Class: XI Chemistry
Chapter 7: EQUILIBRIUM
Module 6 of 6

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- **Buffer Solution**
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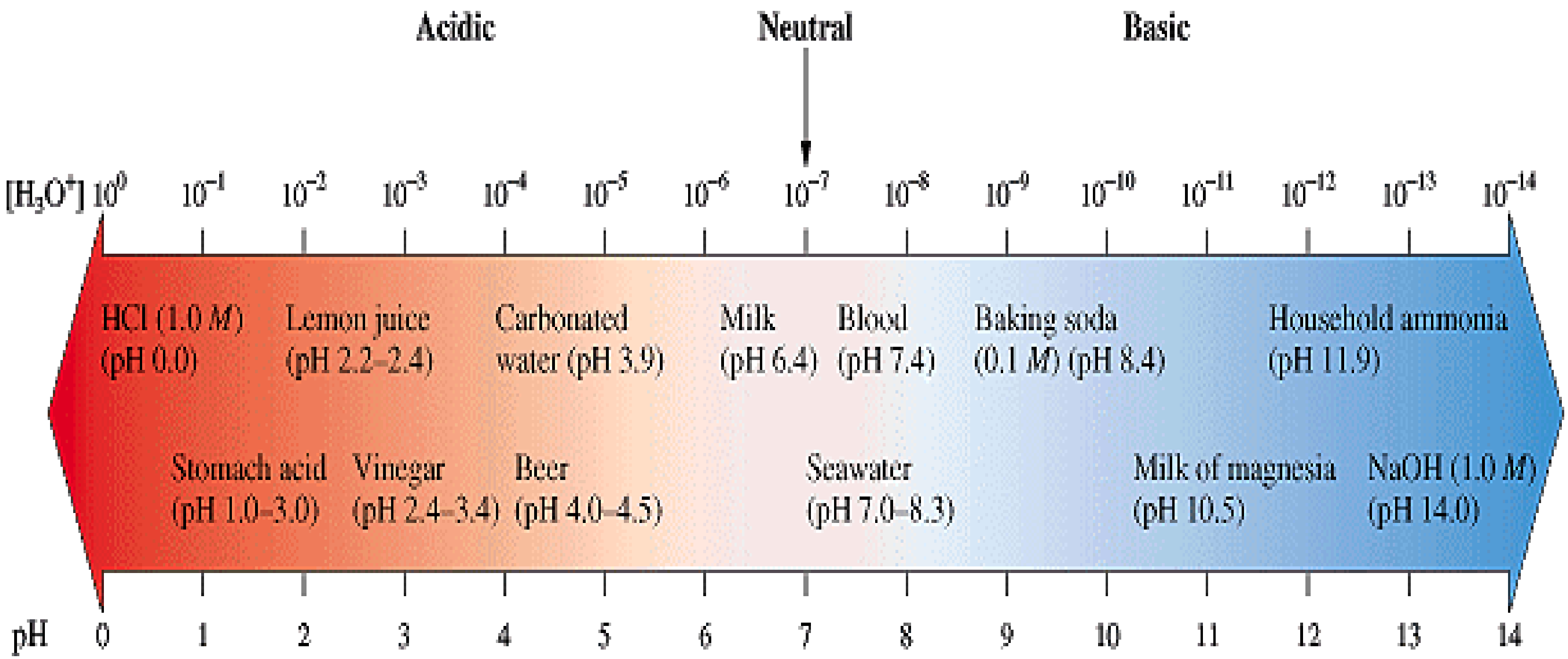
Expressing Hydrogen Ion Concentration
–
pH Scale

Expressing Hydrogen Ion Concentration - pH Scale



Expressing Hydrogen Ion Concentration - pH Scale

- **pH Scale**
  - The pH scale is used to express acidity.
  - For expressing the H_3O^+ ion concentration, a logarithmic scale was devised by P.L.Sorensen.
  - The pH of a solution is defined as the negative logarithm (log) of the hydronium ion (hydrogen ion) concentration.

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

- $[H_3O^+] = 1 \times 10^{-pH}$
- In similar way, pOH is the negative logarithm of the hydroxide ion concentration.

$$pOH = -\log [OH^-]$$

Expressing Hydrogen Ion Concentration - pH Scale

- **In pure water,**

$$K_w = [H_3O^+]\ [OH^-] = 1 \times 10^{-14}\ M^2$$

$$[H_3O^+] = [OH^-] = 1 \times 10^{-7}\ M$$

$$pH = pOH = -\log(1 \times 10^{-7})$$

$$pH = pOH = 7$$

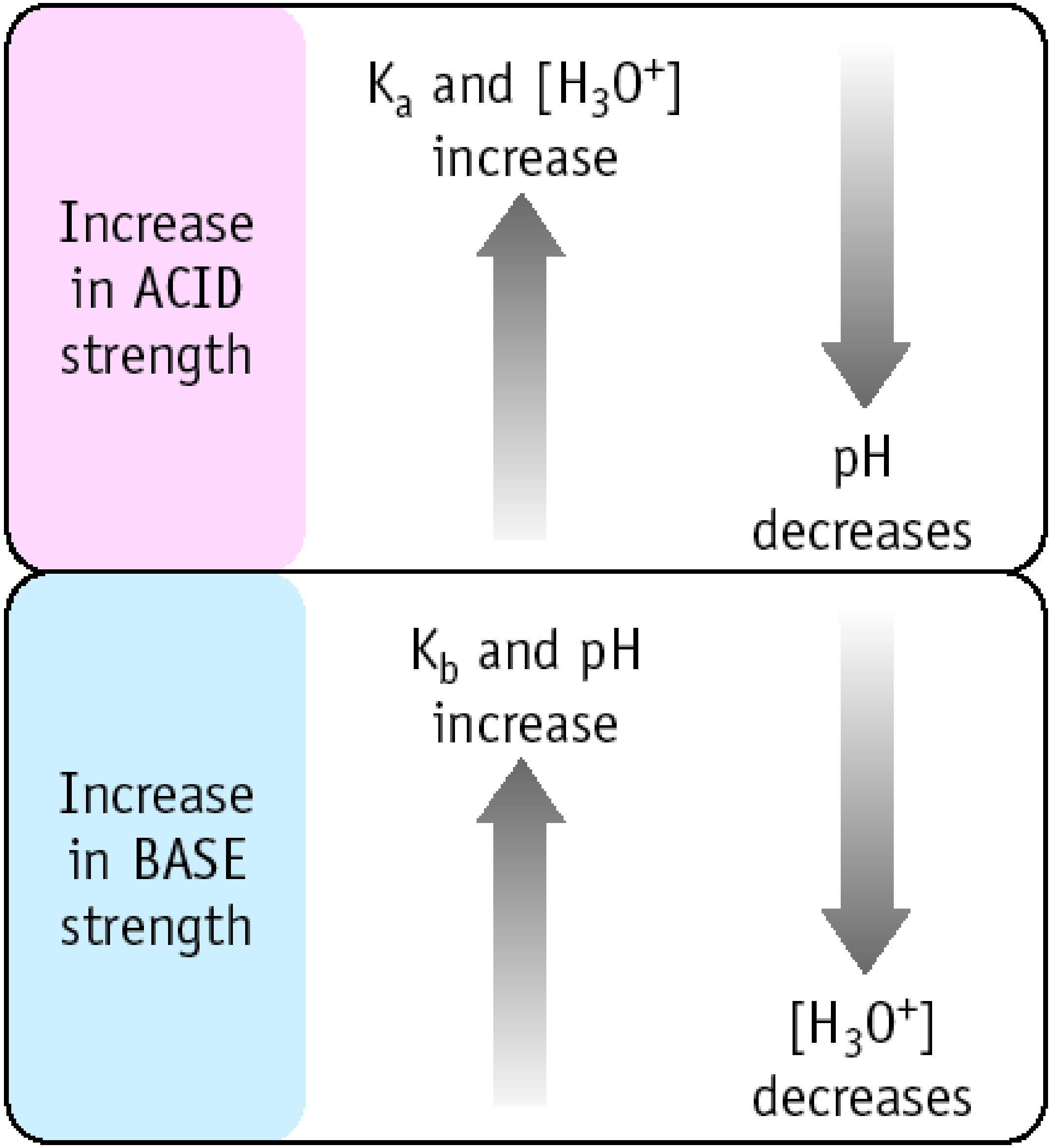
Therefore **$pH + pOH = pK_w = 14$**

Concept is:

At **25 ^oC**,

- **Acidic** solution : **pH < 7.0 ; pOH > 7.0**
- **Basic** solution : **pH > 7.0 ; pOH < 7.0**
- **Neutral** solution : **pH = 7.0 ; pOH = 7.0**

Expressing Hydrogen Ion Concentration - pH Scale



Relation of:

K_a,

K_b,

$[H_3O+]$,

pH

Buffer Solution

Buffer Solution

A solution that maintains its pH when a small amount of a strong acid or a strong base is added to it.

It contains weak acid or weak base with salt that has its conjugate pair.

Two types of buffer solution:

a. acidic buffer solution
b. basic buffer solution

Buffer Solution

- **ACIDIC BUFFER SOLUTION:** Has **pH < 7**.
- An acidic buffer solution is made up of **a weak acid and its salt** (containing its conjugate base)

Example:

- ✓ CH_3COOH / CH_3COONa
- ✓ H_2CO_3 / $NaHCO_3$
- ✓ H_3PO_4 / NaH_2PO_4
- ✓ HCOOH / HCOONa

Buffer Solution

- **pH of Acidic Buffer Solution**
- The pH of acidic is obtained by referring to the equilibrium dissociation of a weak acid, HA.

- Consider buffer solution containing HA and its conjugate, A^-

$$\underset{\uparrow}{HA} \rightleftharpoons H^+ + \underset{\uparrow}{A^{--}}$$

Comes from weak acids

Comes from the ionisation of a salt

Buffer Solution

- **ACIDIC BUFFER SOLUTION:**

Consider an acid buffer solution, containing a weak acid (HA) and its salt (KA) with a strong base(KOH). Weak acid HA ionizes, and the equilibrium can be written as-

$$HA + H_2O \leftrightharpoons H^+ + A^-$$

Acid dissociation constant = Ka

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

Taking, negative log of RHS and LHS:

Cont…

We can write the acidic concentration constant,

$$K_a = \frac{[H^+][A^-]}{[HA]} \quad \text{or} \quad [H^+] = \frac{Ka[HA]}{[A^-]}$$

By applying –log on both sides, we have

$$-\log [H^+] = -\log K_a + \left(-\log \frac{[HA]}{[A^-]} \right)$$

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

pH of acid buffer = pKa + log ([salt]/[acid])

The equation is the Henderson-Hasselbalch equation, popularly known as the Henderson equation.

Buffer Solution

Q. Calculate the pH of a buffer solution containing 0.25 moles / litre of formic acid (HCOOH) and 0.10 moles / litre of sodium formate (HCOONa). Ka for formic acid is 1.8×10^{-4}.

Solution:

Conc. Of acid = 0.25 M

Conc. Of salt = 0.10 M

and **Ka = 1.8×10^{-4}**

So pka is:: pka = -log ka = -log 1.8×10^{-4} = - (log 1.8×10^{-4}) = - (log 1.8 + log 10^{-4}) = - [0.25 + (-4)] = - (- 3.75) = 3.75

Now,

pH = pka + log [salt] / [acid] = 3.75 + log 0.10 / 0.25 = 3.75 – 0.397 = 3.34

Answer: The pH of a buffer solution containing 0.25 M of formic acid and 0.10 M of sodium formate is 3.34.

Buffer Solution

- **Basic Buffer Solution**: Has **pH > 7**
- Basic buffer solution is made up of **a weak base and its salt** (containing its conjugate)
- **Example:**

✓ **NH_4OH / NH_4Cl**

✓ **NH_3 / NH_4Cl**

✓ **NH_3 / $(NH_4)_2CO_3$**

- The **dissociation reactions** are:

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

$$NH_4Cl(aq) \rightarrow NH_4^+(aq) + Cl^-(aq)$$

Buffer Solution

- **<u>pH of Basic Buffer Solution</u>**
- The pH of basic is obtained by referring to the equilibrium dissociation of a weak base, BOH

- Consider buffer solution containing BOH and its conjugate, B+

$$BOH \rightleftharpoons B^{+} + OH^{-}$$

Comes from weak base

Comes from the ionisation of a salt

Buffer Solution

- **Basic Buffer Solution:**

Consider an basic buffer solution, containing a weak base (BOH) and its salt (BCl) with a strong acid(HCl). Weak base BOH ionizes, and the equilibrium can be written as-

$$B(aq) + H_2O(\ell) \rightleftharpoons BH^+(aq) + OH^-(aq)$$

Or

$$BOH(aq) \rightleftharpoons B^+(aq) + OH^-(aq)$$

Base dissociation constant = Kb

$$K_b = \frac{[BH^+]\,[OH^-]}{[B]} \quad \text{or} \quad K_b = \frac{[B^+]\,[OH^-]}{[BOH]}$$

Taking, negative log of RHS and LHS:

Cont…

- The base dissociation constant, K_b

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

By applying – log on both sides:

$$-\log [OH] = -\log K_b + \left(-\log \frac{[B]}{[BH^+]} \right)$$

$$pOH = pK_b + \log \frac{[BH^+]}{[B\ ]}$$

pOH of basic buffer = pKb +log ([salt]/[base])

pH = 14- pOH

The equation is the also called Henderson-Hasselbalch equation, popularly known as the Henderson equation.

Buffer Solution

Buffer Capacity:

The number of millimoles of acid or base to be added to a litre of buffer solution to change the pH by one unit is the Buffer capacity of the buffer.

Formula:

$$\text{BufferCapacity} = \frac{\text{millimoles of acids\& bases}}{(\Delta pH)}$$

Solubility…Begins…

Solubility

The solubility of a substance is the amount of that substance that will dissolve in a given amount of solvent. "Solubility" may be considered to be an equilibrium; the equilibrium is between solid and ions in solution. Any ionic solid is 100% ionized in aqueous solution; once it actually dissolves.

Solubility

Salts are classified on the basis of their solubility in the following table:

Category I	Soluble	Solubility > 0.1M
Category II	Slightly soluble	0.01M< Solubility<0.1M
Category III	Sparingly soluble	Solubility < 0.1M

Solubility

Factors Affecting Solubility

- Temperature
  - Solubility generally increases with temperature.
- Common ion effect
  - Common ions reduce solubility
- pH of solution
  - pH affects the solubility of ionic compounds.
- Formation of complex ion
  - The formation of complex ion increases solubility

Solubility

SOLUBILITY EQUILIBRIUM

- **Some salts are soluble but most are insoluble or slightly soluble in water.**
- **A saturated solution is a solution that contains the maximum amount of solute that can dissolve in a solvent.**
- **The solubility of a salt is the amount of solid that dissolved in a known value of saturated solution.**
- **The unit of solubility used may be g L^{-1} or mol L^{-1}**
- **Molar solubility is the maximum number of moles of solute that dissolves in a certain quantity of solvent at a specific temperature.**

Solubility

- *Different types of solution*

Unsaturated solution: More solute can be dissolved in it.

Saturated solution: No more solute can be dissolved in it. Any more of solute you add will not dissolve. It will precipitate out.

Super saturated solution: Has more solute than can be dissolved in it. The solute precipitates out.

Solubility Product

THE SOLUBILITY PRODUCT CONSTANT, K_{sp}

K_{sp} is the product of the molar concentrations of the ions involved in the equilibrium, each raised to the power of its stoichiometric coefficient in the equilibrium equation.

- **K_{sp} is called the solubility product constant.**
- **The degree of solubility of a salt is shown by the value of K_{sp}.**
- **Soluble salt such NaCl and KNO_3 has an extremely high value of K_{sp} .**
- **The smaller the value of K_{sp} the less soluble the compound in water.**
- **Temperature ↑ , solubility ↑, K_{sp} ↑**

Solubility

- Consider the equilibrium system below :

$$MX\ (s) \rightleftharpoons M^{+}\ (aq) + X^{-}\ (aq)$$

$$K_c = \frac{[M^+]\,[X^-]}{[MX]}$$

$$K_c\,[MX] = [M^+]\,[X^-]$$

*since [MX] is a constant ;

$$K_{sp} = [M^+]\,[X^-]$$

Solubility Product

- ✓ The solubility equilibrium equation for a slightly soluble salt, MA :

$$MA\ (s) \rightleftharpoons M^{+}\ (aq) \quad + \quad A^{-}\ (aq)$$

$$K_{sp} = [M^{+}]\ [A^{-}]$$

- ✓ If we mix a solution containing M^{+} ions with one containing A^{-} ions, the ion product, Q_{sp} is given by :

$$Q_{sp} = [M^{+}]\ [A^{-}]$$

Solubility

- In General, for any sparingly soluble salt AxBy , which dissociates to set-up the equilibrium

$$A_xB_y \rightleftharpoons {}_xA^{y+} + {}_yB^{x-}$$

The solubility product constant may be expressed as

$$K_{sp} = [A^{y+}]^x \; [B^{x-}]^y$$

Suppose at a particular temperature, its solubility is S mol L^{-1} S^{-1}. S moles of salts on ionization will give xS moles of A^{y+} and yS moles of B^{x-}. So

$$K_{sp} = [xS]^x + [yS]^y$$

$$K_{sp} = x^x y^y (S)^{x+y}$$

Solubility Product

<u>Three possible situations:</u>

$Q_{sp} < K_{sp}$;Solution is not saturated.
Solid will dissolve and no precipitate formed.

$Q_{sp} = K_{sp}$; Saturated solution formed.
System is in equilibrium.

$Q_{sp} > K_{sp}$; Solution is supersaturated;
Ions will form precipitate until the ionic concentration product of the system equals the K_{sp} (until the system reaches equilibrium).

Module 6 of 6 Ends…