- **Autoionization of Water**

$$H_2O(l) + H_2O(l) \rightleftharpoons OH^-(aq) + H_3O^+(aq)$$

  - Equilibrium expression
    - $K_c = [H_3O^+][OH^-] = \boldsymbol{K_w}$
      - K_w is the **ion product constant** for water.
      - At 25 °C, $K_w = 1.0 \times 10^{-14}$
  - Ion Product Constant
    - $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$
      - When $[H_3O^+] = [OH^-]$, the solution is **neutral**.
      - When $[H^+] > [OH^-]$, the solution is **acidic**.
      - When $[H^+] < [OH^-]$, the solution is **basic**.
    - Eg. At 25 °C, $[H_3O^+] = 1.0 \times 10^{-14}$, $[OH^-]$ =?

      $[OH^-] * 1.0 \times 10^{-2} = 1.0 \times 10^{-14}$

      $[OH^-] = 1.0 \times 10^{-14}$ (acidic, $[H_3O^+]$ has a higher concentration)

- **pH Scale**
  - In pure water at 25 °C, $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ M
  - **pH = –log[H⁺]** $\mathbf{pH = -\log[H^+]}$
  - At 25 °C: Neutral, pH = 7.00

    Acidic, pH < 7.00

    Basic, pH > 7.00
- **pOH Scale**
  - **pOH = –log[OH⁻]**
  - At 25 °C:

    Neutral, pOH = 7.00

    Acidic, pOH > 7.00

    Basic, pOH < 7.00

- **Relationship between pH and pOH**
  - pH + pOH = pK_w = 14.00

- **In class question**: The autoionization of water is endothermic. What happens to K_w as temperature increases? What happens to the pH of pure water as temperature increases?
  - K_w depends on temperature: K_w increases when T increases
  - pH decreases when K_w increases

- **Measuring pH**
  - pH meters
  - Indicators: molecules that are one color in the acidic form and another color in the basic form.

- **Strong Acids**
  - 7 strong acids: HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_3$, and $HClO_4$.
  - Strong acids are virtually completely dissociated in aqueous solution
  - Assumptions
    - For the monoprotic strong acids, assume $[H_3O^+]$ = [acid]
    - Ignore autoionization most of the time for strong acid problems
      - Special Case: You add HCl to a beaker of water so that [HCl] = 1.0×10^{-9} M
      - pH is around 7 (adding up $1.0 \times 10^{-7} + 1.0 \times 10^{-9} = 1.0 \times 10^{-7}$)

- **Strong Bases**
  - Soluble hydroxides
    - the alkali metal hydroxides (e.g., NaOH)
    - the heavier alkaline earth metal hydroxides [e.g., $Ca(OH)_2$].
  - these substances dissociate completely in aqueous solution
  - eg. Add 4.0g of NaOH to water, so that the final volume is 1.0L. What is the pH? (assuming T = 25 °C)
    NaOH(aq) ⇌ Na+(aq) + OH-(aq)
    4.0g * (1 mol / 40 g) = 0.10 mol/L
    [NaOH] = $1*10^{-1}$ = $[OH^-]$
    pOH = 1, pH = 13

- **Comparing Strong and Weak Acids/Bases**
  - Strong acids/bases *completely* dissociate to ions.
  - Weak acids/bases only *partially* dissociate to ions.

- **How Strong is a Weak Acid?**
  - The equation for dissociation: $HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$
  - The equilibrium constant for this reaction is called the **acid-dissociation constant**, K_a: $K_a = [H_3O^+][A^-] / [HA]$