Warm-UP:

The More Things Change . . .

Le Châtlier's principle states that if the equilibrium of a system is disturbed by a change in concentration, pressure, or temperature, the system will adjust itself to return to equilibrium. The disturbances can be described as either increases or decreases in concentration, pressure, or temperature. By observing the way these disturbances change the equilibrium of a reaction, we can see how such changes should affect any reaction. For example, increasing the temperature of an endothermic reaction increases the K_{eq} for the reaction.

What happens to the equilibrium when you decrease the temperature of an exothermic reaction?

- pH = "parts Hydrogen" → Hydrogen ion concentration
  - ↳ Molarity is too small, so we use pH logarithmic scale
  - ↳ $[H^+]$ = Molarity of H^+
  - ↳ $pH = -\log[H^+]$ → Used for Acids
    $pOH = -\log[OH^-]$ → Used for Bases
  - ↳ Ex 1: If $[H^+] = 1 \cdot 10^{-10}$, what is your pH?
    - ↳ pH = 10
- pH scale = 1 – 14 → 1–7 = Acidic; 7–14 = Basic
  - ↳ $14 = pH + pOH$
- What is the pOH if your pH is 10?
  - ↳ $14 = 10 + pOH$ → $pOH = 4$
- Eq to find $[H^+]$: $10^{-pH} = [H^+]$; $10^{-pOH} = [OH^-]$

Ex 2: If your pH is 6.00, find pOH, $[H^+]$, & $[OH^-]$

- $14 = 6.00 + pOH$ → $pOH = 8.00$
- $10^{-6} = [H^+] = 1 \cdot 10^{-6} M$
- $10^{-8} = [OH^-] = 1 \cdot 10^{-8} M$

- Ex 3: What is the pH of a solution w/ 25g HCl in 1.5 L of water?

$$\frac{25\text{g}}{1} \left| \frac{1\text{mol}}{36.46\text{g}} \right| \frac{1}{1.5\text{L}} = 0.457 = [H^+]$$

$-\log[0.457] = 0.340 = \text{Acidic}$

Ex 4: A solution has a pH of 4.0. If you dilute it from 10 mL to 1000 mL, what is the new pH?

$M_1V_1 = M_2V_2$

$[H^+] = 10^{-4} = 1 \cdot 10^{-4} = M_1$

$1 \cdot 10^{-4} \cdot 10\,mL = M_2 \cdot 1000\,mL$

$M_2 = 1 \cdot 10^{-6}$

$pH = -\log[1 \cdot 10^{-6}] = \boxed{6.0}$