Acid & Base

Arrhenius Definitions

Acid - Something that generates H^+ in H_2O

Base - Somethin that generates OH^- in H_2O

Brönsted-Lowry Definition

$\left.\begin{array}{l}\text{Acid - A proton donor} \\ \text{Base - A proton acceptor}\end{array}\right\}$ Chem

Lewis Definition

Looks at acids & bases as a transfer of e^- $\Rightarrow$ use in Chem 1A

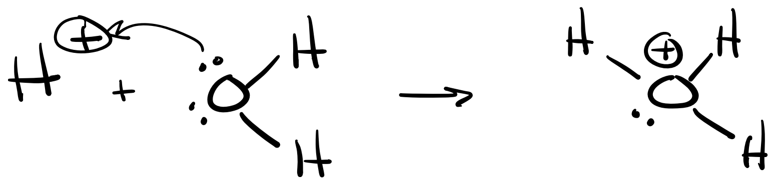
Proton – A hydrogen ion H^+

$$H \longrightarrow H^+ + e^-$$

$1p^+$	$1p^+$ ⟸ only has the 1 proton
$1e^-$	$\emptyset e^-$
$\emptyset n$	$0 n$

Hydrogen ion = Proton

H^+ unstable on their own & like to bond with H_2O



H_3O^+ hydronium ion

H^+ = H_3O^+

proton — hydronium ion

Recognizing Acids

Acids usually written with the acidic proton(s) on left

HCl hydrochloric acid

HBr hydrobromic acid

HNO_3 Nitric acid

H_2SO_4 Sulfuric acid

$HC_2H_3O_2$ acetic acid

not acidic

Recognizing Base

Usually metal hydroxides or anions

$NaOH$ Sodium hydroxide

KOH Potassium hydroxide

$Ca(OH)_2$ Calcium hydroxide

Na_2SO_4 (SO_4^{2-}) Sulfate

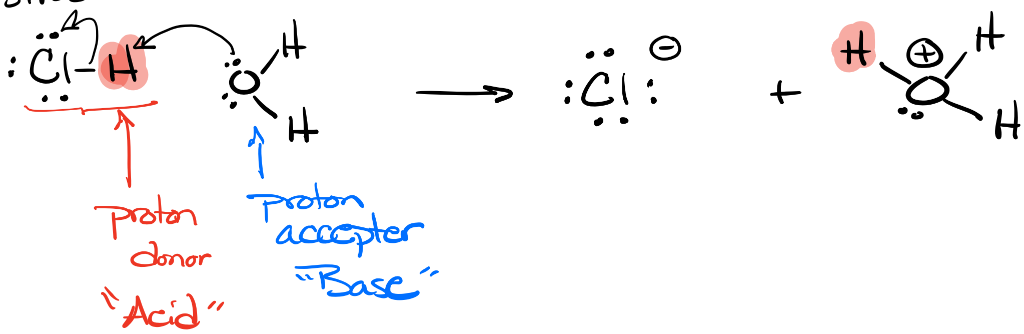
Na_2CO_3 (CO_3^{2-}) Carbonate

KCl (Cl^-) Chloride

Chemical Equation

$$\underset{\text{Acid}}{HCl_{(aq)}} + \underset{\text{Base}}{H_2O_{(l)}} \longrightarrow Cl^{\ominus}_{(aq)} + H_3O^{\oplus}_{(aq)}$$

Lewis structures



Acid

$$H_2PO_4^- + \underset{\text{Base}}{CO_3^{2-}} \longrightarrow HPO_4^{2-} + HCO_3^-$$

Has accepted an H^+

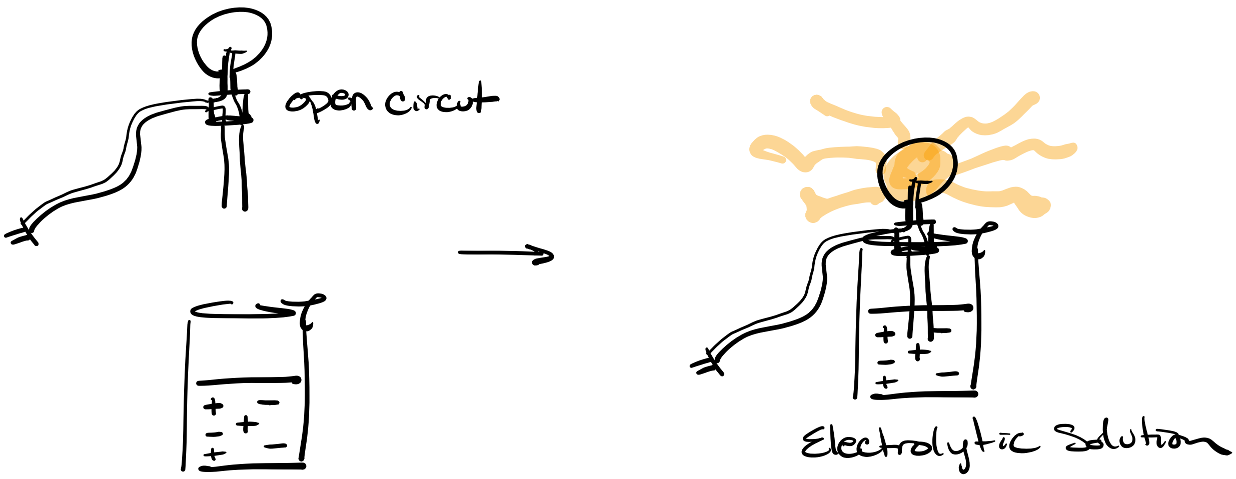
Acid/Base ?

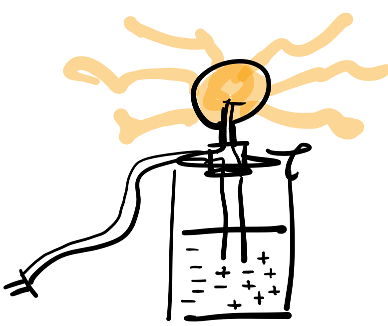
lost an H^+
or donated a proton

Strength of Acid/Base

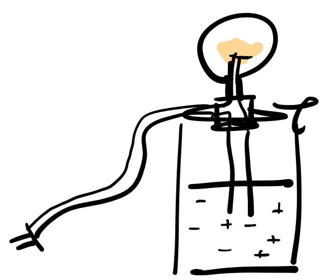
Degree of ionization → formation of ions

Electrolytic solutions → Conduct electricity due to the presence of ions.





Lots of ions = Strong electrolyte & light is bright



Few ions = weak electrolyte & light is dim

Strong electrolytes

$NaCl$

Na_2SO_4

K_3PO_4

NH_4OH

$LiBr$

Group IA cations ⇒ always soluble

NH_4^+ always soluble

All highly soluble

Weak electrolytes

$FeCl_3$

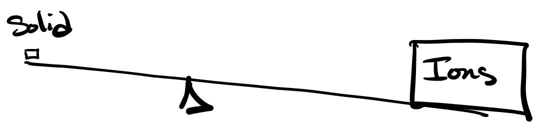
$AlCl_3$

$MgCO_3$

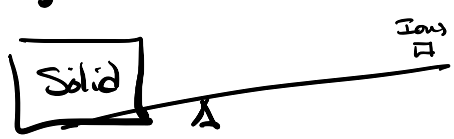
$CaCl_2$

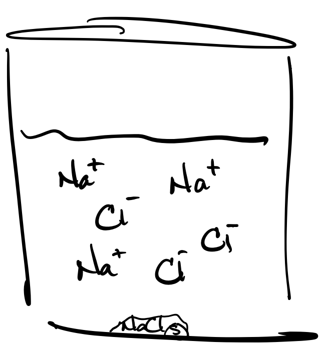
metals & groups that are much less soluble

$$NaCl_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + Cl^-_{(aq)}$$

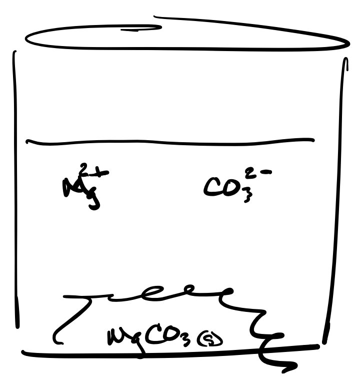


$$MgCO_{3(s)} \xrightarrow{H_2O} Mg^{2+}_{(aq)} + CO_{3(aq)}^{2-}$$





Strong electrolyte

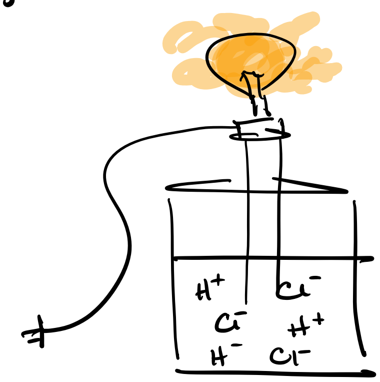


weak electrolyte

Strong Acid = Strong Electrolyte

$\Rightarrow$ dissociates Completely

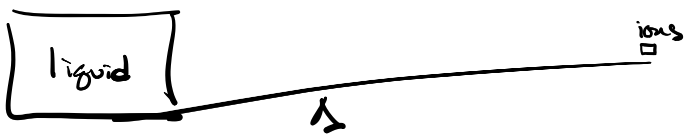
$$HCl_{(g)} \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$$

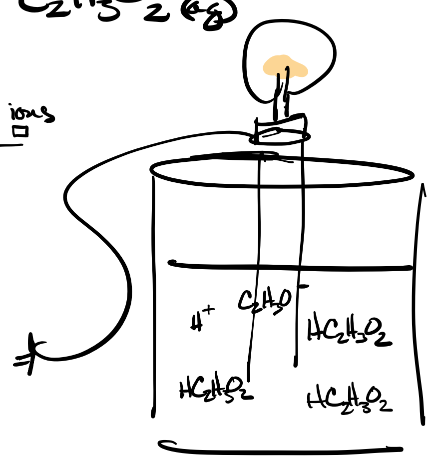


Weak acid = weak Electrolyte

$\Rightarrow$ dissociates only a little

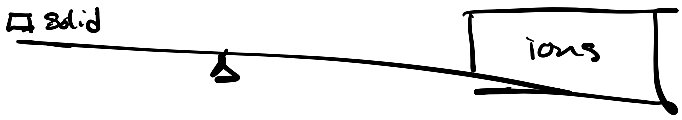
$$HC_2H_3O_{2(l)} \xrightarrow{H_2O} H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$





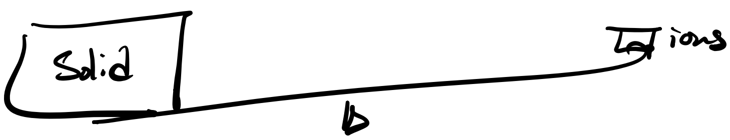
strong base = strong electrolyte

$$NaOH_{(s)} \xrightarrow{H_2O} Na^{\oplus}_{(aq)} + OH^{\ominus}_{(aq)}$$



Weak base = weak electrolyte

$$Na_2CO_{3\,(s)} \longrightarrow 2\,Na^{+}_{(aq)} + CO_{3\,(aq)}^{2-}$$



Strong acids	Strong Base
HCl	metal hydroxides
HNO_3	$NaOH$
H_2SO_4	$LiOH$
	KOH

If not one of these 3 it's a weak acid

If not a metal hydroxide than it's a weak base

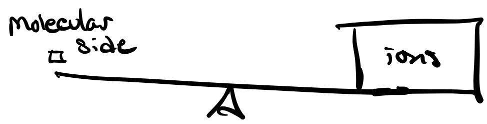
Ways of writing these equations

$\longrightarrow$ forward arrow $\Rightarrow$ implies strong

$\rightleftharpoons$ Equilibrium arrow $\Rightarrow$ implies weak

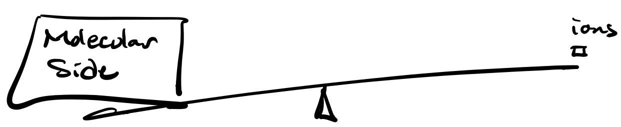
$$HCl_{(g)} + H_2O_{(l)} \xrightarrow{\text{Strong}} H_3O^{+}_{(aq)} + Cl^{-}_{(aq)}$$

Strong Acid



$$HC_2H_3O_{2(l)} + H_2O_{(l)} \overset{\text{weak}}{\rightleftharpoons} H_3O^{+}_{(aq)} + C_2H_3O_2{}^{-}_{(aq)}$$

weak acid



Chemical Species

$$aA + bB \rightleftharpoons cC + dD$$

Coefficients

How far the reaction progresses towards products can be measured.

We measure the progress by measuring the concentration of each species in mols/L

$$\text{mols}/L = [\ ]$$

$$[A] = \text{mols}/L \text{ of } A$$

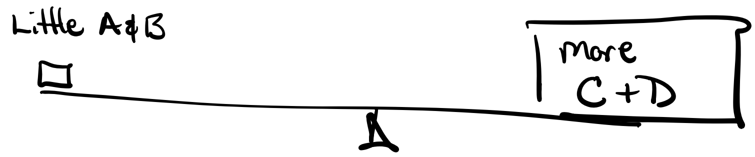
Reaction Quotient = a measure of progress of a reaction

$$aA + bB \rightleftharpoons cC + dD$$

$$\text{Reaction Quotient} = Q = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

When the reaction "stops" $\Rightarrow$ reaches equilibrium, the reaction quotient has a special meaning.

$$aA + bB \underset{}{\overset{Rate_{forward} = Rate_{Reverse}}{\rightleftharpoons}} cC + dD$$



$$\text{Reaction Quotient @ Eq} = Q_{eq} = K = \frac{[C]_{eq}^{c}\,[D]_{eq}^{d}}{[A]_{eq}^{a}\,[B]_{eq}^{b}}$$

Value of Reaction Quotient at eq

What is the value of k when a reaction favors products?

$$aA + bB \rightleftharpoons cC + dD$$

$$k = \frac{[C]^c [D]^d}{[A]^a [B]^b} = \frac{1000}{0.1} = 10000 >> 1$$

When a reaction favors the products the $k >> 1$

What is the k value when reaction favors Reactants?

$$aA + bB \rightleftharpoons cC + dD$$

$$K = \frac{[C]^c [D]^d}{[A]^a [B]^b} = \frac{0.1}{1000} = 0.0001 << 1$$

When a reaction favors Reactants $K << 1$

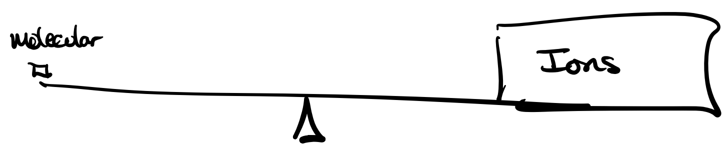
Strong Acid

Higly favored

$$HCl_{(g)} + H_2O \rightleftharpoons \boxed{H_3O^+_{(aq)} + Cl^-_{(aq)}}$$

$K = 1 \times 10^7$ @ 25°C

$1 \times 10^7 >> 1$



Weak Acid

Reactants favored

$$\boxed{HC_2H_3O_{(l)} + H_2O_{(l)}} \rightleftharpoons H_3O^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$

$K = 1 \times 10^{-5}$

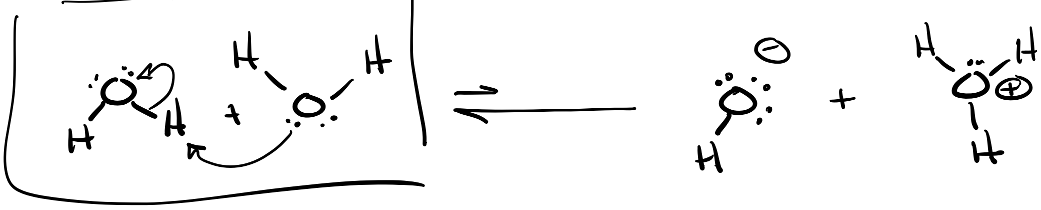
$1 \times 10^{-5} << 1$

$$OH^{\ominus} \xleftarrow{-H^{\oplus}} H_2O \xrightarrow{+H^{\oplus}} H_3O^{+}$$

Base & Acid

Amphoteric – both acid & base

Highly favored



Auto dissociation of water

at equilibrium at 25°C

$$K = 1 \times 10^{-14}$$

$$K = \frac{[H_3O^+]^1 [OH^-]^1}{[H_2O]^2}$$

$$K[H_2O]^2 = \underbrace{K_w}_{\text{equilibrium constant for water}} = [H_3O^+][OH^-] = 1\times10^{-14}\ \frac{\text{mols}^2}{L^2}$$

$[H_3O^+] = 1 \times 10^{-7}$ mols/L

$[OH^-] = 1 \times 10^{-7}$ mols/L

pH = power of hydrogen

= a math function that converts to a log scale

$P = -\log$

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

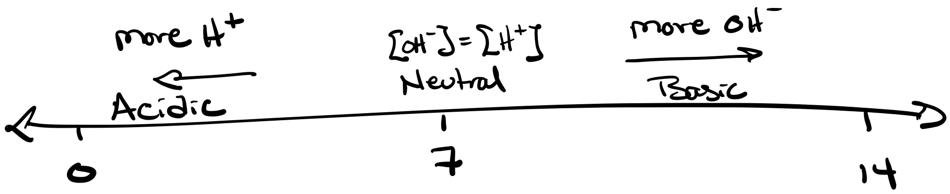
pH when $[H^+] = 1 \times 10^{-7}$ $\cancel{\text{mols/L}}$

$-\log[1 \times 10^{-7}]$

$\log A = 10^{?} = A$

$-(-7)$

$pH = 7$



pH Range

pH	$[H^+]$		$[OH^-]$		Kw
0	1×10^{0}	×	1×10^{-14}		1×10^{-14}
⋮					
4	1×10^{-4}	×	1×10^{-10}		1×10^{-14}
5	1×10^{-5}	×	1×10^{-9}		1×10^{-14}
6	1×10^{-6}	×	1×10^{-8}		1×10^{-14}
7	1×10^{-7}	×	1×10^{-7}	=	1×10^{-14}
8	1×10^{-8}	×	1×10^{-6}		1×10^{-14}
9	1×10^{-9}	×	1×10^{-5}		1×10^{-14}
⋮					
14	1×10^{-14}	×	1×10^{0}		1×10^{-14}

Relationships

$$
\begin{array}{ccccc}
[H^+] & \times & [OH^-] & = & 1\times10^{-14} \\
{\scriptstyle -\log}\downarrow\ \uparrow{\scriptstyle 10^{-pH}} & & {\scriptstyle -\log}\downarrow\ \uparrow{\scriptstyle 10^{-pOH}} & & {\scriptstyle -\log}\downarrow\ \uparrow{\scriptstyle 10^{-14}} \\
pH & + & pOH & = & 14
\end{array}
$$

Ex

What is the pH of a solution with a $[H^+] = 1.62\times10^{-3}$ mols/L ?

$$pH = -\log[H^+] = -\log \overset{3}{1.62}\times10^{-3}$$

$$\sim 3$$

$$= 2.\underbrace{790|484\,985}_{\text{sig figs}}$$

mantissa (exponent on 10)

$$pH = 2.\underbrace{790}_{\text{3 sig figs}}$$

What is the $[H^+]$ for a solution with a pH of 2.691

$$[H^+] = 10^{-pH} = 10^{-2.691}$$

$$= 0.002037042 \text{ mols/L}$$

$$\boxed{= 0.00204 \text{ mols/L} \quad \text{or} \quad 2.04 \times 10^{-3} \text{ mols/L}}$$

What is the pH of a solution with a pOH of 7.923?

$$pH + pOH = 14$$

$$pH = 14 - pOH$$

$$pH = 14 - 7.923 = 6.077$$

$$\boxed{pH = 6.077}$$

what is the pH of a solution with a $[OH^-] = \underbrace{1.792}_{4\ SF} \times 10^{-10}$ mols/L

$$
\begin{tikzcd}
{[H^+]} \arrow[d, shift right, "-\log"'] & \times & {[OH^-]} \arrow[d, shift right, "-\log"'] & = & 1\times10^{-14} \arrow[d, shift right, "-\log"'] \\
pH \arrow[u, shift right, "10^{-pH}"'] & + & pOH \arrow[u, shift right, "10^{-pOH}"'] & = & 14 \arrow[u, shift right, "10^{-14}"']
\end{tikzcd}
$$

$$[H^+][OH^-] = 1\times10^{-14}$$

$$[H^+] = \frac{1\times10^{-14}}{[OH^-]} = \frac{1\times10^{-14}}{\underbrace{1.792}_{4\ SF}\times10^{-10}}$$

$$[H^+] = \overset{4\ SF}{\cancel{6.702412\times10^{-5}}} \quad \text{Error}$$
$$5.580357\times10^{-5}$$

$$pH = -\log[H^+] = -\log \cancel{6.702412\times10^{-5}} \quad \text{Error}$$
$$5.580357\times10^{-5}$$
$$\sim 5$$

$$= \cancel{4.173768823} \quad \text{Error}$$
$$4.253338005$$

$$\boxed{pH = \cancel{4.1738} \quad 4.2533}$$

Error

Activity 22 - Acids and Bases Worksheet

Name____________________________

Section____________ Date_______

Pre-Lab Lecture Questions. *Answer these questions on a separate sheet using complete sentences.*

1. Why do a lemon, grapefruit and vinegar taste sour?
2. What is the acid listed on the label of a bottle of vinegar?
3. What do antacids do? What are some bases listed on the labels of antacids?
4. Why are some aspirin products buffered?

Try It at Home

1. Some natural pigments act as indicators by forming different colors at different hydronium ion concentrations. Prepare an indicator by boiling some red cabbage leaves in water for 5 minutes. Cool the purple solution. Place small amounts of a household product such as vinegar, lemon juice, antacids, cleaners, shampoos and detergents in containers. Add a teaspoon of cabbage juice to each and observe the color. A pink-orange color indicates a pH range of 1-4; a pink-lavender, 5-6; purple, 7; green 8-11 and yellow, 12-13. Classify the products as acidic, neutral or basic. Try other highly colored vegetables or fruits to determine their use as indicators.
2. Place some cabbage indicator in a solution of baking soda made by adding 1 teaspoon of baking soda to a half a glass of water. Carefully add small amounts of vinegar. How does the color change? How do you know that the vinegar (an acid) neutralizes the baking soda (a base)?

Key Words

Use *complete sentences* to describe the following terms:

1. Electrolyte

2. Acid

3. Base

4. pH

5. Neutralization

6. Buffer

Electrolytes - Key Concepts

- Solutions of electrolytes are conductors of electrical current because electrolytes produce ions in aqueous solutions
- Strong electrolytes ionize completely, whereas weak electrolytes ionize only partially. Indicate incomplete ionization using double arrows “ $\rightleftharpoons$ ”.

Exercise A

Write an equation for the dissolving of the following salts as they combine with water to form an aqueous solution:

1. LiCl

2. $Mg(NO_3)_2$

3. Na_3PO_4

4. K_2SO_4

5. $MgCl_2$

Exercise B

Indicate whether aqueous solutions of the following solutes will contain ions, molecules, or both ions and molecules, and write an equation for their dissolution.

6. Glucose, $C_6H_{12}O_6$, a nonelectrolyte **Ions, molecules, or both? (circle one)**

Equation describing the dissolving of the solute:

7. NaOH, a strong electrolyte **Ions, molecules, or both? (circle one)**

Equation describing the dissolving of the solute:

8. K_2SO_4, a strong electrolyte **Ions, molecules, or both? (circle one)**

Equation describing the dissolving of the solute:

9. NH_3, a weak electrolyte that is also a base: **Ions, molecules, or both? (circle one)**

Equation describing the dissolving of the solute:

Acids and Bases – Key Concepts

- In water, an Arrhenius acid produces H^+, and an Arrhenius base produces OH^-.
- According to the Brønsted-Lowry theory, acids are proton (H^+) donors, and bases are proton acceptors.
- Protons form hydronium ions (H_3O^+) in water when they bond to water molecules.

Exercise C

Indicate whether the following characteristics describe and acid (A) or a base (B):

1. A B Turns blue litmus red
2. A B Contains more OH^- ions than H_3O^+ ions
3. A B Tastes bitter
4. A B Contains more H_3O^+ ions than OH^- ions
5. A B Tastes sour
6. A B Neutralizes bases
7. A B Turns red litmus blue
8. A B Neutralizes acids

Exercise D

Fill in the blank spaces with the formula or name of an acid or base

	Formula	Name
1.	HCl	
2.		Sodium hydroxide
3.		Sulfuric acid
4.		Nitric acid
5.	$Ca(OH)_2$	
6.	H_2CO_3	
7.	$Al(OH)_3$	
8.		Potassium hydroxide

Strengths of Acids and Bases – Key Concepts

- In strong acids, all the H^+ in the acid is donated to H_2O; in a weak acid, only a small percentage of acid molecules produce H_3O^+.
- Strong bases are hydroxides of Group 1A and 2A elements that ionize completely in water. An important weak base is ammonia, NH_3.

Exercise E

When do you use the double arrows in ionization equations? Write equations for the ionization of the following acids in water:

1. HCl, a strong acid

2. HF, a weak acid

3. HNO_3, a strong acid

Exercise F

When is water a reactant in the dissolving process? Write equations for the ionization of the following bases in water:

4. NaOH, a strong base

5. NH_3, a weak base

6. $Mg(OH)_2$, a strong base

Acid-Base Neutralization – Key Concepts

- Neutralization is a reaction of an acid and a base that produces water and a salt.
- The net ionic equation for the neutralization of any strong acid with any strong base is

$$H_3O^+ + OH^- \longrightarrow 2\,H_2O$$

- The net ionic equation for the neutralization of a weak acid by a strong base must include the acid written as a molecule:

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

- In a balanced neutralization equation, the number of moles of OH^- utilized must equal the number of moles of protons available for reaction. In other words, one mole of a diprotic acid or a triprotic acid requires 2 or 3 moles of NaOH, respectively, to become neutralized. All of the protons must be converted to water.

Exercise G

Write neutralization equations for the reactions between the following acids and bases:

1. Hydrochloric acid and magnesium hydroxide

2. Sulfuric acid and sodium hydroxide

3. Nitric acid and potassium hydroxide

4. Phosphoric acid and sodium hydroxide

5. Sulfuric acid and ammonia

Ion Product of Water – Key Concepts

- In pure water, a small fraction of the water molecules transfer protons to each other, producing small, but equal amounts of H_3O^+ and OH^-. Both ions have a concentration of 1×10^{-7} M at room temperature.

- The ion product for water is denoted as K_w where $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25 °C. This equilibrium applies to *all* aqueous solutions, not only to pure water.

- In acidic solutions, $[H_3O^+]$ is greater than $[OH^-]$. In basic solutions, $[OH^-]$ is greater than $[H_3O^+]$. In neutral solutions, $[H_3O^+]$ is equal to $[OH^-]$. However K_w always holds as $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25 °C.

Example:

What is the $[H_3O^+]$ in a solution that has $[OH^-] = 1.0 \times 10^{-9}$?

Solution:

$$K_w = 1.0 \times 10^{-14} M^2$$

$$K_w = [H_3O^+][OH^-]$$

$$[H_3O^+] = \frac{K_w}{[OH^-]}$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14} M^2}{[OH^-]}$$

$$[H_3O^+] = \frac{1.0 \times 10^{-14} M^2}{1.0 \times 10^{-9} M} = 1.0 \times 10^{-5} M$$

Exercise H

Calculate $[H_3O^+]$ when the $[OH^-]$ has the following values:

1. 1.0×10^{-10}

2. 1.0×10^{-5}

3. 1.0×10^{-7}

4. 1.2×10^{-4}

5. 3.5×10^{-8}

Exercise I

Calculate $[OH^-]$ when the $[H_3O^+]$ has the following values:

1. 1.0×10^{-3}

2. 1.0×10^{-10}

3. 1.0×10^{-6}

4. 2.8×10^{-13}

5. 8.6×10^{-7}

The pH Scale – Key Concepts

- For commonly encountered solutions, the pH scale ranges from 0 to 14. Its value is related to the hydronium ion concentration of the solution.

- A neutral solution has a pH of 7.00 at 25 °C. In acidic solutions, the pH is less than 7; in basic solutions, it is above 7.

- The mathematical definition is $pH = -\log [H_3O^+]$

- The number of decimal places in a pH value is the same as the number of sig. figs. in the concentration used to calculate that pH value.

Exercise J

Circle the most acidic pH in each group:

1. 5   2

2. 12   9   2

4. 3   7   10

5. 7.5   4.4   3.2

3. 0.2 1.5 2.3

6. 5.5 3.8 11.2 1.6

Exercise K

Calculate the pH of the following solutions at 25 °C. Indicate whether the solution is acidic, basic, or neutral.

	pH	Acidic, Basic, or Neutral
1. $[H_3O^+] = 1.0 \times 10^{-8}$ *M*	______________	A B N
2. $[H_3O^+] = 0.0010$ *M*	______________	A B N
3. $[OH^-] = 1.0 \times 10^{-12}$ *M*	______________	A B N
4. $[OH^-] = 2.0 \times 10^{-5}$ *M*	______________	A B N
5. $[OH^-] = 1.0 \times 10^{-7}$ *M*	______________	A B N

Exercise L

Indicate whether the following pH values are acidic, basic, or neutral at 25 °C:

1. A B N plasma, pH = 7.40
2. A B N soft drink, pH = 2.80
3. A B N maple syrup, pH = 6.80
4. A B N beans, pH = 5.00
5. A B N tomatoes, pH = 4.20
6. A B N lemon juice, pH = 2.20
7. A B N saliva, pH = 7.00
8. A B N eggs, pH = 7.80
9. A B N lime (CaO, not citrus), pH = 12.40
10. A B N strawberries, pH = 3.00

Exercise M

Complete the following table for solutions at 25ºC:

	$[H_3O^+]$	$[OH^-]$	pH	acidic, basic, neutral
1.		1.0×10^{-12}		
2.			8.32	
3.	5.0×10^{-8}			
4.				neutral
5.			1.00	

Buffers – Key Concepts

- A buffer solution minimizes the change in pH when small amounts of acid or base are added.
- Virtually all buffers contain a weak acid and its conjugate base. The weak acid reacts with added OH^- and the conjugate base (which is also weak) reacts with added H_3O^+.
- Buffers are important in maintaining the pH of blood.

Exercise N

State whether or not mixtures 1 through 4 below represent a buffer system, and explain why or why not:

1. HCl and NaCl

2. K_2SO_4

3. H_2CO_3

4. H_2CO_3 and $NaHCO_3$

5. A buffer is prepared by adding 26.8 mL of 0.200 M HCl to 50.0mL of 0.200 M tris-(hydroxymethyl)aminomethane (Tris). The mixture is then diluted to a total volume of 200 mL with water. $TrisH^+$ has a pK_a value of 8.3 at 20^oC. What is the pH of the above buffer solution? (You may use the Henderson-Hasselbalch equation (see below) in your calculation.) pH = pKa + log([Tris]/[$TrisH^+$])

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