Acids and Bases

Ms. Peace

Lesson 1

Theories of Acids and Bases

Main
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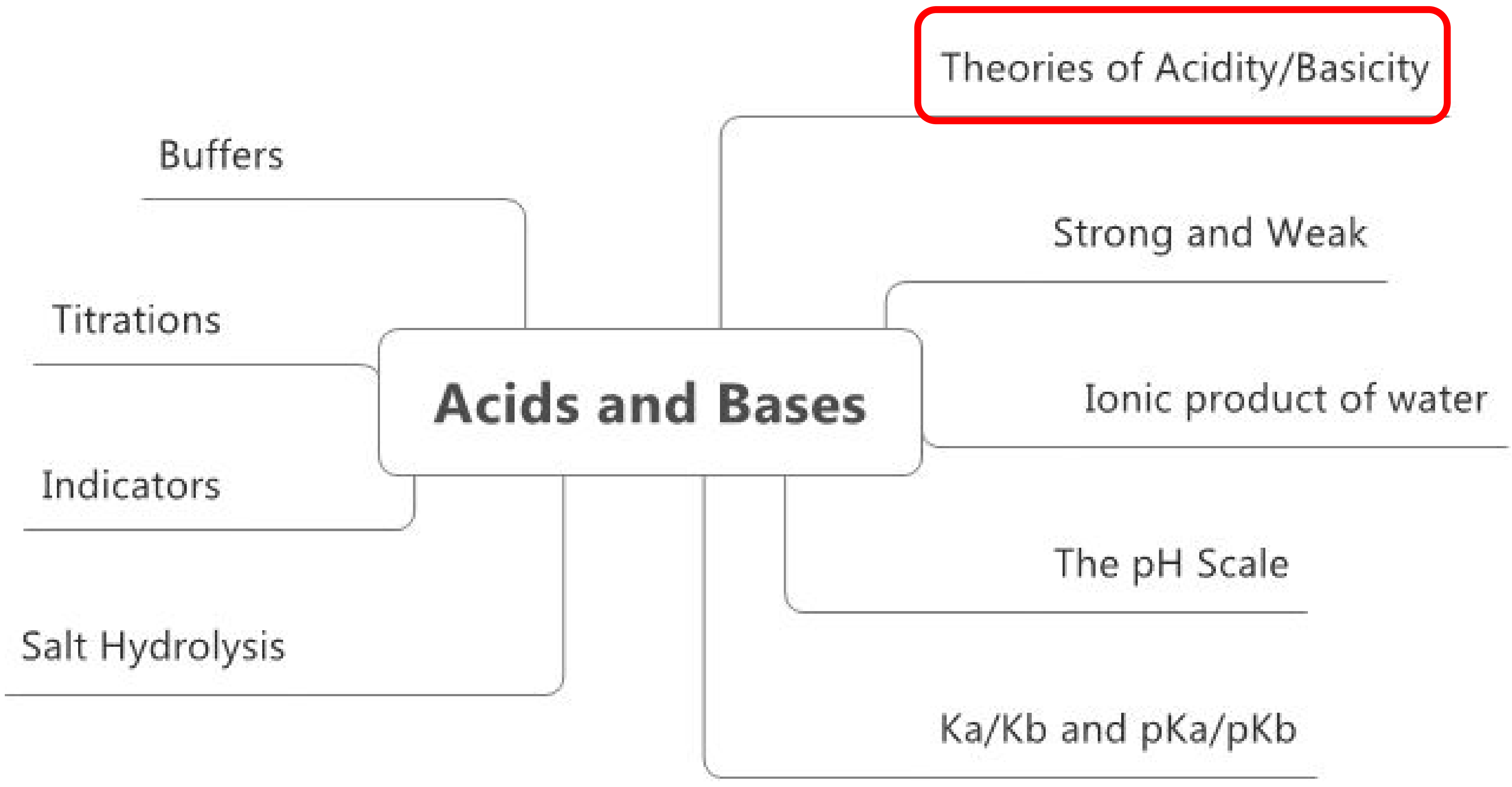
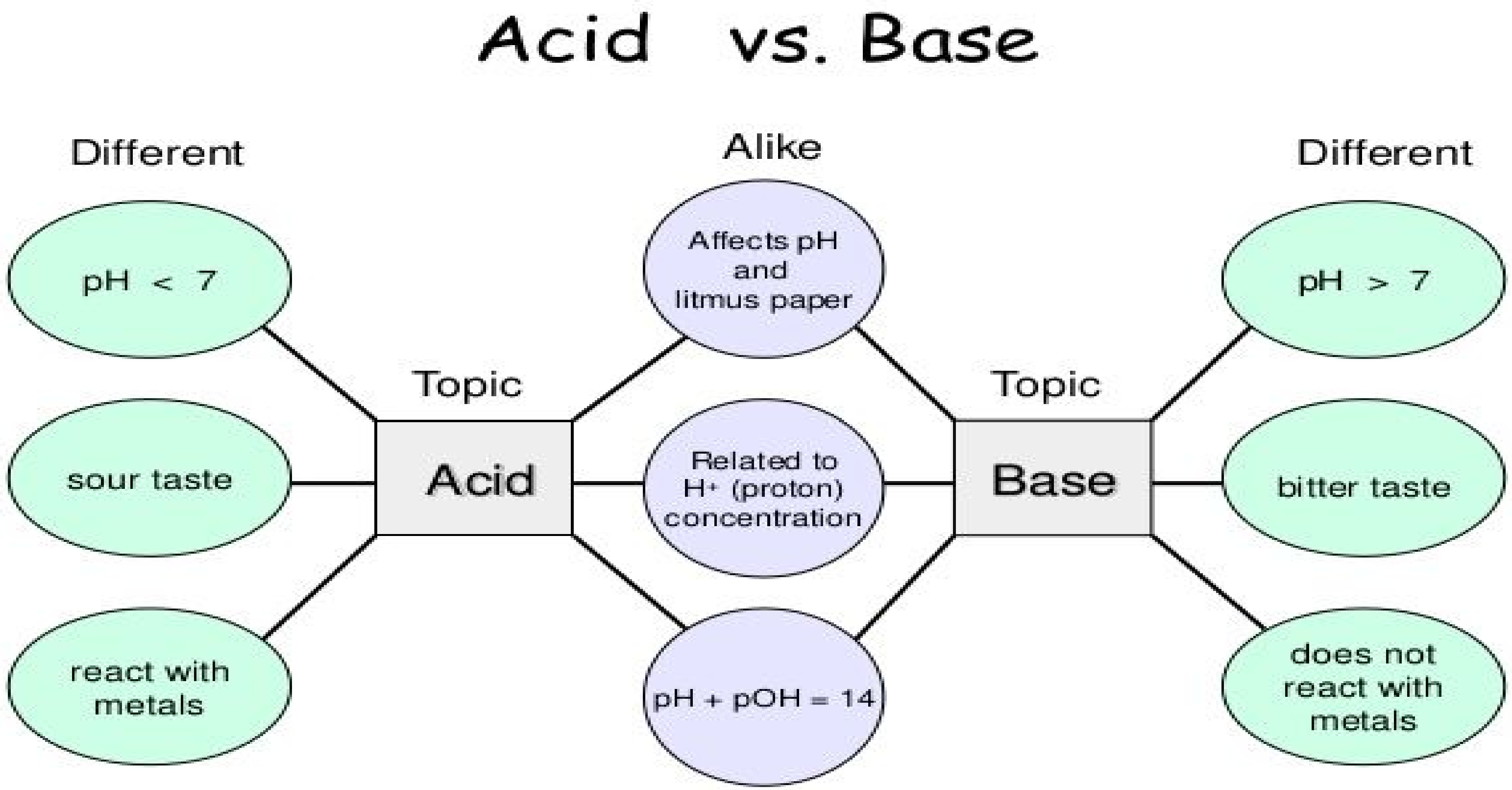


Main

Acids vs Bases



Arrhenius Theory

- Svante August Arrhenius defined an acid as a substance that ionizes in water to produce hydrogen ions, H^+
- An alkali, a soluble base, produces hydroxide ions, OH^-
- The combination of an acid and base is neutralization

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

Arrhenius Theory

- Arrhenius theory has its limitations

- The reaction between the weak base, ammonia, and hydrogen chloride gas could not be explained, as ammonia does not contain hydroxide ions

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

Brønsted-Lowry Theory

- Johannes Brønsted and Thomas Lowry developed a definition of acids and bases that broadened Arrhenius's theory
- Referring to a hydrogen ion as a proton they proposed that an acid could be defined as a proton donor and a base a proton acceptor
- In an aqueous solution, a proton can be represented as either the hydrogen ion H^+ or the hydronium ion H_3O^+

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

Main

Brønsted-Lowry Acids and Bases

- It's all about protons (H^+)

- **Acid:** Proton donor
  - $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
  - $H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$

- **Base:** Proton acceptor
  - $NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$
  - $OH^-(aq)^* + H^+(aq) \rightarrow H_2O(l)$

*From any soluble hydroxide or other alkali

- If we mention acid/base without mentioning the type, we generally mean a Bronsted-Lowry one.

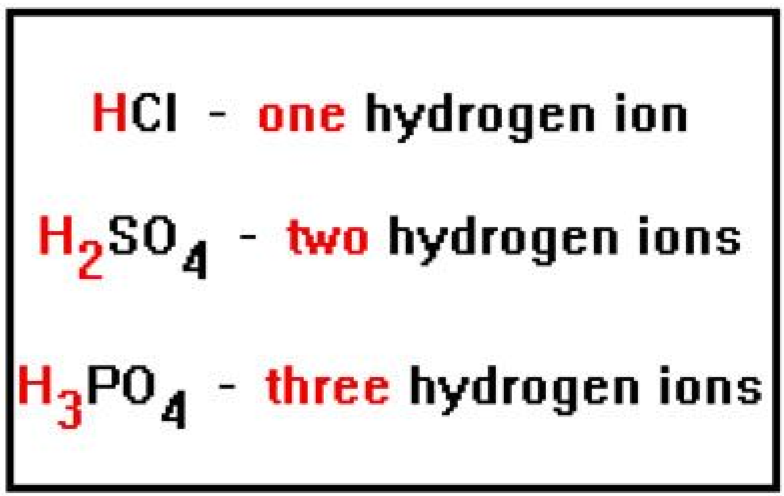
Common Acids

Acid	Formula
Hydrochloric	HCl
Sulfuric	H_2SO_4
Nitric	HNO_3
Carbonic	H_2CO_3
Ethanoic	CH_3COOH
Benzoic	C_6H_5COOH

The H that is lost as the H^+ is shown in red

Dissociation of Acids

- Monoprotic:
  - $HCl(aq) \rightarrow H^{+}(aq) + Cl^{-}(aq)$
- Diprotic:
  - $H_2SO_4(aq) \rightarrow 2H^{+}(aq) + SO_4^{2-}(aq)$
- Triprotic:
  - $H_3PO_4(aq) \rightleftarrows 3H^{+}(aq) + PO_4^{3-}(aq)$



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Dissociation of Acids

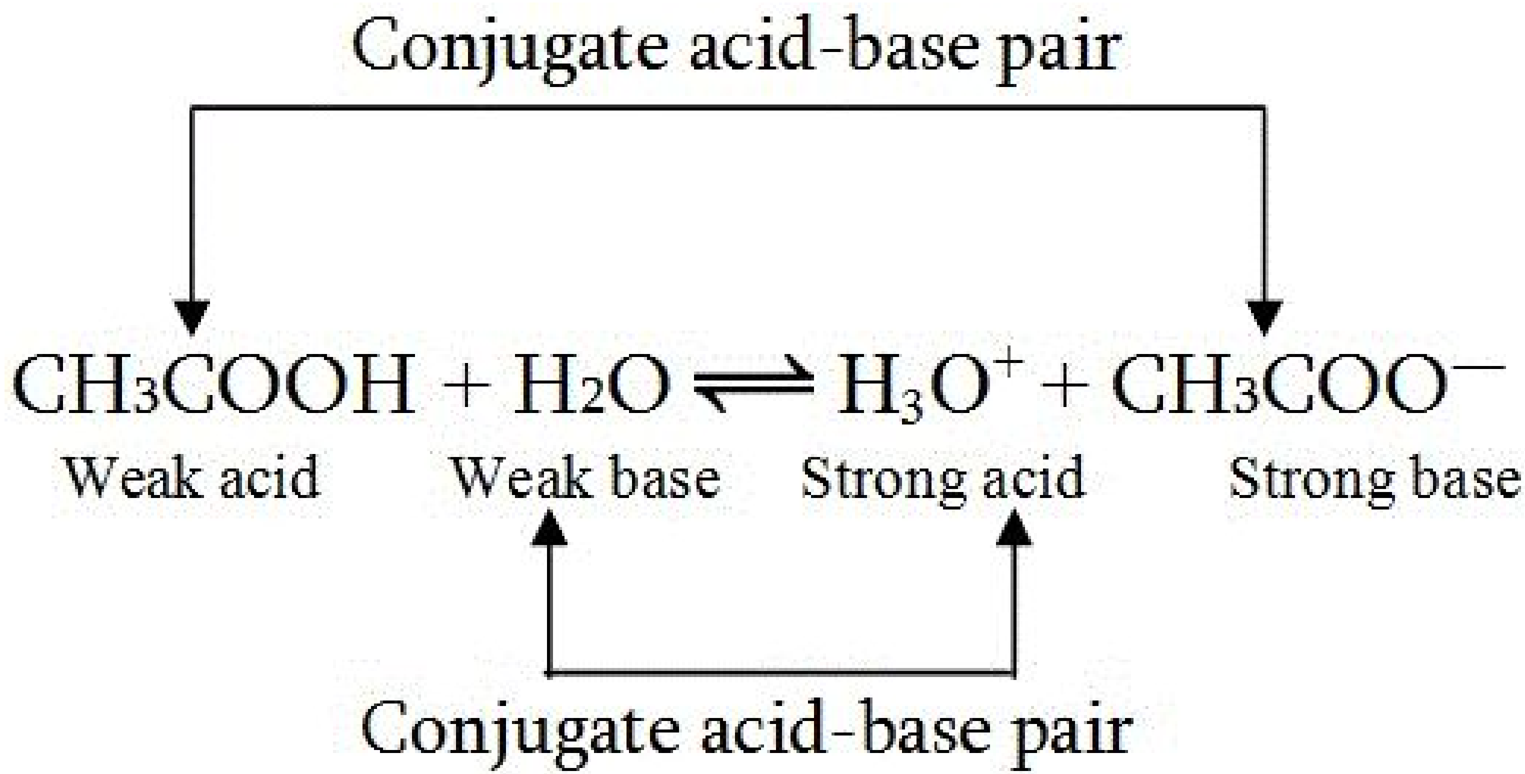
- Ethanoic Acid:
  - $CH_3COOH(aq) \rightleftarrows CH_3COO^-(aq) + H^+(aq)$

OR

  - $\mathbf{CH_3COOH(aq) + H_2O(l) \rightleftarrows CH_3COO^-(aq) + H_3O^+(aq)}$

- In the forward reaction:
  - $CH_3COOH(aq)$: Brønsted-Lowry acid
  - $H_2O(l)$: Brønsted-Lowry base
- In the reverse reaction:
  - $CH_3COO^-(aq)$: Proton acceptor
  - $H_3O^+(aq)$: Proton donor

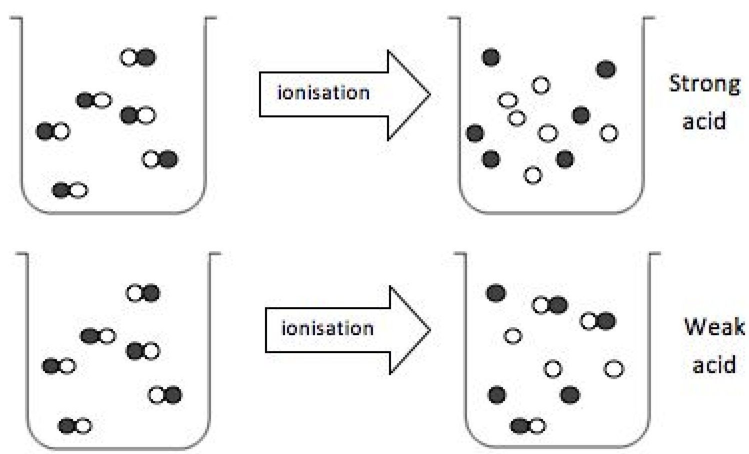
Conjugate Acid-Base Pair



Conjugate acid-base pairs differ by a single proton

Strong vs Weak Acids

- A strong acid is assumed to undergo complete dissociation in water
- A weak acid undergoes only partial dissociation in water, establishing an equilibrium, and a solution of a weak acid is only a weak electrolyte



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Time to practice

- Give the formula of the conjugate base for each of the following:
  1. HF
  2. H_2SO_4
  3. H_3PO_4
  4. CH_3COOH
  5. H_2O
  6. NH_4^+

- Give the formula of the conjugate acid for each of the following:
  1. OH^-
  2. SO_4^{2-}
  3. HPO_4^{2-}
  4. $(CH_3)_2NH$
  5. H_2O

Amphiprotic Species

- Amphiprotic species are substances that have the ability to act as either a Brønsted-Lowry acid or a Brønsted-Lowry base

- Polyprotic species are usually involved in reactions that behave amphiprotically

$$H_2PO_4^-(aq) + OH^-(aq) \rightleftarrows HPO_4^{2-}(aq) + H_2O(l)$$

$$H_2PO_4^-(aq) + H_3O^+(aq) \rightleftarrows H_3PO_4(aq) + H_2O(l)$$

Amphoteric Species

- Amphoteric species are substances that can act as an acid and a base
  - Most commonly metals and hydroxides

- As an acid:
  - $Al(OH)_3(s) + 3HCl(aq) \rightarrow AlCl_2(aq) + 3H_2O(l)$
- As a base:
  - $Al(OH)_3(s) + NaOH(aq) \rightarrow Na[Al(OH)_4](aq)$

Lesson 2

Properties of Acids and Bases

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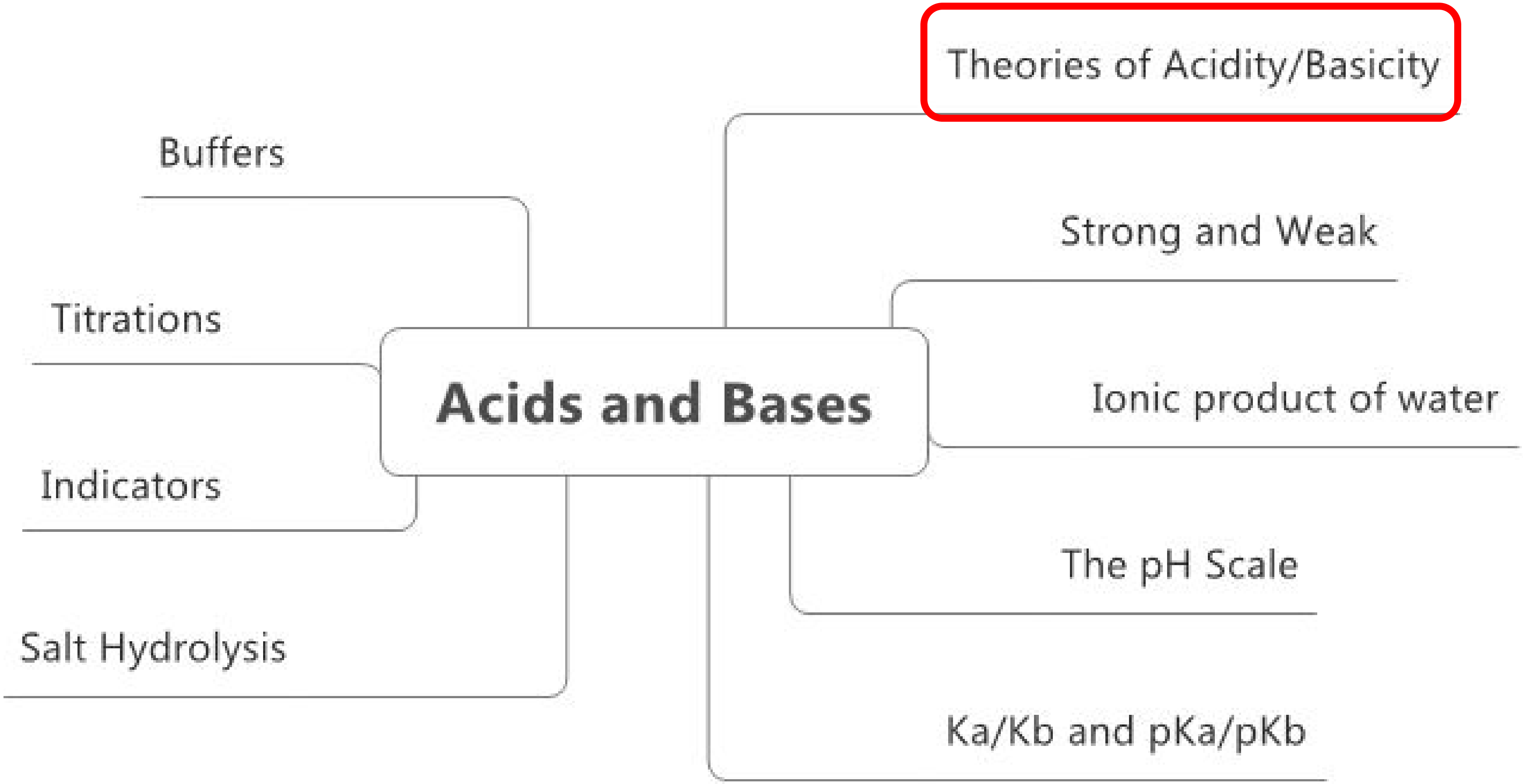


Main

Properties of Acids and Bases

Acids	Bases
Taste sour	Taste bitter
pH < 7.0	pH > 7.0
Litmus is red	Litmus is blue
Phenolphthalein is colorless	Phenolphthalein is pink
Methyl orange is red	Methyl orange is yellow

Common Acids and Bases

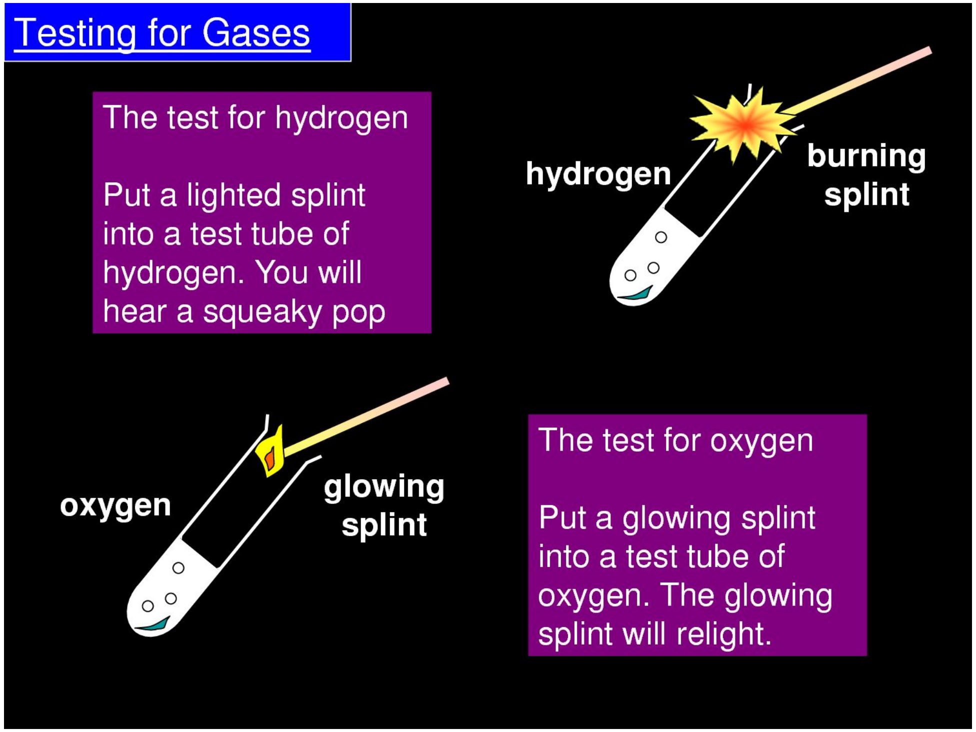
Caustic Acid (Sodium Hydroxide)	Base	Dissolves grease and oil deposits that can block domestic and commercial drains
Phosphoric Acid	Acid	Rust remover
Ammonia	Base	General cleaner
Vinegar (Acetic Acid)	Acid	Treating wasp stings

Acids with Metals

- ACID + METAL → SALT + HYDROGEN
  - $2HCl(aq) + Zn(s) \rightarrow ZnCl_2(aq) + H_2(g)$
  - $H_2SO_4(aq) + Fe(s) \rightarrow FeSO_4(aq) + H_2(g)$
  - $2CH_3COOH(l) + 2Na(s) \rightleftarrows 2CH_3COONa(l) + H_2(g)$
- The rates at which hydrogen gas are produced depend on the reactivity of the metal and the strength and concentration of the acid
- The salt produced depends on the acid from which it was produced

Testing for Hydrogen

- Putting a lighted splint in hydrogen gas results in a combustion reaction, and a "squeaky pop" is heard
- Hydrogen gas is highly flammable



Neutralization

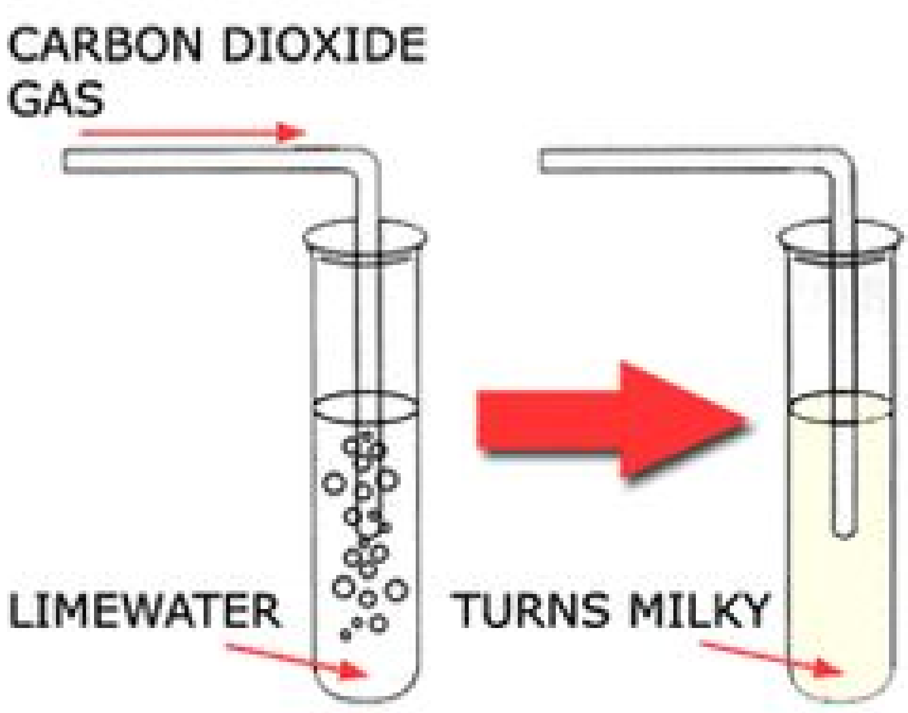
- ACID + ALKALI → SALT + WATER
  - $2HCl(aq) + Ca(OH)_2(aq) \rightarrow CaCl_2(aq) + 2H_2O(l)$
  - $H_2SO_4(aq) + CaO(s) \rightarrow CaSO_4(s) + H_2O(l)$
  - $CH_3COOH(aq) + NH_4OH(aq) \rightleftarrows CH_3COONH_4(aq) + H_2O(l)$
- This is always an exothermic reaction resulting in a $-\Delta H$
- Calcium oxide does not directly react with aqueous acids and must be dissolved to create alkaline solutions
- A soluble base is called an alkali
- Some bases are insoluble

Acids with Metal Carbonates or Hydrogencarbonates

- ACID + METAL CARBONATE/ HYDROGEN CARBONATE → SALT + CARBON DIOXIDE + WATER
  - $2HCl(aq) + Na_2CO_3(s) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$
  - $HCl(aq) + NaHCO_3(s) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$

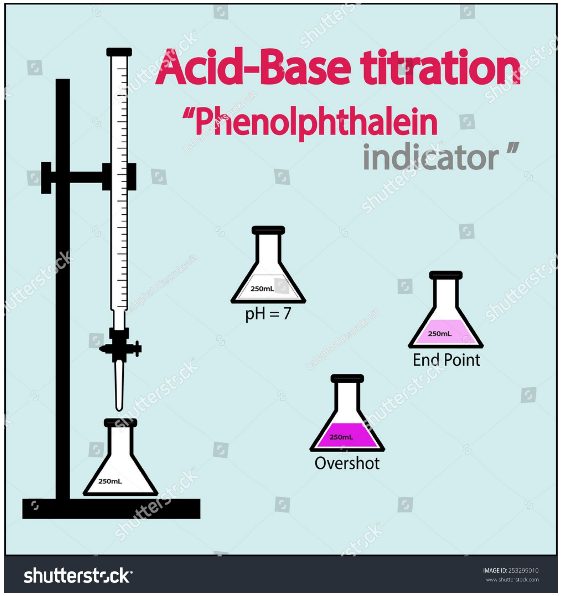
Testing for Carbon Dioxide

- Passing carbon dioxide gas through limewater (calcium hydroxide) results in a cloudy suspension of insoluble calcium carbonate



Acid-Base Titrations

- Titration is a volumetric analysis technique that involves a reaction between a substance of unknown concentration with a standard solution (titrant)
- The titrant is delivered from a burette into the solution being analyzed in small increments
- An acid-base indicator undergoes a color change as the titration approaches and reaches equivalence point



Main

Lesson 3

The pH Scale

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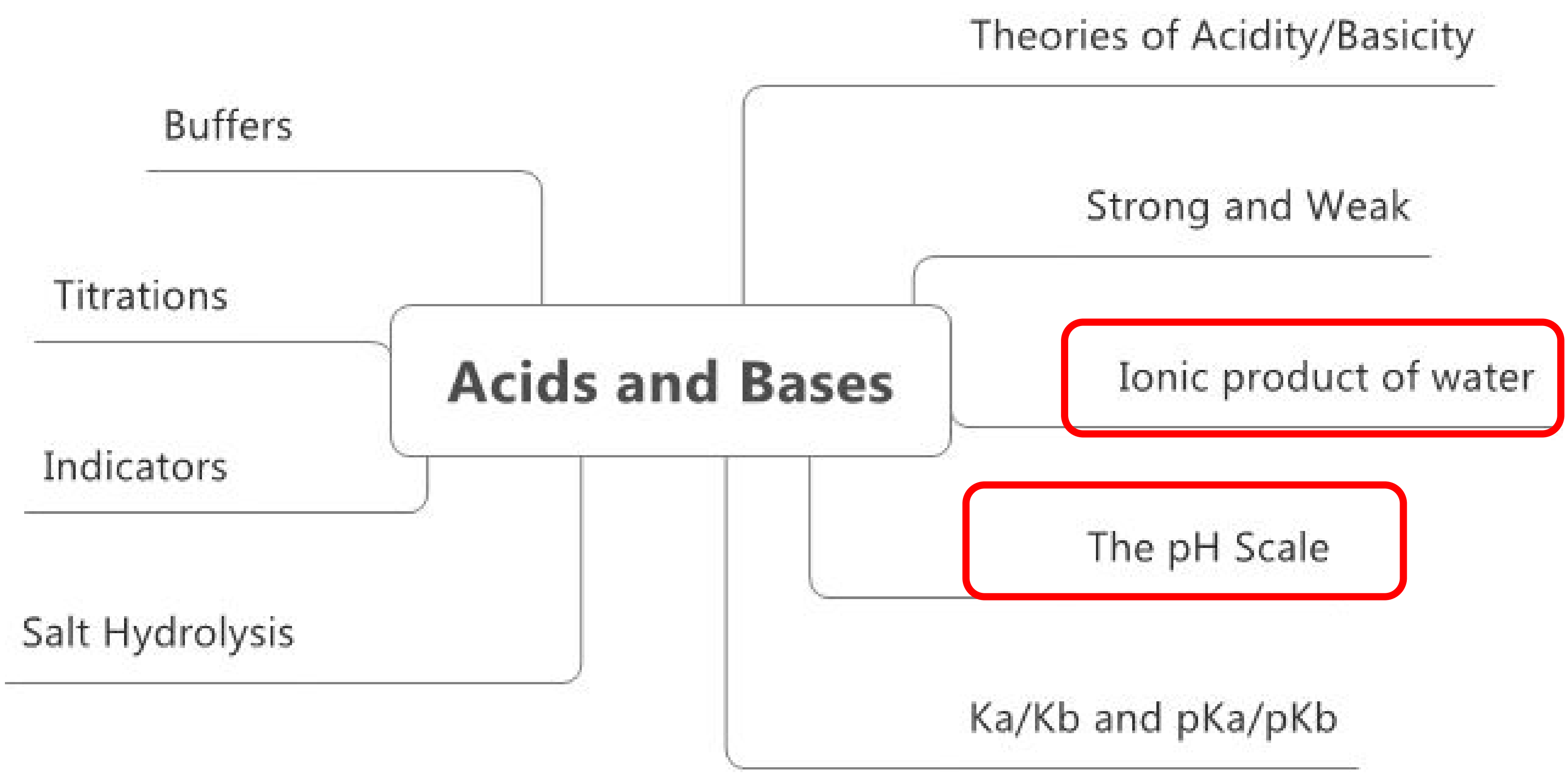
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The pH Scale
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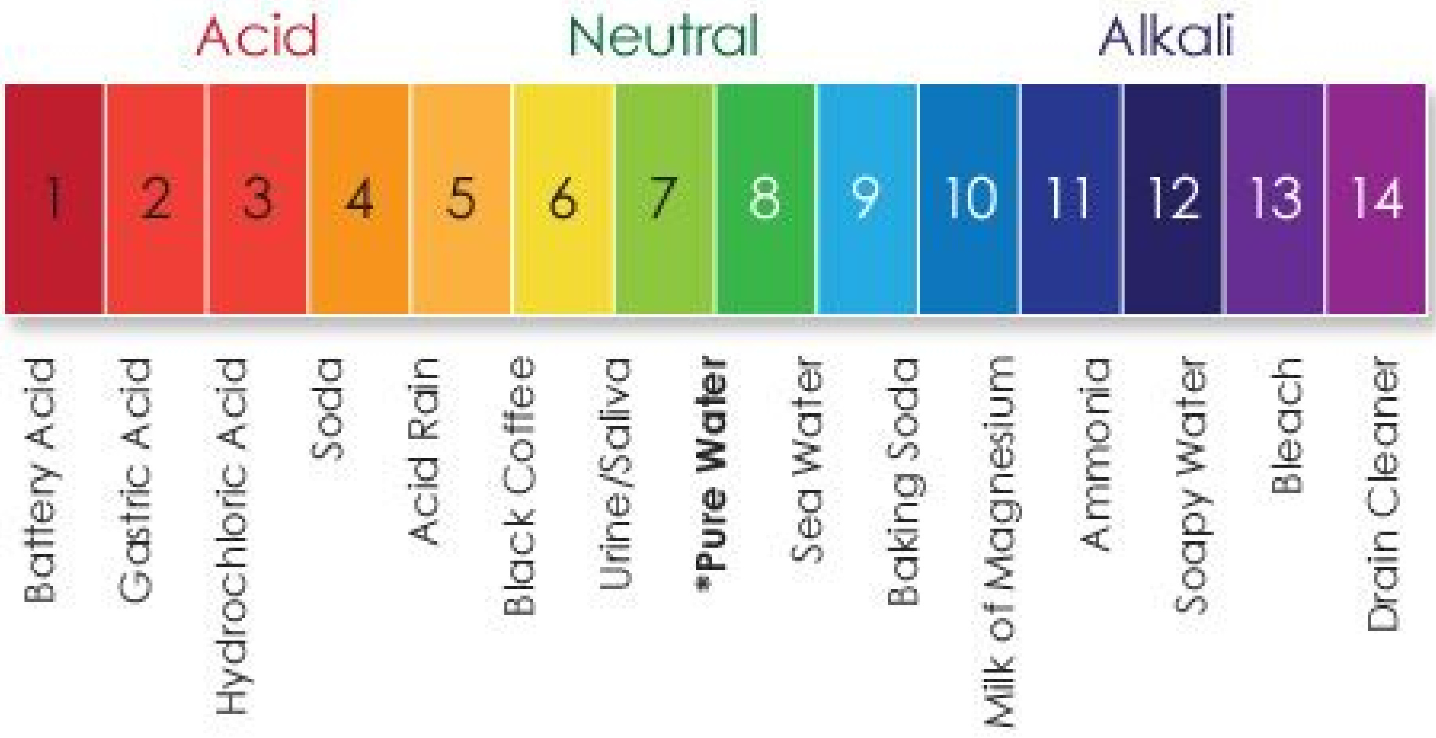
The pH Scale

- The pH scale represents the concentration of hydrogen ions $[H^+]$ in a solution



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Calculating pH

- $pH = -log[H^+(aq)]$ OR $pH = -log [H_3O^+(aq)]$
- $[H^+] = 10^{-pH}$
- The concentration of strong monoprotic acids will be the same as the concentration of the hydrogen ion
- A 0.1 mol dm^{-3} solution of HCl (aq) equates to $[H^+]$ = 0.1 mol dm^{-3}

pH	$[H^+]$
1	1×10^{-1}
2	1×10^{-2}
5	1×10^{-5}
7	1×10^{-7}
10	1×10^{-10}
14	1×10^{-14}

Ionization of Water

- Water undergoes auto-ionization
- $H_2O \rightleftarrows H^+ + OH^-$
- At 298K

$$K_w = \frac{[H^+][OH^-]}{[H_2O]}$$

- $[H_2O]$ is co $K_w = 1 x 10^{-14}$

- For pure water: $K_W = [H^+][OH^-] = 1 \times 10^{-14}$
  - $[H^+]=[OH^-] = \sqrt{1.0x10^{-14}} = 1.0x10^{-7}$

Solutions are..

- Neutral if $[H^+ (aq)] = [OH^- (aq)]$
- Acidic if $[H^+ (aq)] > [OH^- (aq)]$
- Basic if $[H^+ (aq)] < [OH^- (aq)]$

Lesson 4

Strong and Weak Acids and Bases

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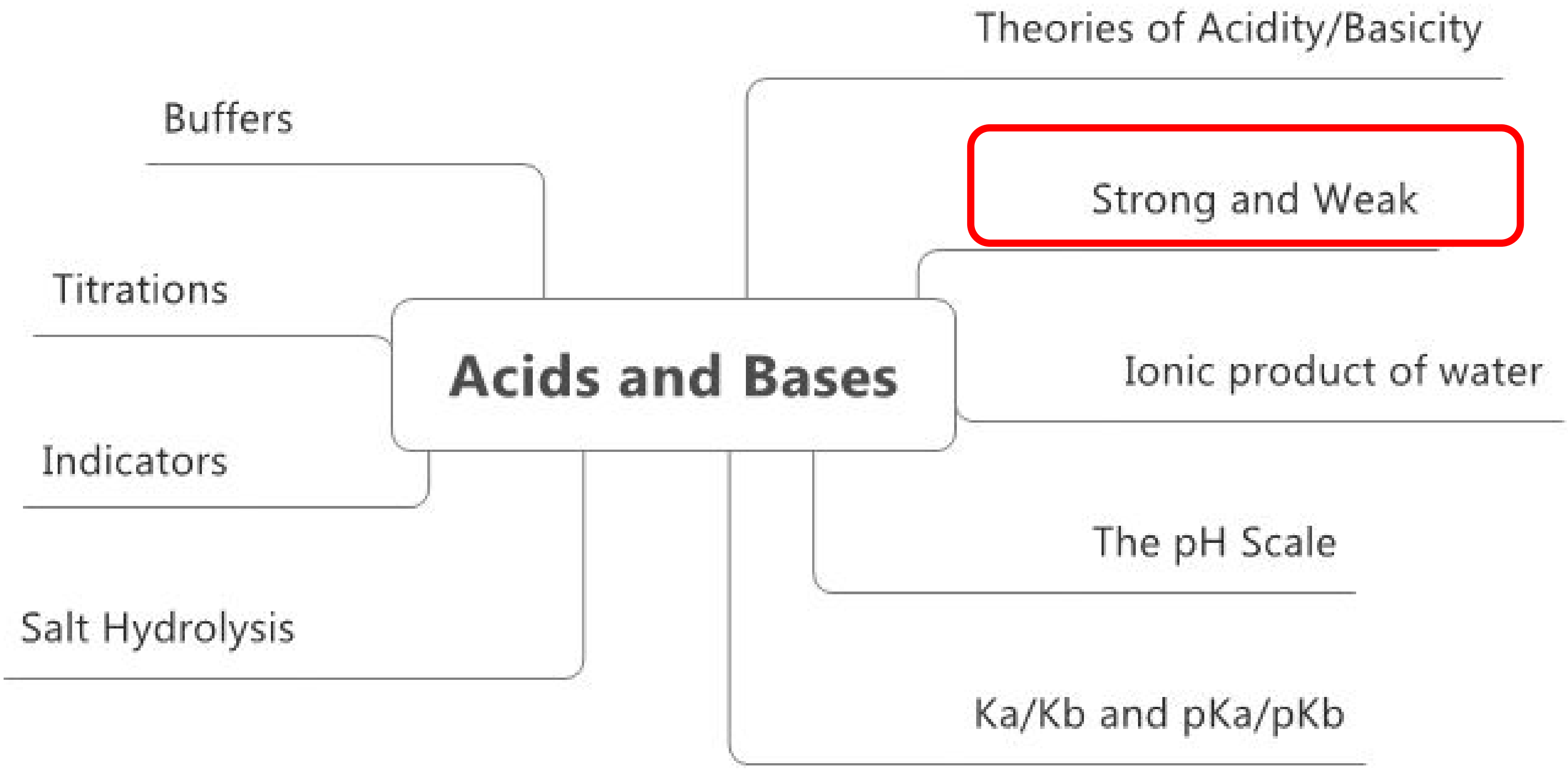
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Strengths of Acids and Bases

- The strength of an acid or base depends on the degree to which it ionizes or dissociates in water
- A strong acid is an effective proton donor that is assumed to completely dissociate in water
- Strong acids have weak conjugate bases and vice versa

Strengths of Acids and Bases

- A weak acid dissociates only partially in water; it is a poor proton donor
- The dissociation of a weak acid is a reversible reaction that reaches equilibrium
- At equilibrium only a small portion of the acid molecules have dissociated

Strengths of Acids and Bases

strong acid:

$$HCl + H_2O \longrightarrow H^+ + Cl^-$$

weak acid:

$$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H^+$$

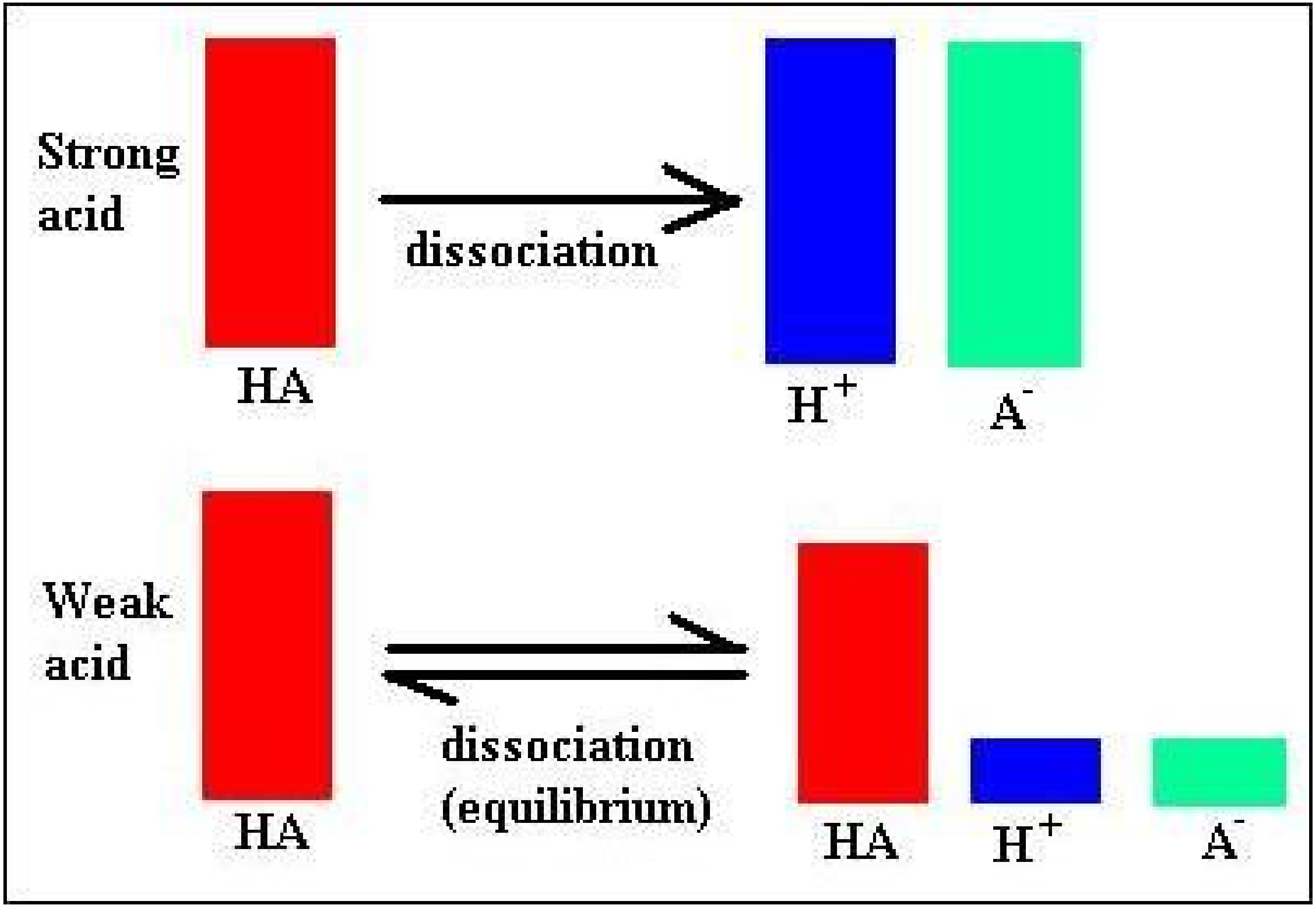
strong base:

$$NaOH + H_2O \longrightarrow Na^+ + OH^-$$

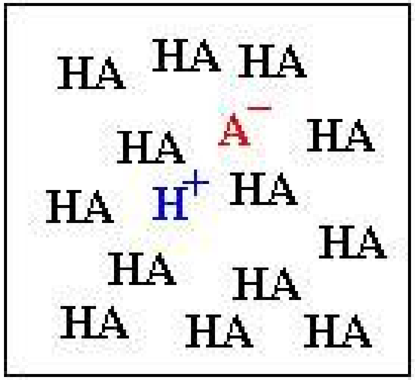
weak base:

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

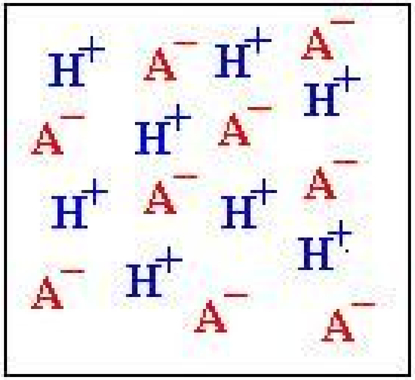
Strengths of Acids and Bases



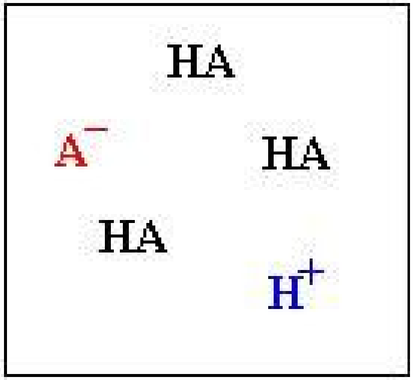
Concentration vs Dilution



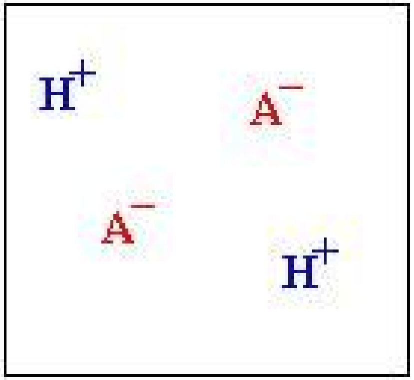
Concentrated weak acid - a lot present, but little dissociation of acid



Concentrated strong acid - a lot present with a lot of dissociation to form many hydrogen ions



Dilute weak acid - little acid present with little dissciation of acid



Dilute strong acid - much acid present with a high degree of dissociation

Strengths of Acids and Bases

- A strong base completely dissociates in water
- Group 1 metal hydroxides are all soluble in water
  - $NaOH(aq) \rightarrow Na^{+}(aq) + OH^{-}(aq)$
  - $KOH(aq) \rightarrow K^{+}(aq) + OH^{-}(aq)$
- A metal hydroxide does not act a a Brønsted-Lowry base because it does not have the capacity to accept a proton
- In solution, hydroxide ions act as a Brønsted-Lowry base
- NH_3 is an example of a weak base
  - $NH_3(aq) + H_2O(l) \rightarrow NH_4^{+}(aq) + OH^{-}(aq)$

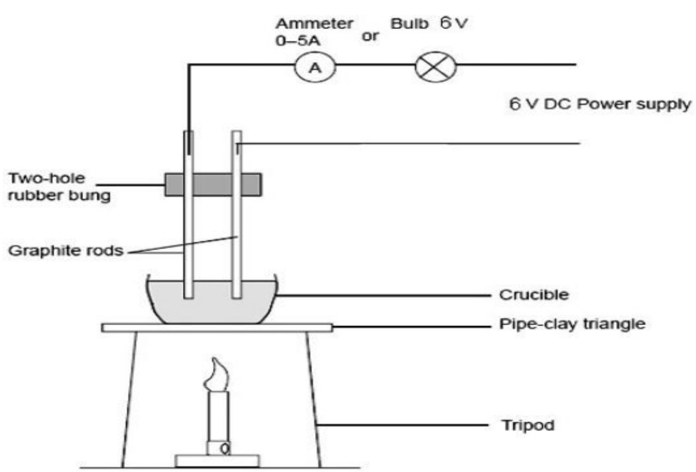
6 Strong Acids		6 Strong Bases	
$HClO_4$	perchloric acid	LiOH	lithium hydroxide
HCl	hydrochloric acid	NaOH	sodium hydroxide
HBr	hydrobromic acid	KOH	potassium hydroxide
HI	hydroiodic acid	$Ca(OH)_2$	calcium hydroxide
HNO_3	nitric acid	$Sr(OH)_2$	strontium hydroxide
H_2SO_4	sulfuric acid	$Ba(OH)_2$	barium hydroxide

Simulation

- http://phet.colorado.edu/en/simulation/acid-base-solutions

Experimental Determination

- Several techniques can be utilized to compare acids and bases of equal concentration
- Conductivity: the conductivity of aqueous solutions depends on the concentration of ions present
  - Strong acids and bases are strong electrolytes that display higher conductivity than weak acids and bases
- Ammeters reads the current from electrolytic solutions



So what?

- The equilibrium has a profound effect on the properties of the acid/base

- Compared with strong acids of the same concentration, weak acids:
  - Have lower electrical conductivity
  - React more slowly
  - pH is higher (less acid)
  - Change pH more slowly when diluted
  - However, they neutralise the same volume of alkali

- Weak bases follow a similar pattern

Neutralization

- The neutralization reaction removes ionized species from the dissociation reaction, driving the reaction to completion
- Recall a strong acid or base is completely dissociated in solution resulting in an exothermic reaction
- The ionization of a weak acid or base is mildly endothermic
- The weaker the acid, the more endothermic the dissociation reaction becomes therefore lowering the enthalpy of change neutralization

Neutralization

Reactants	*Strength*		$\Delta_n H^{\ominus}$
	Acid	*Base*	*(kJ mol^{-1})*
$HCl(aq) + NaOH(aq)$	strong	strong	– 57.1
$HNO_3(aq) + NaOH(aq)$	strong	strong	– 57.1
$CH_3COOH(aq) + NaOH(aq)$	weak	strong	– 56.1
$HCl(aq) + NH_3(aq)$	strong	weak	– 53.4
$CH_3COOH(aq) + NH_3(aq)$	weak	weak	– 50.4

Lesson 5

Acid Deposition

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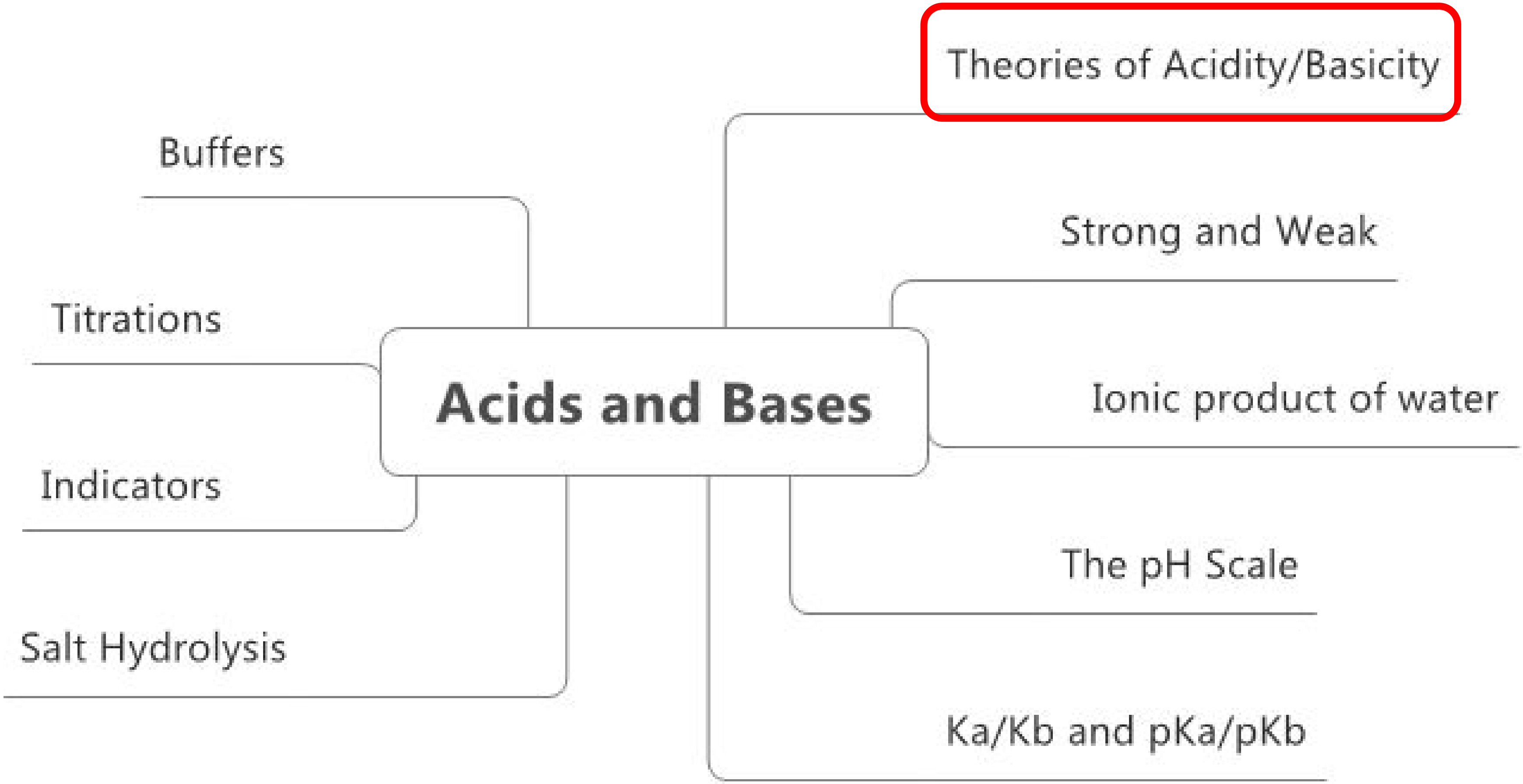
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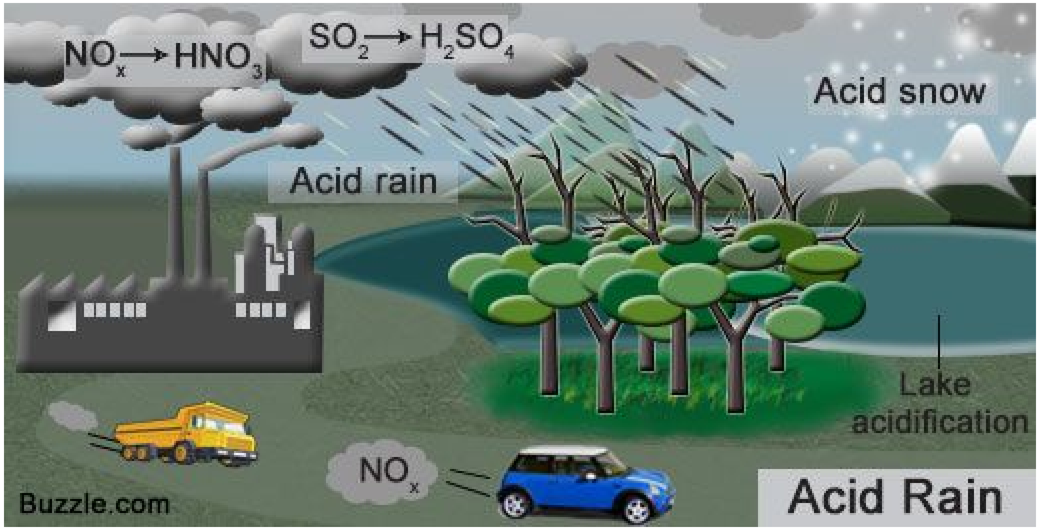
Acid Deposition

- Acid deposition is the process by which acid-forming pollutants are deposited on the Earth's surface
- Increased industrialization and economic development have led to increased emissions of nitrogen and sulfur oxides that cause acid rain



Acid Deposition

- Acid deposition leads to:

1. Deforestation
2. The leaching of minerals from soils leading to elevated acid levels in lakes and rivers
3. The uptake of toxic minerals from soil by plants, reduction in the pH of lake and river systems
4. Increased uptake of toxic metals by shellfish and other marine life which can affect the fishing industry and people's health
5. Corrosive effects on marble, limestone, and metal buildings, bridges, and vehicles

Problems associated with acid deposition

- **Effect on vegetation**-the acid can displace metal ions from the soil. Mg^{2+} ions are needed to produce chlorophyll. Acid rain causes aluminum ions to dissolve from rocks, which damages plant roots and limits water uptake, stunting growth

- **Lakes and rivers**-insect larvae, fish, and invertebrates, and others cannot survive below pH 5.2. Acid rain can dissolve minerals from rocks , which can accumulate in lakes

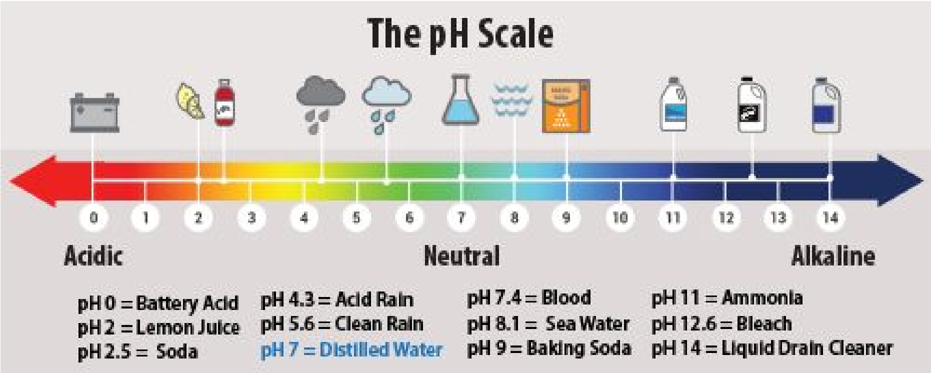
Problems associated with acid deposition

- **Buildings**-limestone and marble are eroded by acid rain and dissolve away

- **Human health**-acids cause respiratory illnesses such as asthma and bronchitis. Acidic water can dissolve heavy metal compounds and release poisonous ions such as Cu^{2+}, Pb^{2+}, and Al^{3+} which may be linked to Alzheimer's disease

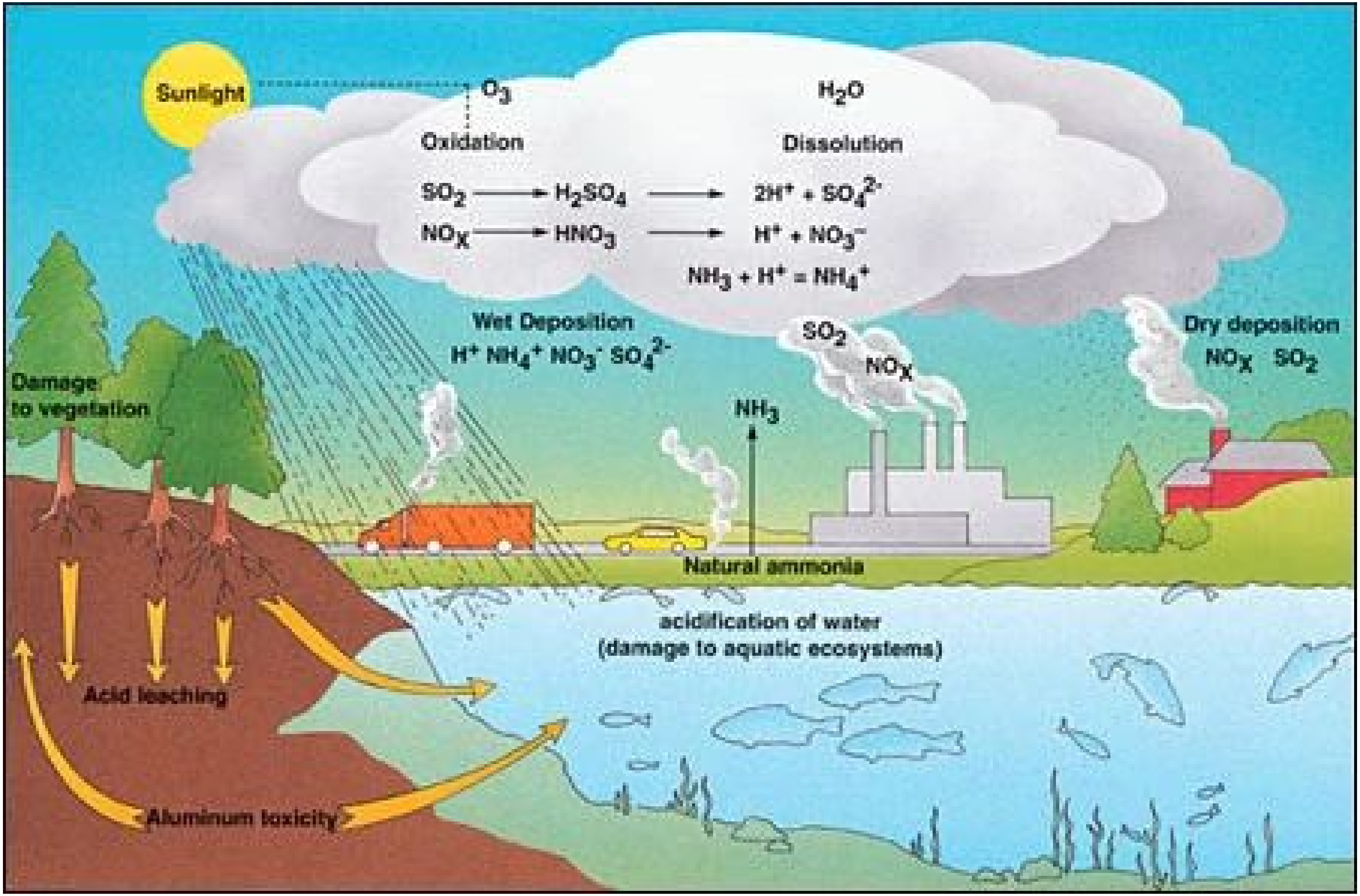
Acid Rain

- Rain is naturally acidic due to presence of dissolved CO_2 which forms weak carbonic acid, H_2CO_3
- Rain has a typical pH of 5.6.
  - $CO_2(g) + H_2O(l) \rightleftarrows H_2CO_3(aq)$
  - $H_2CO_3(aq) \rightleftarrows H^+(aq) + HCO_3^-(aq)$
  - $HCO_3^-(aq) \rightleftarrows H^+(aq) + CO_3^{2-}(aq)$
- Acid rain is anything with a pH lower than 5.6.



Causes of Acid Deposition

- Major pollutants that cause acid deposition are sulfur dioxide, SO_2 and nitrogen oxides, NO and NO_2
- These are products of natural occurrences such as volcanic eruptions, decomposition of vegetation, man made pollutants from combustion of fossil fuels
- Acid rain results from the formation of two strong acids, nitric acid, $HNO_3(aq)$, nitrous acid, $HNO_2(aq)$, sulfuric acid, $H_2SO_4(aq)$, and sulfurous acid $H_2SO_3(aq)$
- Is considered a major global environmental problem



Sunlight
O3
Oxidation
H2O
Dissolution
SO2 → H2SO4 → 2H+ + SO4 2-
NOx → HNO3 → H+ + NO3-
NH3 + H+ = NH4+
Wet Deposition
H+ NH4+ NO3- SO4 2-
SO2
NOx
Dry deposition
NOx SO2
Damage to vegetation
NH3
Natural ammonia
acidification of water
(damage to aquatic ecosystems)
Acid leaching
Aluminum toxicity

Ex: Internal Combustion Engine of a Car or Jet

1. $N_2(g) + O_2(g) \rightarrow 2NO(g)$
   a. $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$ — NO_2 causes brown smog
   b. $2NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + HNO_2(aq)$

2. $N_2(g) + O_2(g) \rightarrow 2NO(g)$
   a. $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$
   b. $2HNO_2(aq) + O_2(g) \rightarrow 2HNO_3(g)$

All oxides of nitrogen eventually produce nitric acid

1. $SO_2(g) + H_2O(l) \rightleftarrows H_2SO_3(aq)$
   a. $H_2SO_3(aq) + H_2O(l) \rightleftarrows HSO_3^-(aq) + H_3O^+(aq)$

2. $S(s) + O_2(g) \rightleftarrows SO_2(g)$ — Coal contains 3% sulfur
   a. $2SO_2(g) + O_2(g) \rightleftarrows 2SO_3(g)$
   b. $SO_3(g) + H_2O(l) \rightleftarrows H_2SO_4(aq)$

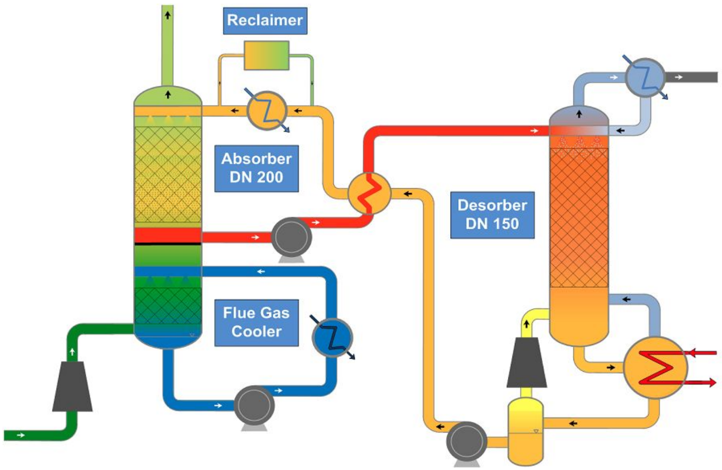
Pre-Combustion Methods

- Techniques to reduce sulfur emissions on fuels before their combustion
- Physical cleaning or mineral beneficiation involves crushing coal followed by flotation that reduces the amounts of sulfur and other impurities
- Combinations of pre-combustion methods result in the removal of up to 80-%-90% of inorganic sulfur

Main

Post-Combustion Methods

- Focus on complementary technologies to remove sulfur dioxide, nitrogen oxides, heavy metals and dioxins from the combustion gases before they are released into the atmosphere
- Ex: Calcium oxide or lime will react with sulfur dioxide and remove it from gases that are released into the atmosphere via flue pipes
  - $CaO(s) + SO_2(g) \rightarrow CaSO_3(s)$



Lesson 6
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Lewis Acids and Bases

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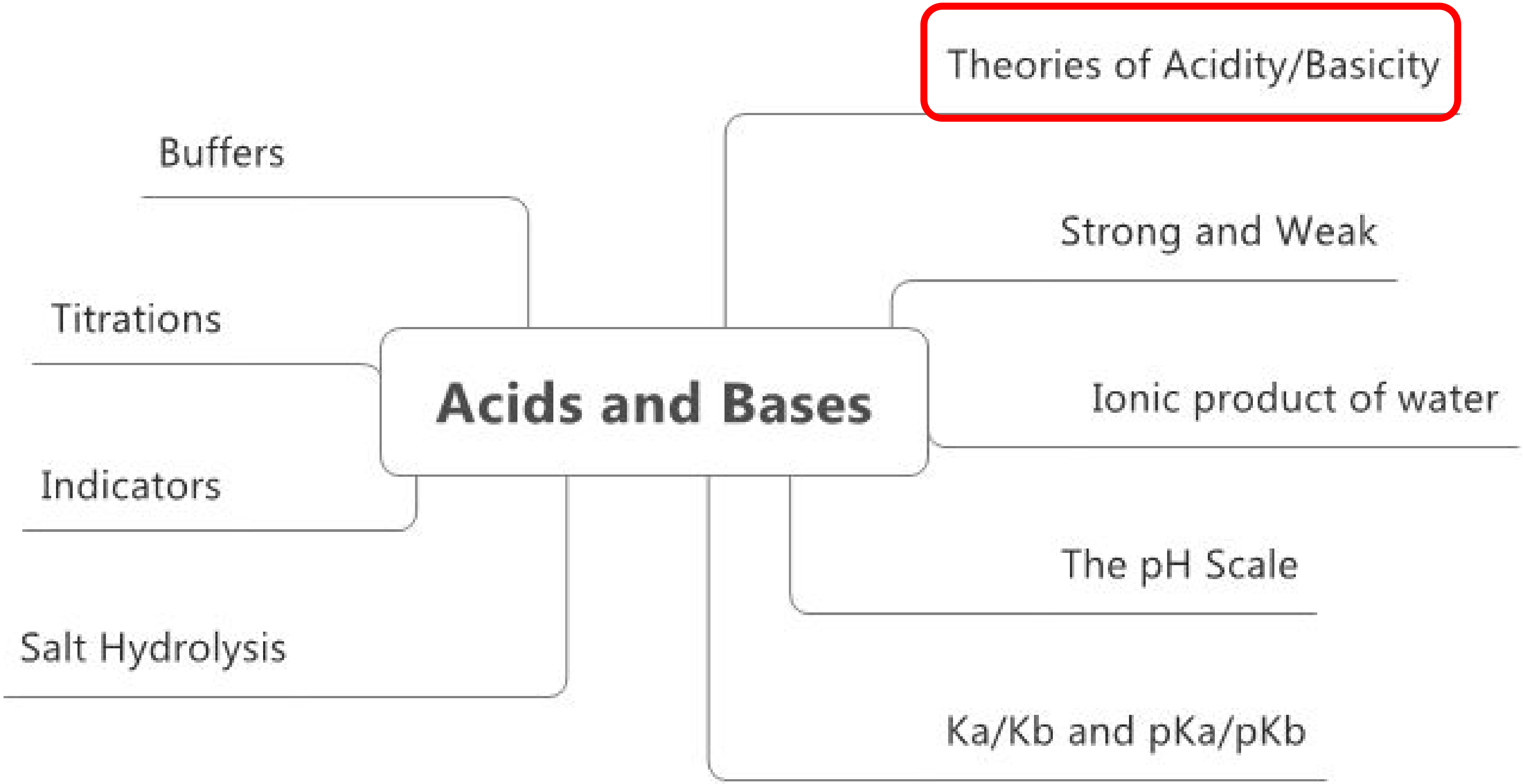
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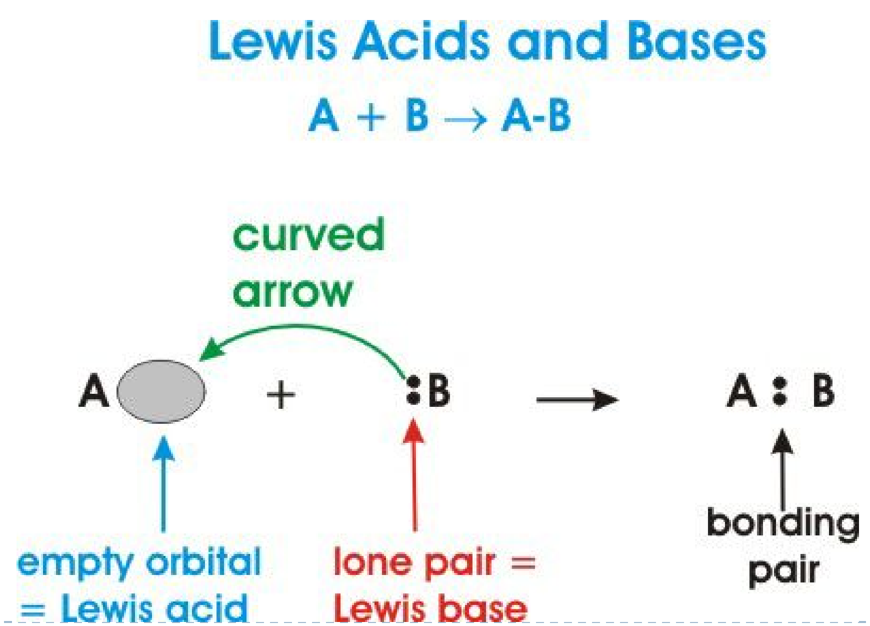
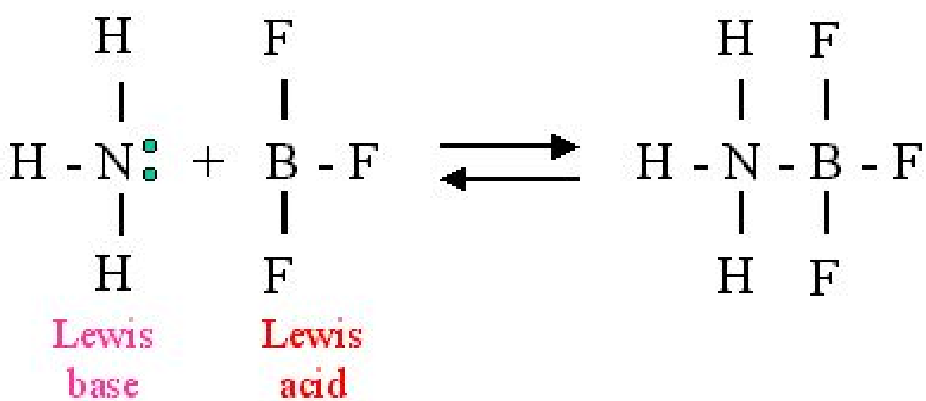
Lewis Acids and Bases

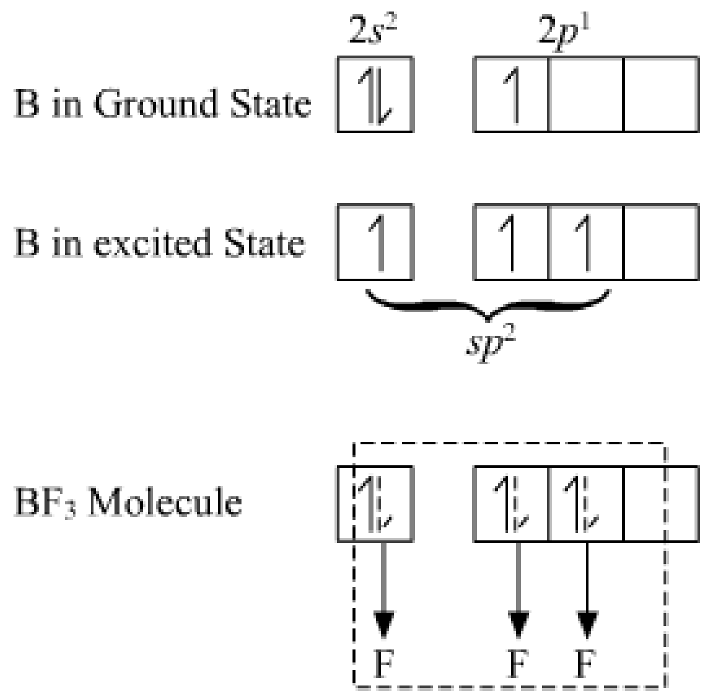
- Lewis acids are defined as electron pair acceptors (electrophile)
- Lewis bases are electron pair donors (nucleophile)
- This general definition of acids and bases allowed a wider range of substances to be included



Coordinate Bonds

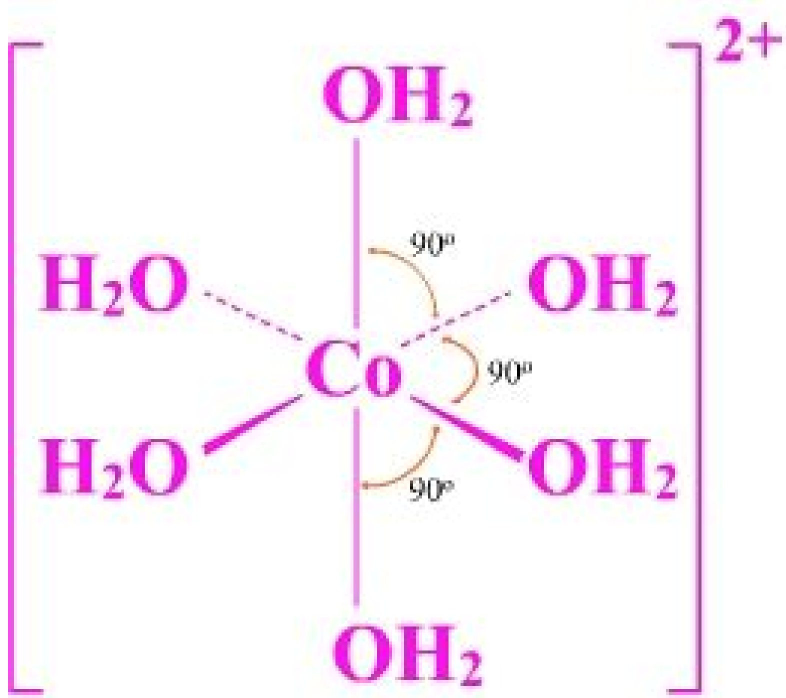
- In coordinate bonds, a lone pair of electrons is donated forming a covalent bond





Coordinate Bonds

- Transition elements have a partially occupied d subshell so they can form complex ions with ligands that possess a lone pair of electrons



Lesson 7
HL

Calculations Involving Acids and Bases

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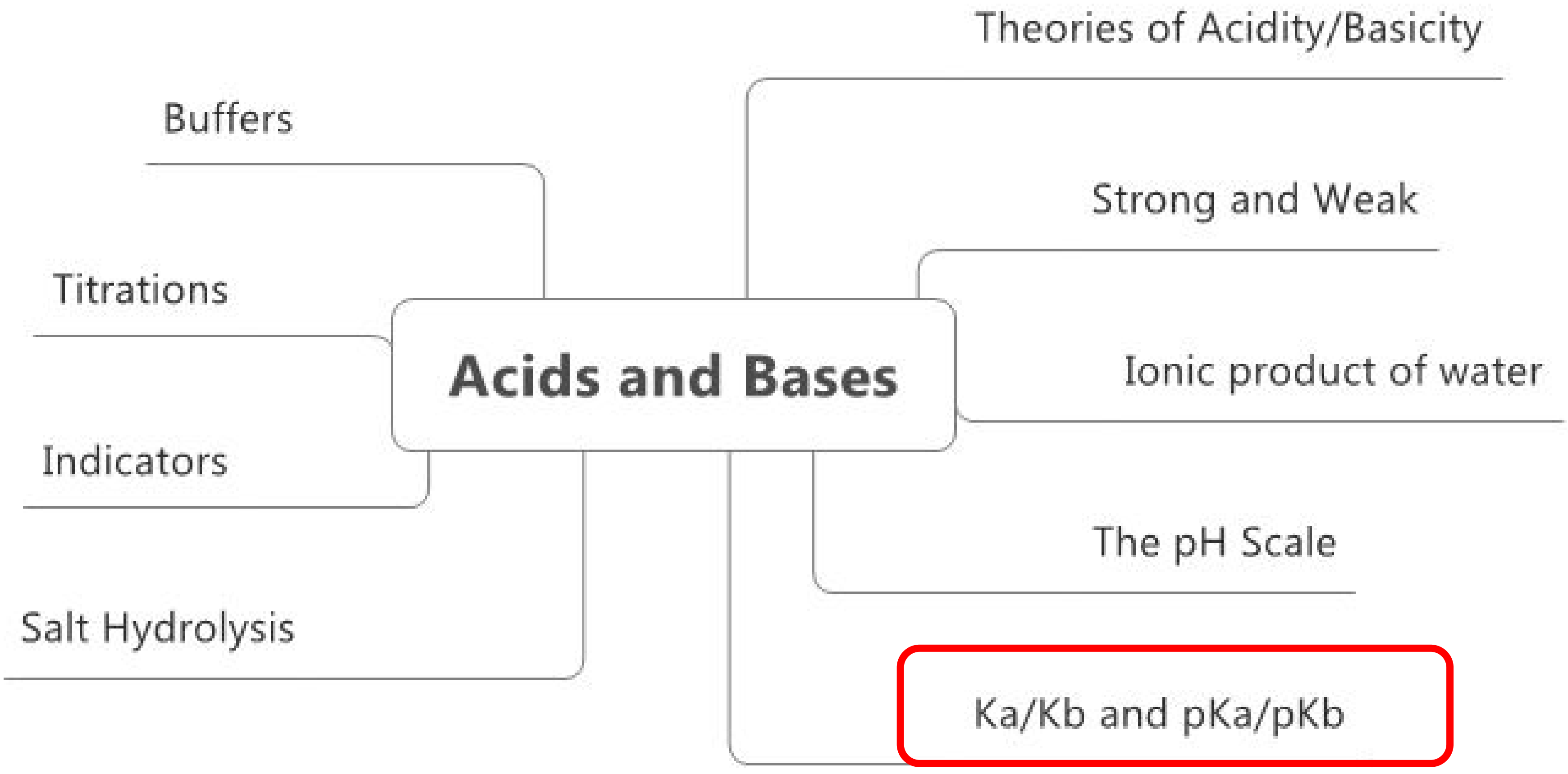

Theories of Acidity/Basicity
Buffers
Strong and Weak
Titrations
Acids and Bases
Ionic product of water
Indicators
The pH Scale
Salt Hydrolysis
Ka/Kb and pKa/pKb

Acid Dissociation Constant

- We can determine the concentration of a dissociated weak acid using the relationship between concentrations of reactants and products and the equilibrium position

$$HA + H_2O \underset{K_a}{\overset{K_d}{\rightleftharpoons}} H^+ + A^-$$

- H_2O is consid

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

Base Dissociation Constant

$$B_{(aq)} + H_2O_{(l)} \rightleftharpoons BH^+_{(aq)} + OH^-_{(aq)}$$

Weak Base

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

Calculating K_a and K_b

- Acids:
  - The stronger the acid the larger the K_a
  - The stronger the acid the weaker the conjugate base
  - The stronger the acid the smaller the K_b of the conjugate base
- Bases:
  - The stronger the base the larger the K_b
  - The stronger the base the weaker the conjugate acid
  - The stronger the base the smaller the K_a of the conjugate acid
- $K_a K_b = [H^+][OH^-] = K_w$
- $K_a K_b = K_w$

Main

Temperature Dependence and K_a

- The ionization of water is an endothermic process
- Per Le Châtlier's principle, a change in temperature will result in a change in the position of equilibrium
- A rise in temperature will result in the forward reaction being favored, increasing the concentration of the hydrogen and hydroxide ions. This represents an increase in K_w and a decrease in pH

Temperature	Kw	pH
15	0.453×10^{-14}	7.17
20	0.684×10^{-14}	7.08
25	1.00×10^{-14}	7.00
30	1.47×10^{-14}	6
35	2.09×10^{-14}	6.84

Main

pK_a and pK_b

- The values for K_a are often very small, so we use 'pK_a' to make them easier to handle:

  $pK_a = -\log_{10}(K_a)$ $K_a = 10^{-pKa}$

- The same goes for K_b

  $pK_b = -\log_{10}(K_b)$ $K_b = 10^{-pKb}$

K_a and pK_a in action

- In order of decreasing acid strength:

Acid	K_a	pK_a
Hydronium ion, H_3O^+	1.00	0.00
Oxalic acid, HO_2CCO_2H	5.9×10^{-2}	1.23
Hydrofluoric, HF	7.2×10^{-4}	3.14
Methanoic, CHOOH	1.77×10^{-4}	3.75
Ethanoic, CH_3COOH	1.76×10^{-5}	4.75
Phenol, C_6H_5OH	1.6×10^{-10}	9.80

- Smaller K_a → weaker acid
- Smaller pK_a → stronger acid

K_b and pK_b in action

- In order of decreasing base strength:

Base	K_b	pK_b
Diethylamine	$1.3x10^{-3}$	2.89
Ethylamine	$5.6x10^{-4}$	3.25
Methylamine	$4.4x10^{-4}$	3.36
Ammonia	$1.8x10^{-5}$	4.74

- Smaller K_b → weaker base
- Smaller pK_b → stronger base

pH to pKa

- Henderson-Hasselbalch Equation

$$\mathrm{HA} \underset{}{\overset{K_a}{\rightleftharpoons}} \mathrm{A^-} + \mathrm{H^+}$$

$$K_a = \frac{[\mathrm{H^+}][\mathrm{A^-}]}{[\mathrm{HA}]}$$

$$\log_{10} K_a = \log_{10}[\mathrm{H^+}] + \log_{10}\frac{[\mathrm{A^-}]}{[\mathrm{HA}]}$$

$$-\mathrm{p}K_a = -\mathrm{pH} + \log_{10}\frac{[\mathrm{A^-}]}{[\mathrm{HA}]}$$

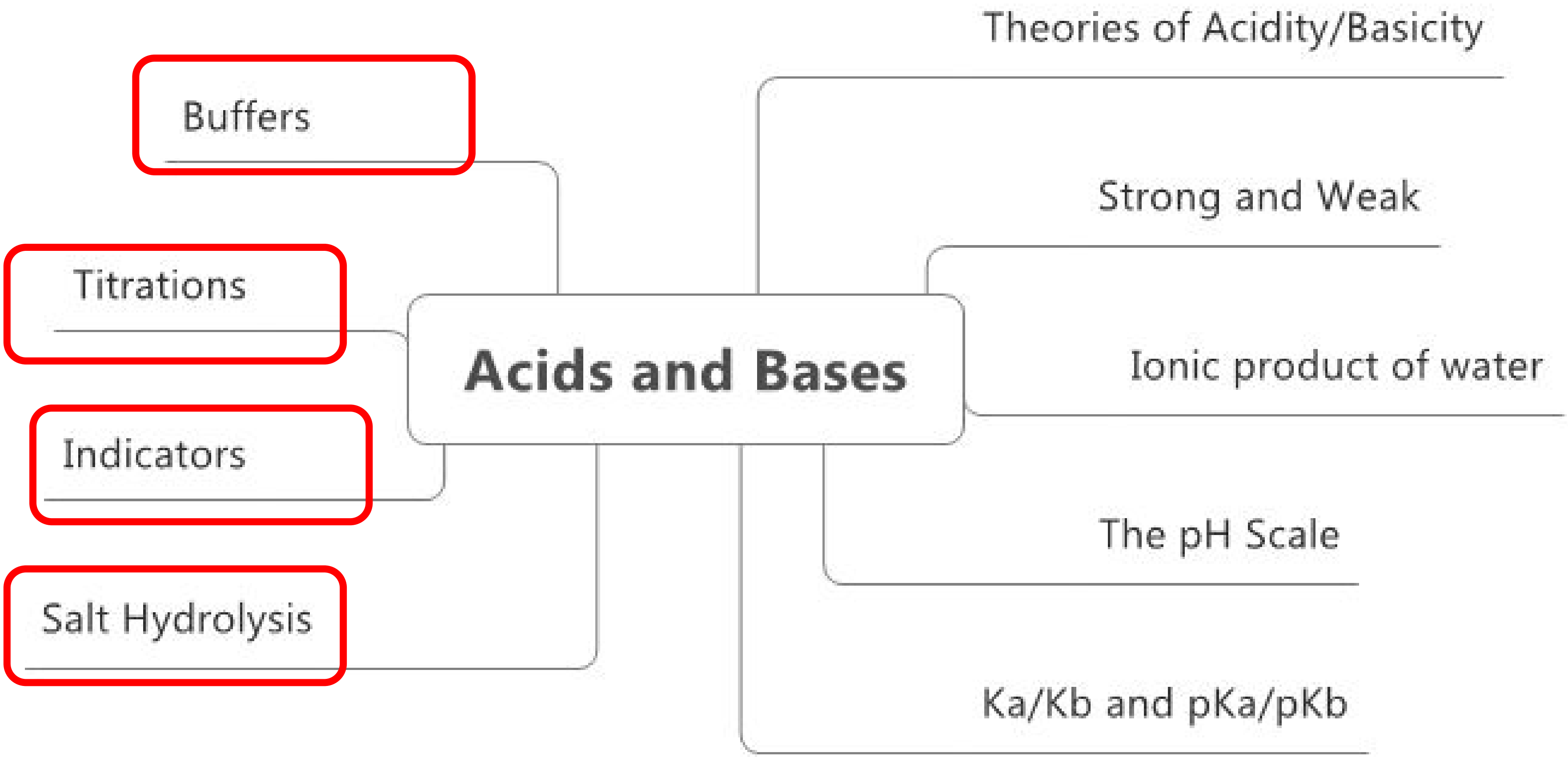
$$\mathrm{pH} = \mathrm{p}K_a + \log_{10}\frac{[\mathrm{A^-}]}{[\mathrm{HA}]} \quad (1)$$

Lesson 8
HL

pH Curves

Main Menu

We Are Here


Theories of Acidity/Basicity
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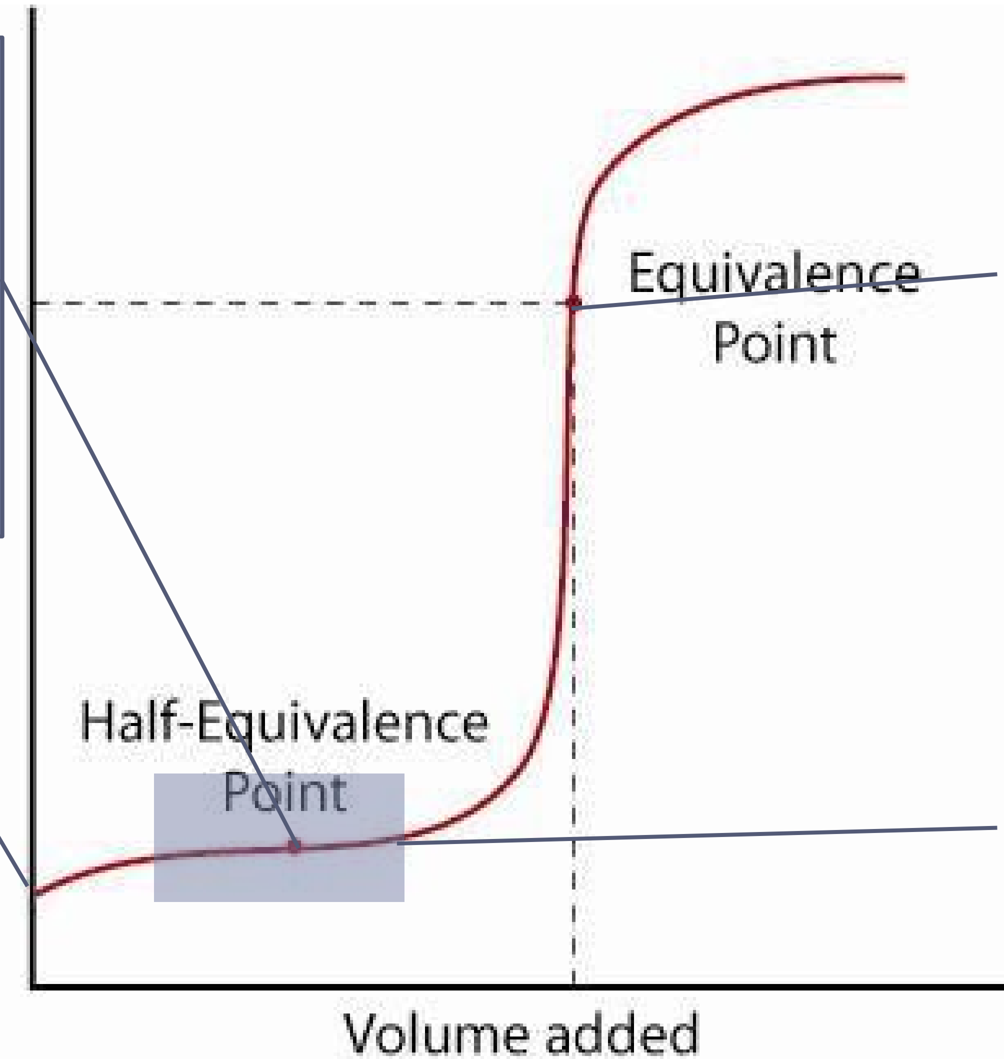
Buffer Solutions

- Buffers are solutions that resist a change in pH upon the addition of small amounts of a strong base or strong acid
- A buffer may be composed of a weak acid and its conjugate base or a weak base and its conjugate acid

Anatomy of an acid-base titration curve

Half-equivalence point:
i.e. Half the volume added at the equivalence point
The pH of this point is equal to the pKa of the weak acid.



Equivalence point:
i.e. The point of inflection (where the gradient starts to drop again).
Represents the point at which the acid has just been neutralised.

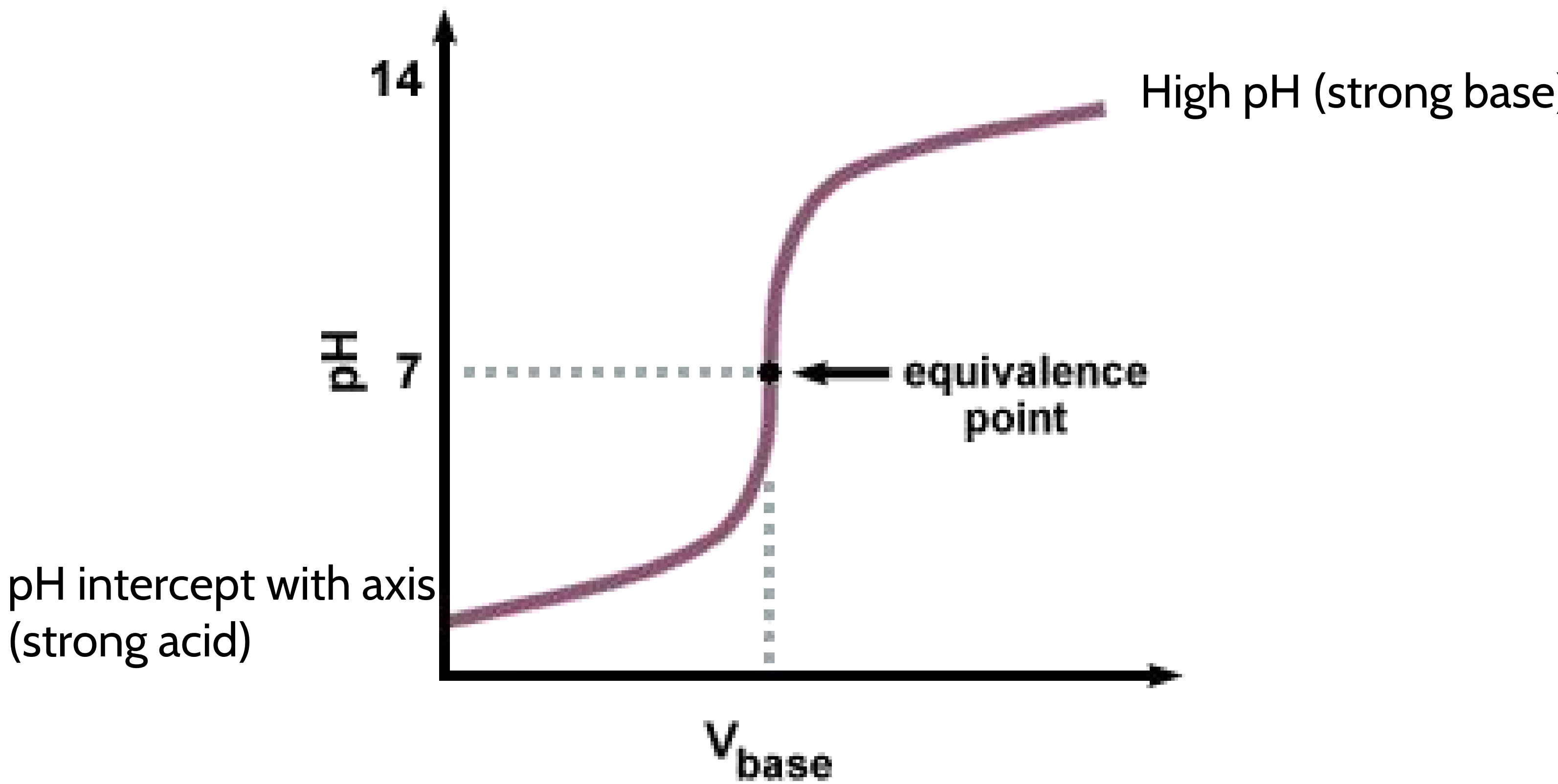
pH Intercept:
Higher pH if a weaker acid used

Buffer Region:
A lot of base has to be added to result in only a small change of pH

Main

Strong Acid with Strong Base Curve

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



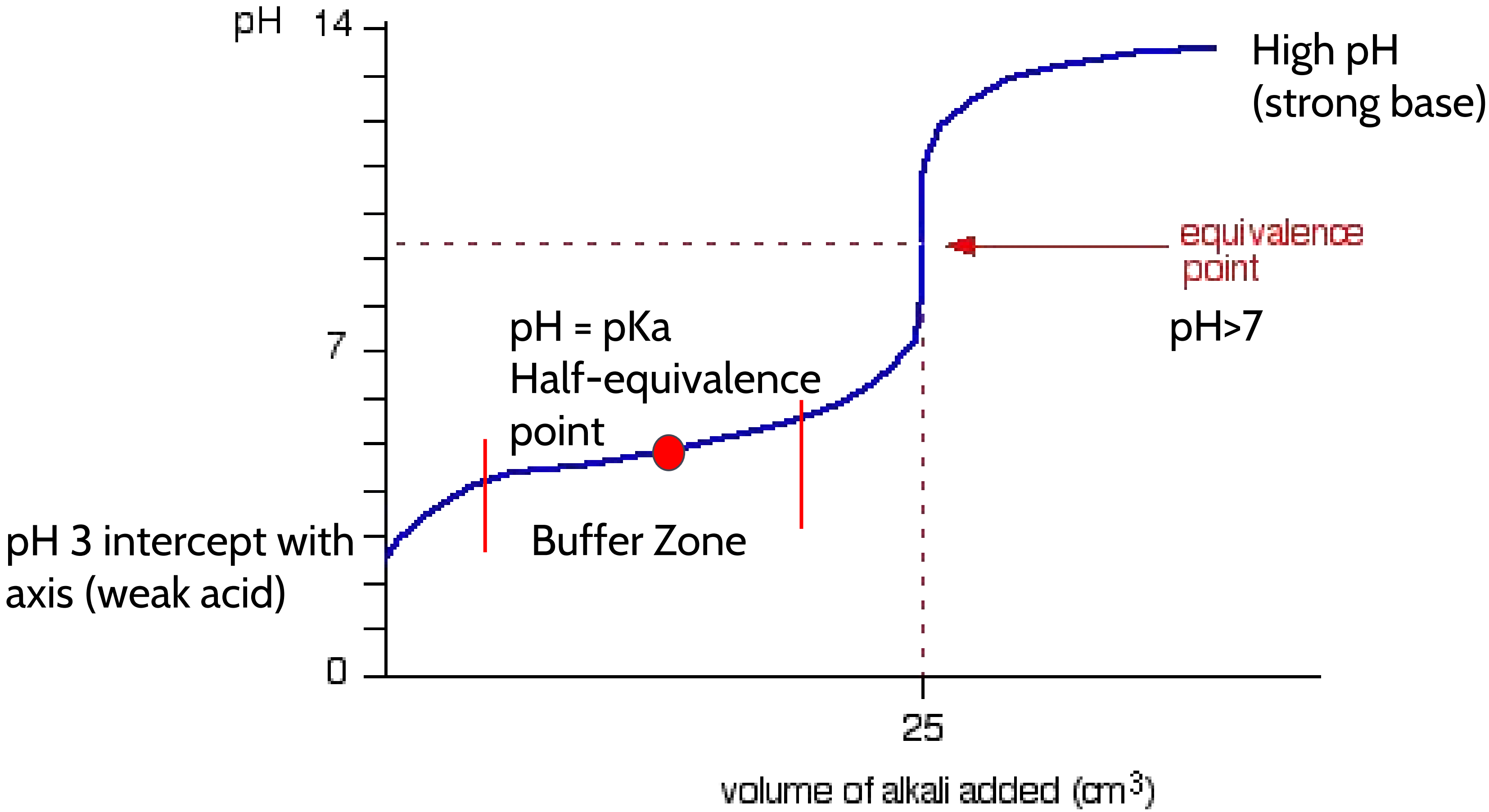
Strong Acid with Strong Base Curve

- Characteristics:
  - The starting point on the pH axis indicates initial pH value
  - There is a gradual rise in the pH as the titration approaches the equivalence point
  - The sharp rise in pH at the equivalence point (pH=7) is described as the point of inflection of the curve
  - Once there is no remaining acid to be neutralized, the curve flattens and finishes at a high pH reflecting the strong base

Main

Weak Acid with Strong Base Curve

$$CH_3COOH(aq) + NaOH(aq) \rightarrow CH_3COONa(aq) + H_2O(l)$$

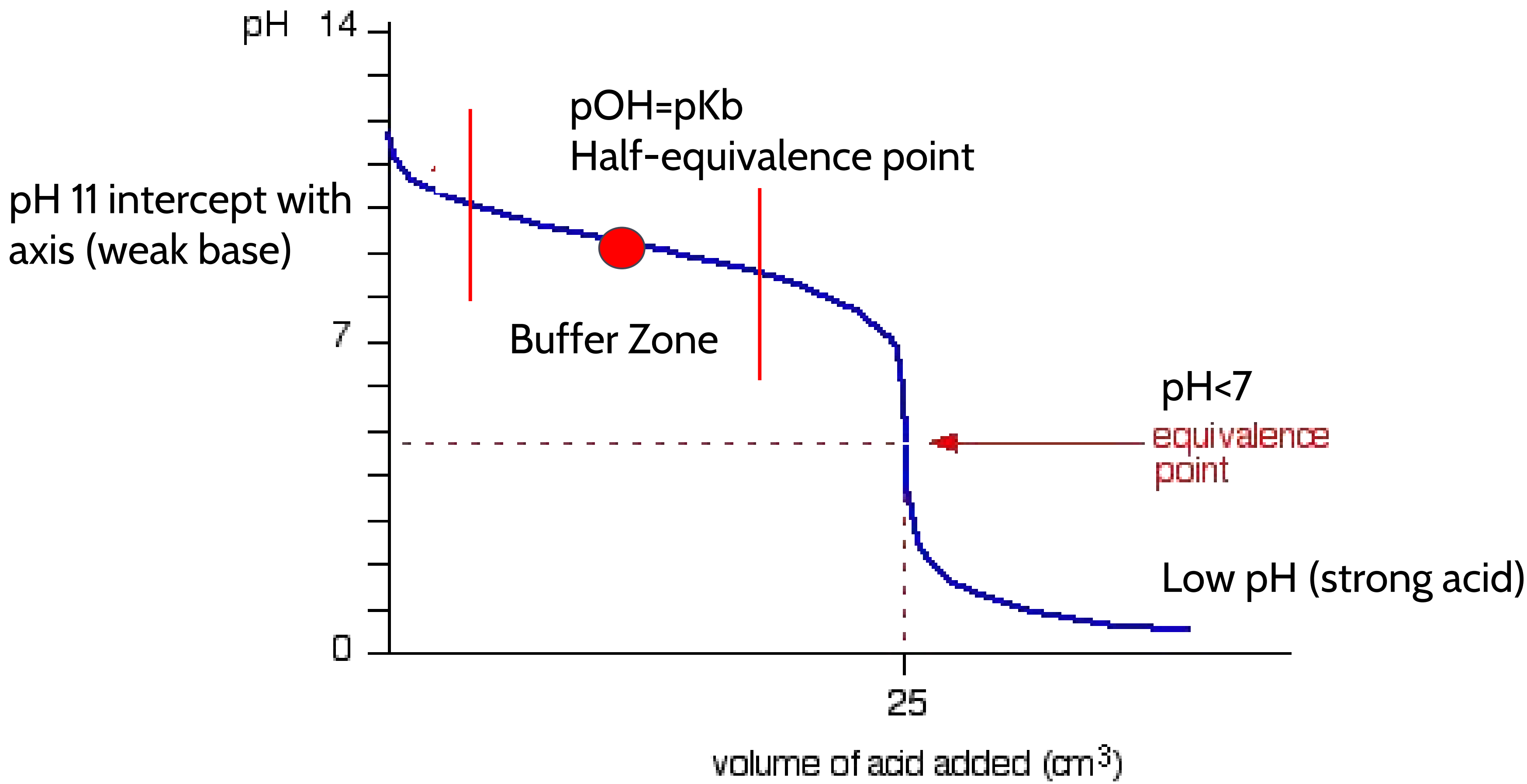


Weak Acid with Strong Base Curve

- Characteristics:
  - The starting point on the pH axis indicates initial pH value
  - The initial rise is steep, as a strong base is being added with a weak acid and neutralization is rapid
  - As the weak acid begins to be neutralized the strong conjugate base is formed creating a buffer that resists change in pH
  - The continued addition of base results in forward reaction being favored. This results in a very gradual change in pH in this region of the curve
  - The half-equivalence is the stage at which half the amount of weak acid has been neutralized. pKa=pH
  - There is a sharp rise in pH at the equivalence point (pH>7). The equivalence point is the result of salt hydrolysis. Hydrolysis is the ionization of water that results from reaction with ionic salts.
  - With no remaining acid to be neutralized, the curve flattens and finishes at a high pH due to the presence of excess base

Weak Base with Strong Acid

$$HCl(aq) + NH_3(aq) \rightarrow NH_4Cl(aq)$$



Main

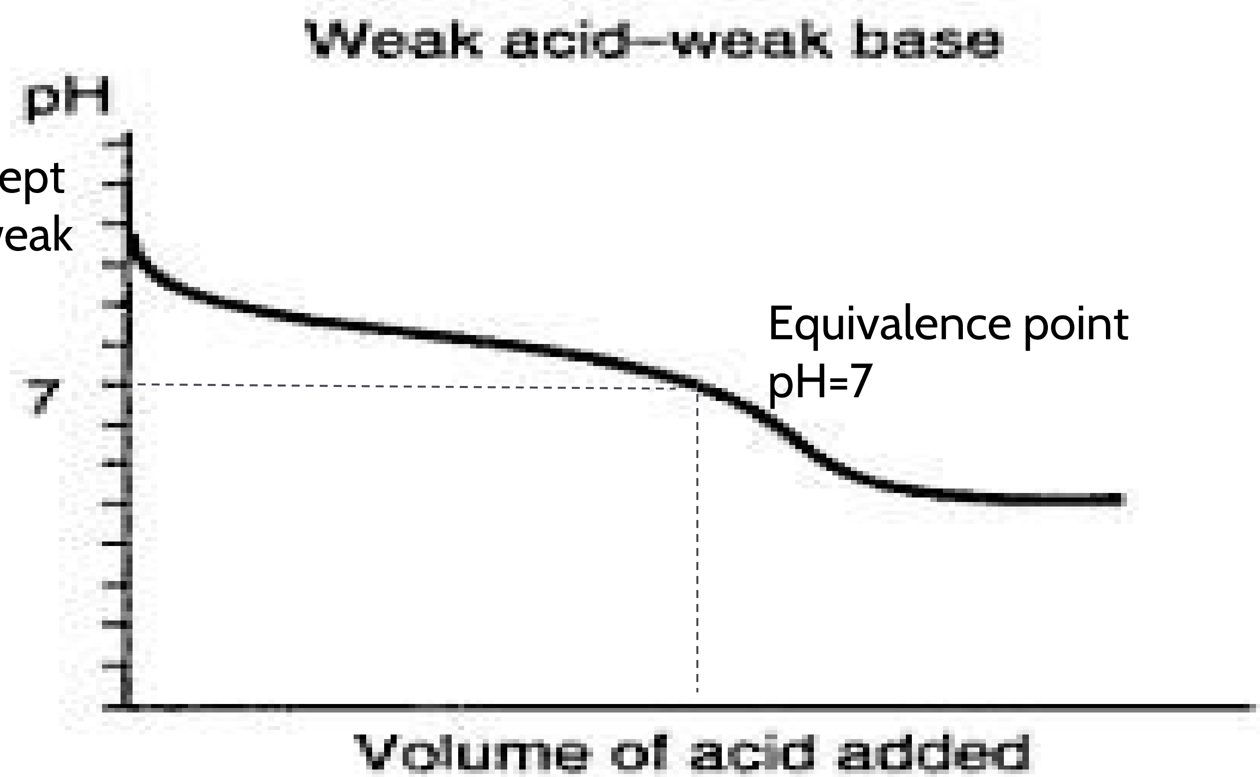
Weak Base with Strong Acid

- Characteristics
  - The starting point on the pH axis indicates initial pH value
  - As the weak base begins to be neutralized, the conjugate acid is created resulting in a buffer that resists change in pH
  - At the half-equivalence point half of the amount of weak base has been neutralized. $pOH = pK_b$
  - There is a gradual fall in the pH due to the buffering effect as the titration approaches equivalence point
  - The pH falls sharply at the equivalence point (pH<7). The equivalence point is the result of salt hydrolysis
  - With no remaining base to be neutralized, the curve flattens and ends at a low pH due to the presence of excess acid

Weak Base with Weak Acid

$$NH_3(aq) + CH_3COOH(aq) \rightarrow CH_3COONH_4(aq)$$

pH 11 intercept with axis (weak base)



Equivalence point pH=7

Main

Weak Acid with Weak Base Curve

- Characteristics:
  - The starting point on the pH axis indicates initial pH value
  - The change in pH throughout the titration is very gradual
  - The point of inflection in the pH curve is not as steep as in the previous curves. The point of equivalence is difficult to determine, so this kind of titration has little or no practical us
  - With no remaining base to be neutralized, the curve flattens and ends at a pH that indicates the presence of a weak acid

Indicators

- An indicator is typically a weak acid or a weak base that displays a different color in acidic or alkaline environments

$$HIn(aq) \rightleftarrows H^{+}(aq) + In^{-}(aq)$$

Color A Color B

$$HIn(aq) + H_2O(l) \rightleftharpoons H_3O^{+}(aq) + In^{-}(aq) \qquad K_a \text{ of } HIn = \frac{[H_3O^{+}][In^{-}]}{[HIn]}$$

Therefore:

$$\frac{[HIn]}{[In^{-}]} = \frac{[H_3O^{+}]}{[K_a]}$$

- The midpoint of color change is observed when $[HIN] = [In^{-}]$
- For an indicator which is a weak base: $BOH(aq) \rightleftarrows B^{+}(aq) + OH^{-}(aq)$

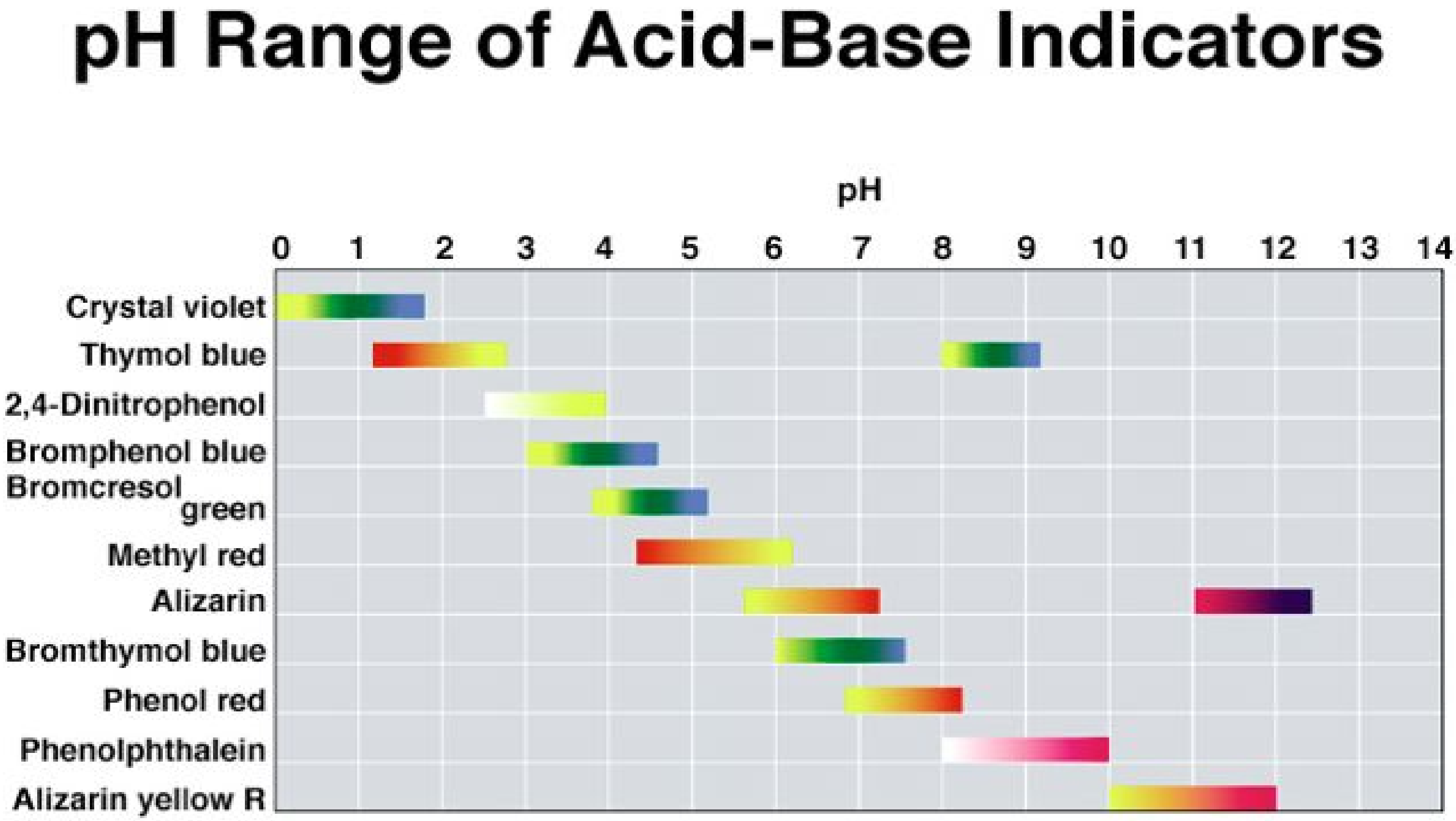
Color A Color B

Selection of an Indicator

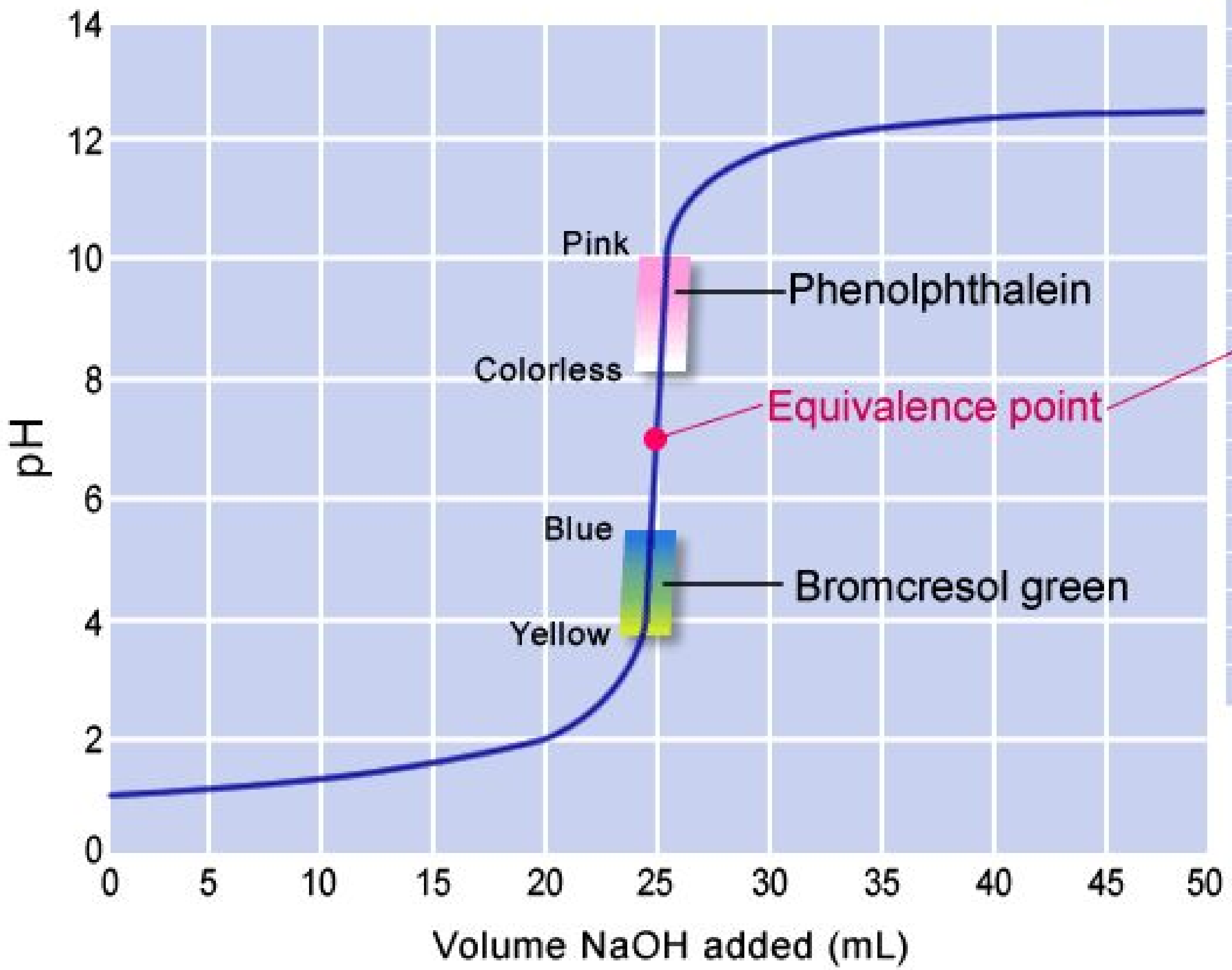
- Most indicators change range within ±1.0 of their pK_a
- pK_a data for indicators can be found in your data booklet

Indicator	pK_a	pH Range	Colour Change Acid	Colour Change Alkali
methyl orange	3.46	3.2–4.4	Red	Yellow
bromophenol blue	4.10	3.0–4.6	Yellow	Blue
bromocresol green	4.90	3.8–5.4	Yellow	Blue
methyl red	5.00	4.8–6.0	Red	Yellow
bromothymol blue	7.30	6.0–7.6	Yellow	Blue
phenol red	8.00	6.6–8.0	Yellow	Red
phenolphthalein	9.50	8.2–10.0	Colourless	Pink

Indicators


pH Range of Acid-Base Indicators
pH
0 1 2 3 4 5 6 7 8 9 10 11 12 13 14
Crystal violet
Thymol blue
2,4-Dinitrophenol
Bromphenol blue
Bromcresol green
Methyl red
Alizarin
Bromthymol blue
Phenol red
Phenolphthalein
Alizarin yellow R

Titration of a Strong Acid



Volume NaOH added (mL)	pH
0	1.00
5	1.18
10	1.37
15	1.60
20	1.95
21	2.06
22	2.20
23	2.38
24	2.69
25	7.00
26	11.29
27	11.59
28	11.75
29	11.87
30	11.96
35	12.22
40	12.36
45	12.46
50	12.52