) = – log (anything)**
  We generally use the pH scale.

  $pH = -\log (H^+)$ and $pOH = -\log (OH^-)$

  $pH + pOH = 14$

- **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 phenol-phthalein 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.