Outline

- Acids and Bases
- Acid and Base Reactions
- Aqueous Solutions
- Ionization of Water
- Buffers

Acids and Bases

Brønsted – Lowry acids are substances capable of donating hydrogen ions (H^+)

Brønsted – Lowry bases are substances capable of accepting hydrogen ions (H^+)

If water accepts a hydrogen ion it forms hydronium ion

$$H_2O + H^+ \rightarrow H_3O^+$$

An acid reacts with a base by transferring a hydrogen ion from the acid to the base

$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$

acid base c. acid c. base

When an acid loses a hydrogen ion, it becomes a <u>conjugate base</u>

When a base gains a hydrogen ion, it becomes a <u>conjugate acid</u>

$NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$

base acid c. acid c. base

$HCl(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$

acid base c. acid c. base

Acid and Base Reactions

Acids usually react with metals to produce hydrogen gas…

$2HCl + Zn \rightarrow H_2 + ZnCl_2$

$3H_2SO_4 + 2Al \rightarrow 3H_2 + Al_2(SO_4)_3$

Acids react with metal oxides to produce water…

$2HCl + Na_2O \rightarrow H_2O + 2NaCl$

$2H_3PO_4 + 3CaO \rightarrow 3H_2O + Ca_3(PO_4)_2$

Sodium hydroxide reacts with aluminum to produce hydrogen gas!

Acids and bases react to produce salts (ionic compounds) in neutralization reactions

Amount of acid or base in solution is determined in a titration…

Acid or base is slowly added to until indicator in solution changes color (end point)

The end point indicates location of equivalence point…

mol acid = mol base at the equivalence point!

36.0 mL of 0.105 M NaOH are needed to neutralize 25.0 mL of HCl. What's the HCl molarity?

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

$$36.0\ mL \Rightarrow 0.0360\ L \quad \text{and} \quad 25.0\ mL \Rightarrow 0.0250\ L$$

$$0.0360\ L \times \frac{0.105\ M\ NaOH}{L} \times \frac{1\ mol\ HCl}{1\ mol\ NaOH} = 0.00378\ mol\ HCl$$

$$\frac{0.00378\ mol\ HCl}{0.0250\ L} = \underline{0.151\ M\ HCl}$$

45.5 mL of 0.105 M NaOH are needed to neutralize 36.0 mL of H_2SO_4. What's the H_2SO_4 molarity?

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

$$45.5\,\text{mL} \Rightarrow 0.0455\,\text{L} \quad \text{and} \quad 36.0\,\text{mL} \Rightarrow 0.0360\,\text{L}$$

$$0.0455\,\text{L} \times \frac{0.105\,\text{M NaOH}}{\text{L}} \times \frac{1\,\text{mol } H_2SO_4}{2\,\text{mol NaOH}} = 0.0023\underline{8}8\,\text{mol } H_2SO_4$$

$$\frac{0.0023\underline{8}8\,\text{mol } H_2SO_4}{0.0360\,\text{L}} = \underline{0.0663\,\text{M } H_2SO_4}$$

How many mL's of 0.25 M NaOH are needed to neutralize 36 mL of 1.0 M H_3PO_4?

$$H_3PO_4(aq) + 3NaOH(aq) \rightarrow Na_3PO_4(aq) + 3H_2O(l)$$

$$36.0\ \text{mL} \Rightarrow 0.0360\ \text{L}$$

$$0.0360\ \text{L} \times \frac{1.0\ \text{mol}\ H_3PO_4}{\text{L}} \times \frac{3\ \text{mol NaOH}}{1\ \text{mol}\ H_3PO_4} = 0.1\underline{0}8\ \text{mol NaOH}$$

$$0.1\underline{0}8\ \text{mol NaOH} \times \frac{1\ \text{L}}{0.25\ \text{mol NaOH}} \times \frac{1000\ \text{mL}}{1\ \text{L}} = \underline{430\ \text{mL NaOH}}$$

How many mL's of 0.18 M HNO_3 are needed to neutralize 48 mL of 2.0 M $Ba(OH)_2$?

$$2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O(l)$$

$$48\ \text{mL} \Rightarrow 0.048\ \text{L}$$

$$0.048\ \text{L} \times \frac{2.0\ \text{mol}\ Ba(OH)_2}{\text{L}} \times \frac{2\ \text{mol}\ HNO_3}{1\ \text{mol}\ Ba(OH)_2} = 0.1\underline{9}2\ \text{mol}\ HNO_3$$

$$0.1\underline{9}2\ \text{mol}\ HNO_3 \times \frac{1\ \text{L}}{0.18\ \text{mol}\ HNO_3} \times \frac{1000\ \text{mL}}{1\ \text{L}} = \underline{1100\ \text{mL}\ HNO_3}$$

Aqueous Solutions

Salts (ionic compounds) dissolve in a process called dissociation

Acid molecules are “ripped” into ions when dissolved into water

$$HCl(g) \rightarrow H^{+}(aq) + Cl^{-}(aq)$$

This process is called ionization

Acids that ionize completely are strong acids

HCl, HBr, HI, H_2SO_4, HNO_3, $HClO_4$

Acids that ionize partially are weak acids

An <u>electrolyte</u> is a substance that when dissolved in water produces a solution that can conduct electricity

Strong electrolytes: completely ionized

Weak electrolytes: partially ionized

Nonelectrolytes: do not ionize

Strong acids are strong electrolytes

Bases acting as strong electrolytes are <u>strong bases</u>

alkali metal hydroxides, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

In net-ionic equations…

Strong electrolytes will be soluble salts, strong acids, and strong bases…

Write in “ionic form”!

Weak electrolytes will be insoluble salts, weak acids, and weak bases…

Write in “molecular form”!

Nonelectrolytes will be soluble molecular compounds and water…

Write in “molecular form”!

Ionic or molecular form?

HNO_3	ionic form	H^+ and NO_3-
$BaCl_2$	ionic form	Ba^{2+} and Cl^-
CO_2	molecular form	CO_2
NaOH	ionic form	Na^+ and OH^-
NH_3	molecular form	NH_3
HF	molecular form	HF

Ionization of Water

Water molecules ionize to a very small extent…

$$H_2O(l) \rightleftarrows H^+(aq) + OH^-(aq)$$

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

Equal amounts of H_3O^+ and OH^- are produced…

$[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ M in pure water at 25 ^oC

The product of $[H_3O^+]$ and $[OH^-]$ is constant…

$$[H_3O^+] \times [OH^-] = (1.0 \times 10^{-7}) \times (1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

Acidity of a solution depends on amount of hydronium...

If $[H_3O^+] = [OH^-]$, then solution is neutral

If $[H_3O^+] > [OH^-]$, then solution is acidic

If $[H_3O^+] < [OH^-]$, then solution is basic

Basic or acidic if $[H_3O^+] = 3.2 \times 10^{-9}$ M?

$$[OH^-] = \frac{1.0 \times 10^{-14}}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{3.2 \times 10^{-9}} = 3.1 \times 10^{-6} \text{ M}$$

$$[H_3O^+] < [OH^-] \quad \Rightarrow \quad \underline{\text{Basic}}$$

pH scale used to provide a convenient way to represent solution acidity

$pH = -\log[H_3O^+]$

pH of pure water $$pH = -\log(1.0 \times 10^{-7}) = 7.00$$

if $[H_3O^+] = 1.45 \times 10^{-11}$ M $$pH = -\log(1.45 \times 10^{-11}) = 10.839$$

if $[H_3O^+] = 4.2 \times 10^{-5}$ M $$pH = -\log(4.2 \times 10^{-5}) = 4.38$$

if $[OH^-] = 1.62 \times 10^{-8}$ M $$pH = -\log\left(\frac{1.0 \times 10^{-14}}{1.62 \times 10^{-8}}\right) = 6.21$$

Acidity of a solution depends on the pH…

If pH = 7.00, then solution is neutral

If pH $<$ 7.00, then solution is acidic

If pH $>$ 7.00, then solution is basic

Basic or Acidic?

Lemon juice, pH = 2.3	acidic
Household ammonia, pH = 11.0	basic
Urine, pH = 6.0	acidic
Blood, pH = 7.4	basic

The hydronium ion concentration is found from pH using antilog…

$$pH = -\log[H_3O^+] \Rightarrow -pH = \log[H_3O^+] \Rightarrow 10^{-pH} = [H_3O^+]$$

What's $[H_3O^+]$ if…

pH = 3.00 $\quad [H_3O^+] = 10^{-3.00} = 1.0 \times 10^{-3}$ M

pH = 11.0 $\quad [H_3O^+] = 10^{-11.00} = 1 \times 10^{-11}$ M

pH = 14.2 $\quad [H_3O^+] = 10^{-14.2} = 6 \times 10^{-15}$ M

pH = 1.436 $\quad [H_3O^+] = 10^{-1.436} = 0.0366$ M

Buffers

Buffers are solutions that resist changes in pH

Buffers are solutions that contains significant amounts of a weak acid (HA) and its conjugate base (A^-)

Buffer or Not?

HF and F^-	Yes!
H_2CO_3 and CO_3^{2-}	No!
H_3PO_4 and $H_2PO_4^-$	Yes!
HCl and Cl^-	No!
HPO_4^{2-} and $H_2PO_4^-$	Yes!
H_2SO_3 and HSO_4^-	No!

The weak acid neutralizes added strong base (OH^-):

$$HA + OH^- \rightarrow A^- + H_2O$$

The conjugate base neutralizes added strong acid (H_3O^+):

$$A^- + H_3O^+ \rightarrow HA + H_2O$$

Resistance to pH change increases with increasing amounts of the weak acid and conjugate base