Acids and Bases

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Characteristics Of Acids

Acids can be characterized by:

1. A sour taste.
2. It turns blue litmus paper red
3. It tastes sour. Try drinking lemon juice (citric acid)

Characteristics of Bases

A Base is characterized by:

1. A bitter taste. (Milk of Magnesia)
2. It feels slippery. (Soapy Water)
3. It turns Red Litmus Blue.

Acids & Bases

Types

Arrhenius

Acid - Substances in water that increase the concentration of hydrogen ions (H^+).

Base - Substances in water that increase concentration of hydroxide ions (OH^-).

Acids & Bases (Con't)

Bronsted-Lowry

Acid - Neutral molecule, anion, cation which donates a proton.

Base - Neutral molecule, anion, cation which accepts a proton.

$$HA + :B \rightarrow HB^{+} + :A^{-}$$

$$HCl + H_2O \rightarrow H_3O^{+} + Cl^{-}$$

Acid Base Conj Acid Conj Base

1. $H_3O^+(aq) + Cl^-(aq) + NH_3 \rightarrow Cl^-(aq) + NH_4^+(aq)$
2. $HCl(benzene) + NH_3(benzene) \rightarrow NH_4Cl(s)$
3. $HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$

Both definitions work for the first example, where water is the solvent and hydronium ion is formed. The next two reactions do not involve the formation of ions but are still proton transfer reactions. In the second reaction hydrogen chloride and ammonia (dissolved in benzene) react to form solid ammonium chloride in benzene solvent and in the third gaseous HCl and NH_3 combine to form the solid.

Bronsted-Lowry Acid/Base Examples

Acid: proton donor

Base: proton acceptor

Note: Water can act as acid or base

Acid		Base		C Acid		C Base
HCL	+	H_2O	$\Leftrightarrow$	H_3O^+	+	Cl^-
$H_2PO_4^-$	+	H_2O	$\Leftrightarrow$	H_3O^+	+	HPO_4^{2-}
NH_4^+	+	H_2O	$\Leftrightarrow$	H_3O^+	+	NH_3

Base		Acid		C Acid		C Base
$:NH_3$	+	H_2O	$\Leftrightarrow$	NH_4^+	+	OH^-
PO_4^{3-}	+	H_2O	$\Leftrightarrow$	HPO_4^{2-}	+	OH^-

Acids & Bases

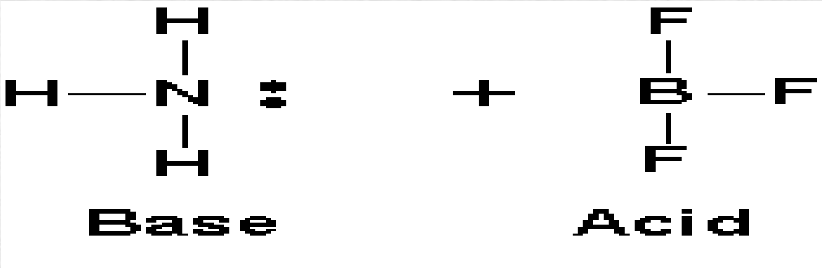
Types

Lewis

Acid - An electron pair acceptor

Base - An electron pair donor

All Brønsted acids are also Lewis acids, but not all Lewis acids are Brønsted acids.



Conjugate Acid-Base Pairs

Conjugate Base	- Species remaining after an acid has transferred its proton.
Conjugate Acid	- Species produced after base has accepted a proton.
HA & $:A^-$	- conjugate acid/base pair
$:A^-$	- conjugate base of acid HA
$:B$ & HB^+	- conjugate acid/base pair
HB^+	- conjugate acid of base :B

Acid – Base Strength

Strong Acid - Transfers all protons to water;
completely ionizes;
strong electrolyte;

conjugate base is weaker and has
negligible tendency to be protonated.

Weak Acid - Fraction of protons transferred to water;
partly ionized;
weak electrolyte;
conjugate base is stronger readily
accepting protons from water

- As acid strength decreases, base strength increases.
- The stronger the acid, the weaker its conjugate base
- The weaker the acid, the stronger its conjugate base

Dissociation Constants of Acids & Bases

Strong Base - All molecules accept a proton;
completely ionizes;
strong electrolyte;
conjugate acid is weaker with negligible tendency to donate Protons.

Weak Base - Fraction of molecules accept proton;
partly ionized;
weak electrolyte;
conjugate acid is stronger that readily donates protons.

- As base strength decreases, acid strength increases.
- The stronger the base, the weaker its conjugate acid.
- The weaker the base the stronger its conjugate acid.

Acid – Base Dissociation

Acid-base reactions are equilibrium processes.

The relationship between the relative concentrations of the reactants and products is a constant (temperature dependent) referred to as the Acid or Base Dissociation Constant.

The stronger the acid or base, the larger the value of the dissociation constant.

HA <-----> A^- + H^+

the equilibrium constant can be calculated from the equation:

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

For an acid in water

$$K_{eq} = \frac{[:A^-][H_3O^+]}{[HA][H_2O]}$$

For a base in water

$$K_{eq} = \frac{[HB][OH^-]}{[:B^-][H_2O]}$$

Note :

$$H_3O^+ = [H^+]$$

$[H_2O]$ **in dilute solutions is constant.**

$$\therefore \quad K_{eq}[H_2O] = K_a = \frac{[:A^-][H^+]}{[HA]} \qquad K_{eq}[H_2O] = K_b = \frac{[HB][OH^-]}{[:B^-]}$$

Water as an Equilibrium System

Water has the ability to act as either a Bronsted-Lowry acid or base.

Autoionization – Spontaneous formation of low concentrations of [H^+] and OH^-] ions by proton transfer from one molecule to another.

Equilibrium Constant for Water

$$K_c = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

$$K_c[H_2O]^2 = [H_3O^+][OH^-]$$

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14} \text{ (at } 25^\circ\text{C)}$$

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ (at } 25^\circ\text{C)}$$

In Pure Water :

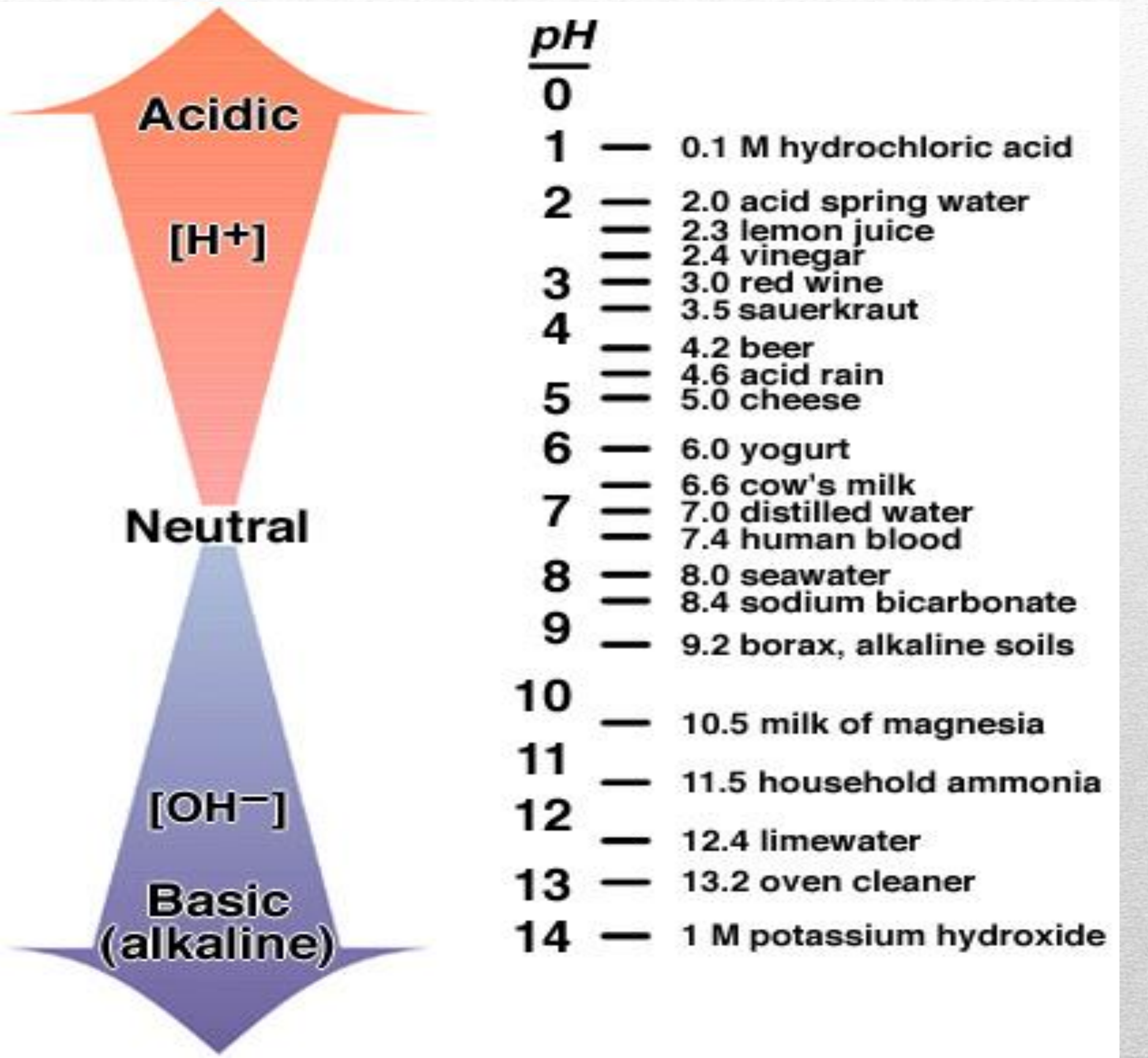
$$[H^+] = [OH^-] = 1.0 \times 10^{-7}$$

Add protons to the solution, i.e., add an acid

- Number of OH^- ions will decrease by reacting with the protons.
- Number of water molecules will increase.
- Why? Value of K_w must be maintained.

What is the pH scale

- The pH scale measures how *acidic* or *basic* a solution is.



The pH scale

- The pH scale is the concentration of hydrogen ions in a given substance.

$$pH = -\log\left[H^+\right]$$

Identifying Acids and Bases

- *Acids* have a ph from 0-7
- Lower pH value indicates a stronger acid.

- *Bases* have a pH from 7-14
- Higher pH value indicates a stronger base.

Acid

- Any substance which has a pH of value of less than 7 is considered an acid

0---------------7----------------14

Acid Neutral Base

Base

- Any substance which has pH value greater than 7 is a base

0--------------7---------------14

Acid Neutral Base

pH 7

- A pH of 7 is called neutral—neither acid nor base.

0------------7------------14

Acid Neutral Base

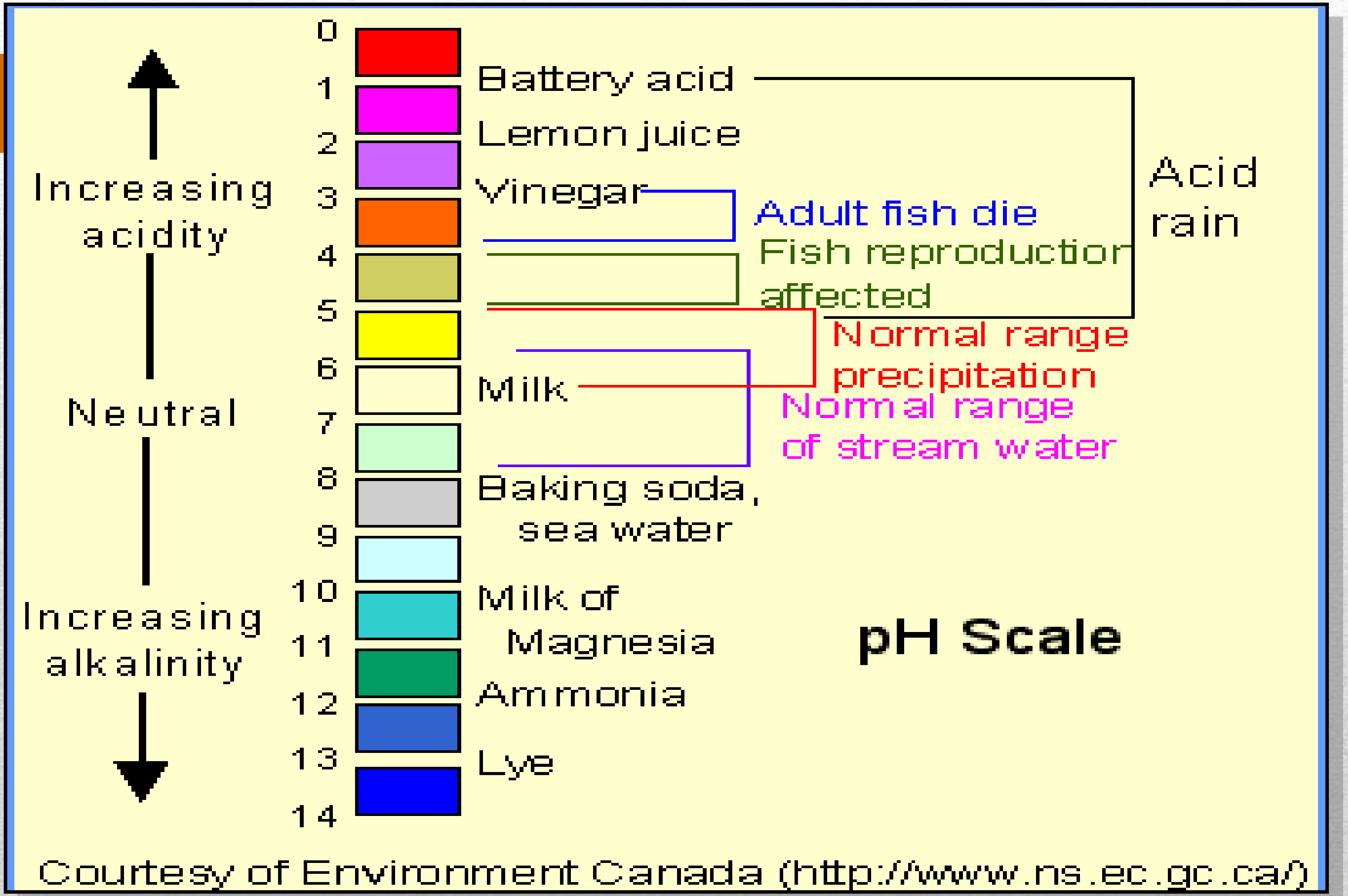
Acidic or Basic

- If the number is less than 7 the soil or water is acidic
- If the number is more than 7 the soil or water is basic

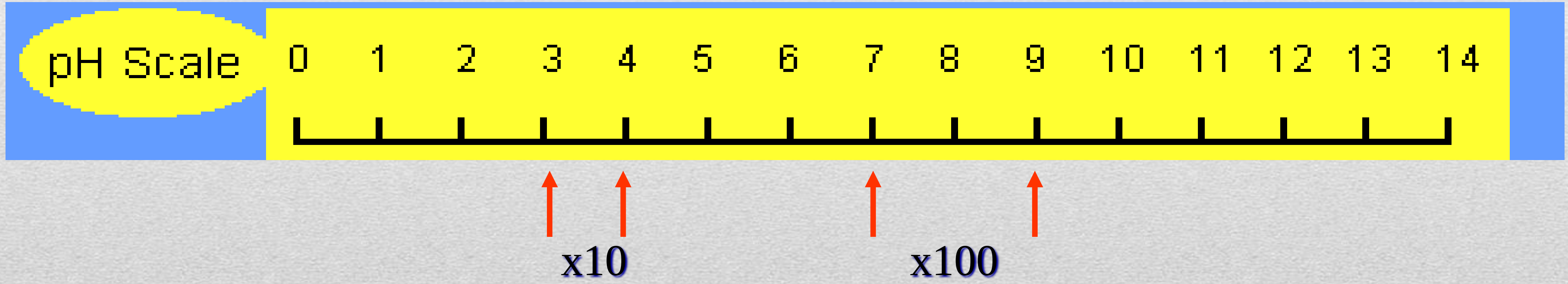
The pH Scale

- pH scale ranges from 0 -14
- pH 7 is neutral; neither acid nor base
- Pure water is pH 7
- Low pH = acid
- High pH = base
- The closer to the ends of the scale, the stronger the solution is

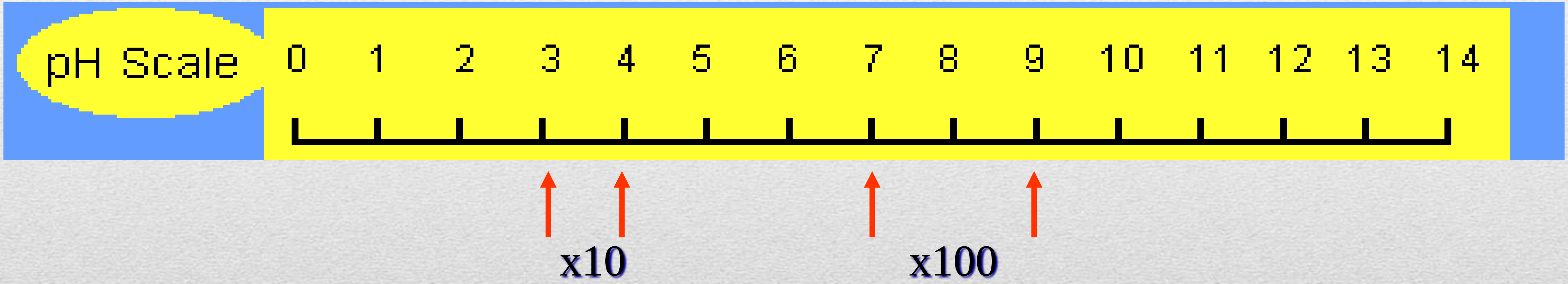
The pl


Increasing
acidity
Neutral
Increasing
alkalinity
0
1
2
3
4
5
6
7
8
9
10
11
12
13
14
Battery acid
Lemon juice
Vinegar
Milk
Baking soda,
sea water
Milk of
Magnesia
Ammonia
Lye
Adult fish die
Fish reproduction
affected
Normal range
precipitation
Normal range
of stream water
Acid
rain
pH Scale
Courtesy of Environment Canada (http://www.ns.ec.gc.ca/)

- Each pH unit is *10 times* as large as the previous one
- A change of 2 pH units means *100 times* more basic or acidic

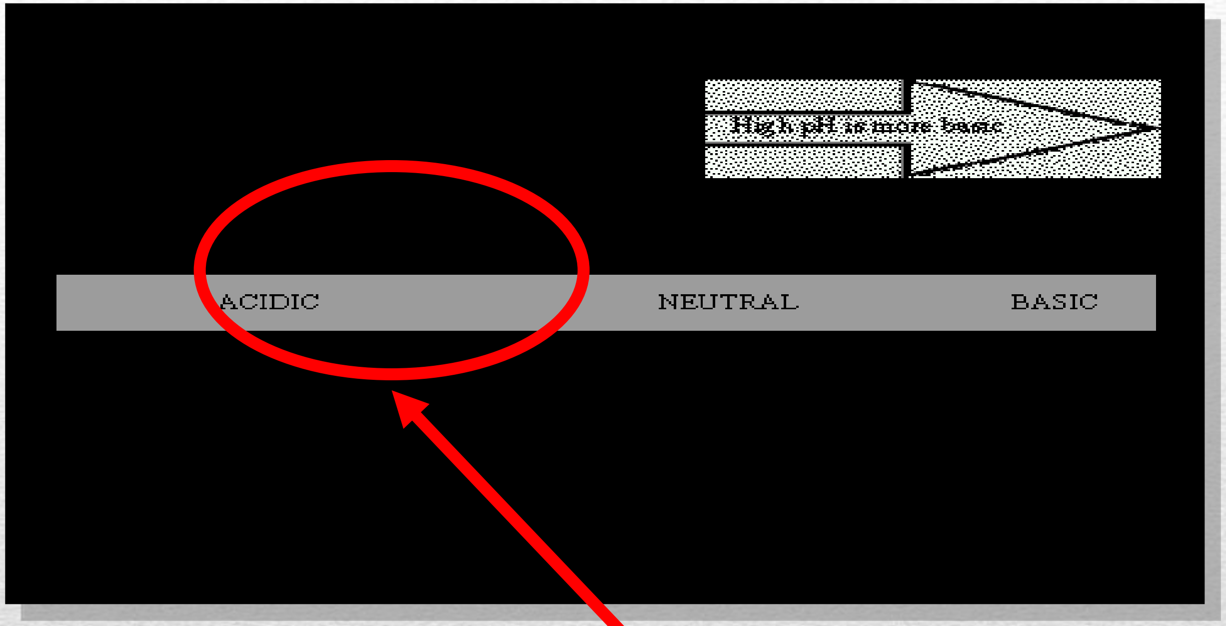


The pH Scale



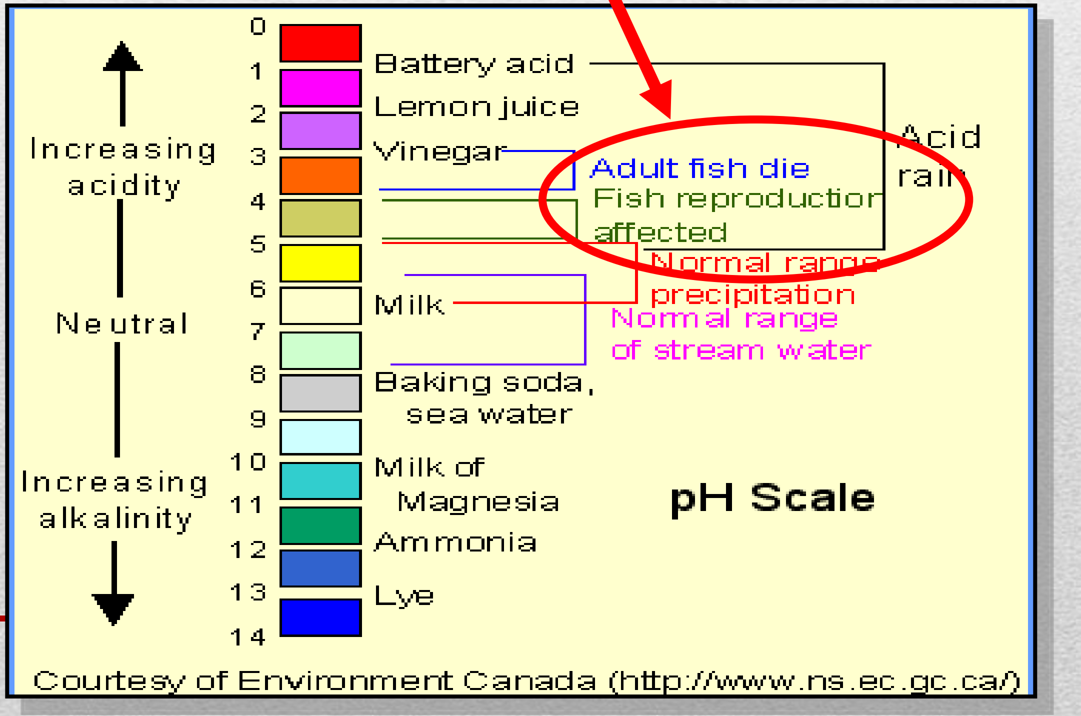
- Careful measurement is important
- A mistake of *one* pH unit means *10 times* too much or too little!

Why is pH important?

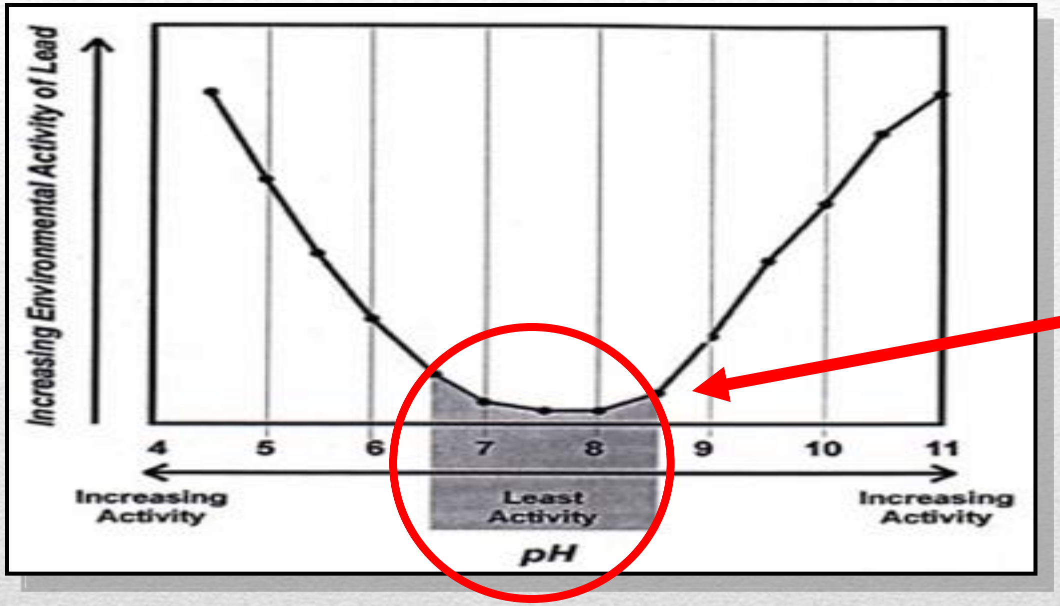


Fish can't live if the pH is too high or too low

Soil has to be in a certain pH range for plants to grow and stay healthy.



pH and People



Water that has too high or low pH may contain harmful dissolved chemicals.

Water plant operators keep a careful watch on the pH of our drinking water, to keep it safe.

pH

1. Acidity or Acid Strength is a function of Hydrogen Ion Concentration ($[H^+]$)
2. The pH system is a logarithmic representation of the Hydrogen Ion concentration (or OH^-) as a means of avoiding using large numbers and powers.

$$pH = -\log[H^+] = \log(1 / H^+])$$

$$pOH = -\log[OH^-] = \log(1 / OH^-])$$

3. In pure water $[H^+] = 1 \times 10^{-7}$ mol / L

$$\therefore\ pH = -\log(1 \times 10^{-7)} = -(0 - 7) = 7$$

4. pH range of solutions: 0 - 14

$pH < 7$ (Acidic) $[H^+] > 1 \times 10^{-7}$ m / L

$pH > 7$ (Basic) $[H^+] < 1 \times 10^{-7}$ m / L

5. pH is measured directly with a pH meter

Questions?

BUFFERS

Why is pH important in biology?

- pH affects solubility of many substances.
- pH affects structure and function of most proteins - including enzymes.
- Many cells and organisms (esp. plants and aquatic animals) can only survive in a specific pH environment.
- Important point -
  - pH is dependent upon temperature

Buffers

- Definition: a solution that resists change in pH
  - Typically a mixture of the acid and base form of a chemical
  - Can be adjusted to a particular pH value

$$H_2O + CO_2 \leftrightarrow H_2CO_3 \leftrightarrow H^+ + HCO_3^-$$

Blood: pH = 7.35-7.45

Too acidic? Increase respiration rate expelling CO2, driving reaction to the left and reducing H+ concentration.

Excretory system – excrete more or less bicarbonate

Buffers

- Definition: a solution that resists change in pH
  - Typically a mixture of the acid and base form of a chemical
  - Can be adjusted to a particular pH value
- Why use them?
  - Enzyme reactions and cell functions have optimum pH’s for performance
  - Important anytime the structure and/or activity of a biological material must be maintained

How buffers work

- Equilibrium between acid and base.
- Example: Acetate buffer
  - $CH_3COOH \leftrightarrow CH_3COO^- + H^+$
- If more H^+ is added to this solution, it simply shifts the equilibrium to the left, absorbing H^+, so the $[H^+]$ remains unchanged.
- If H^+ is removed (e.g. by adding OH-) then the equilibrium shifts to the right, releasing H^+ to keep the pH constant

Limits to the working range of a buffer

- Consider the previous example:
  - $CH_3COOH \leftrightarrow CH_3COO^- + H^+$
- If too much H^+ is added, the equilibrium is shifted all the way to the left, and there is no longer any more CH_3COO^- to “absorb” H^+.
- At that point the solution no longer resists change in pH; it is useless as a buffer.
- A similar argument applies to the upper end of the working range.

Chemistry of buffers

- K_a = equilibrium constant for H+ transfer… also described as the dissociation constant…the tendancy of an acid to dissociate. AH → A- (base conjugant) + H+
- K_a = [A-] [H+]/ [AH] = [base] [H+] / [acid]
- Weak acids have low values… contribute few H+ ions…
- Because we are usually dealing with very small concentrations, log values are used…
- The log constant = $pK_a = -\log_{10} K_a$

Chemistry of buffers

- K_a = [A-] [H+]/ [AH] = [base] [H+] / [acid]
- Weak acids have low values… contribute few H+ ions…
- Because we are usually dealing with very small concentrations, log values are used…
- The log constant = $pK_a = -\log_{10} K_a$
- SO! Since pK is the negative log of K, weak acids have high values … (-2 – 12).
- HCl = -9.3 – very low ~complete dissociation

Chemistry of buffers

- First rearrange the first equation and solve for [H+]
  - $[H+] = K_a \times [acid]/[base]$
- Then take the log of both sides
  - $\log_{10}[H+] = \log_{10}K_a + \log_{10} [acid]/[base]$

$-pH$ $\qquad$ $-pK_a$

Chemistry of buffers

- $-pH = -pKa + \log_{10} [acid]/[base]$
- Multiply both sides by –1 to get the **Henderson-Hasselbach equation**
    - $\mathbf{pH = pKa - \log_{10} [acid]/[base]}$

Chemistry of buffers

- What happens when the concentration of the acid and base are equal?
  - Example: Prepare a buffer with 0.10M acetic acid and 0.10M acetate
    - pH = pKa - $\log_{10}$ [acid]/[base]
    - pH = pKa - $\log_{10}$ [0.10]/[0.10]
    - pH=pKa
    - Thus, the pH where equal concentrations of acid and base are present is defined as the pKa

- A buffer works most effectively at pH values that are $\pm$ 1 pH unit from the pKa (the **buffer range**)

Factors in choosing a buffer

- Be sure it covers the pH range you need
  - Generally: pK_a of acid $\pm$ 1 pH unit
  - Consult tables for ranges or pK_a values
- Be sure it is not toxic to the cells or organisms you are working with.
- Be sure it would not confound the experiment (e.g. avoid phosphate buffers in experiments on plant mineral nutrition).

What to report when writing about a buffer:

- The identity of the buffer (name or chemicals)
- The molarity of the buffer
- The pH of the buffer
- Examples:
  - "We used a 0.5M Tris buffer, pH 8.0."
  - "The reaction was carried out in a 0.1M boric acid – sodium hydroxide buffer adjusted to pH 9.2."

Three basic strategies for making a buffer

1. Guesswork – mix acid and base at the pH meter until you get the desired pH.
   - Wasteful on its own, but should be used for final adjustments after (2) or (3).
2. Calculation using the Henderson-Hasselbach equation.
3. Looking up recipe in a published table.

Calculating buffer recipes

- Henderson-Hasselbach equation
  - **$pH = pKa - \log_{10} [acid]/[base]$**
- Rearrange the equation to get
  - **$10^{(pKa-pH)} = [acid]/[base]$**
- Look up pKa for acid in a table. Substitute this and the desired pH into equation above, and calculate the approximate ratio of acid to base.
- Because of the log, you want to pick a buffer with a pKa close to the pH you want.

Example

- You want to make about 500 mL of 0.2 M acetate buffer (acetic acid + sodium acetate), pH 4.0.
- Look up pKa and find it is 4.8.
- **$10^{(4.8 - 4.0)} = 10^{0.8} = 6.3 = [\text{acid}]/[\text{base}]$**
- If you use 70 mL of base, you will need 6.3X that amount of acid, or 441 mL. Mix those together and you have 511 mL (close enough).