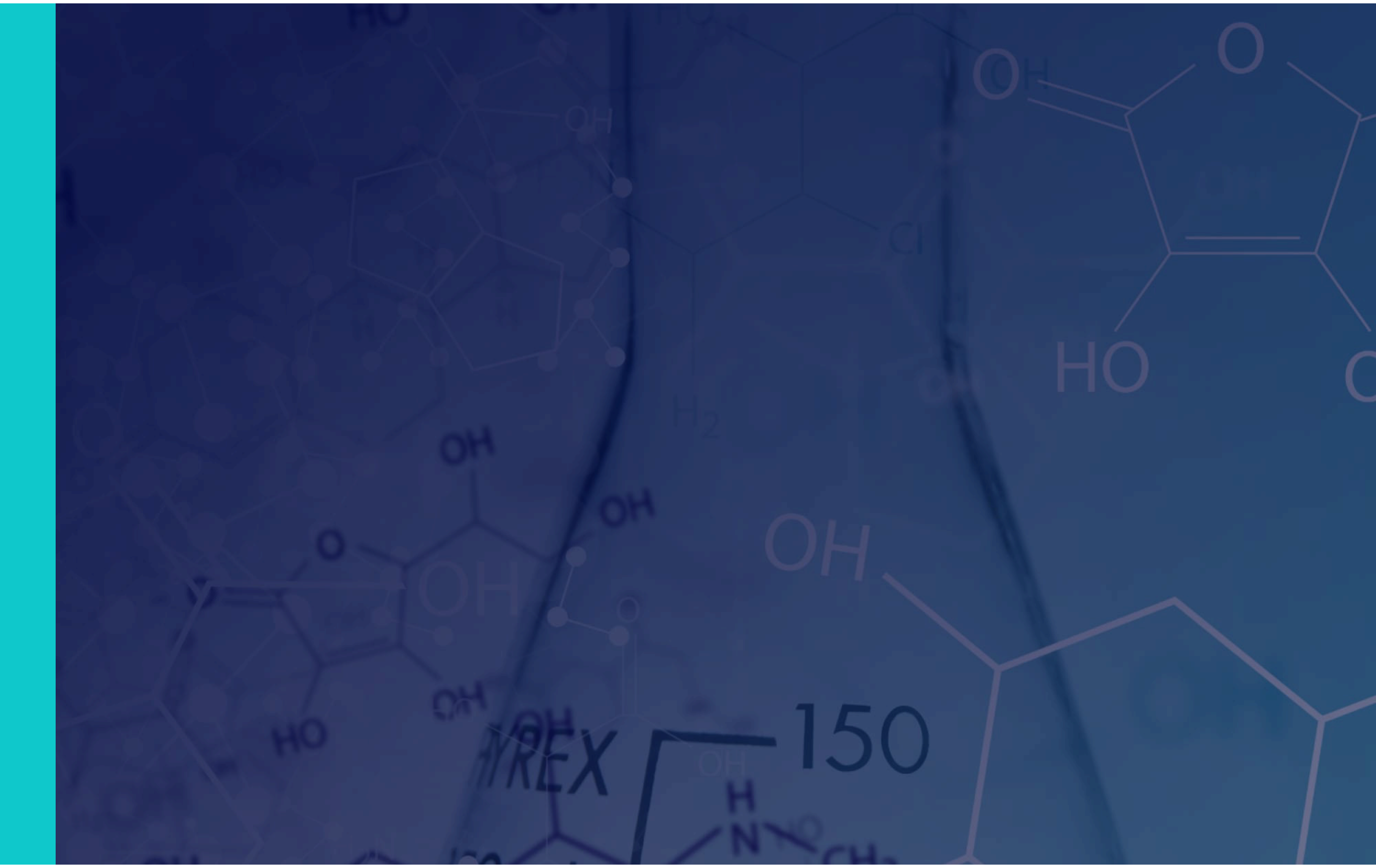


UWO

CHEM 1302

Winter 2025, Chapter 9 Notes

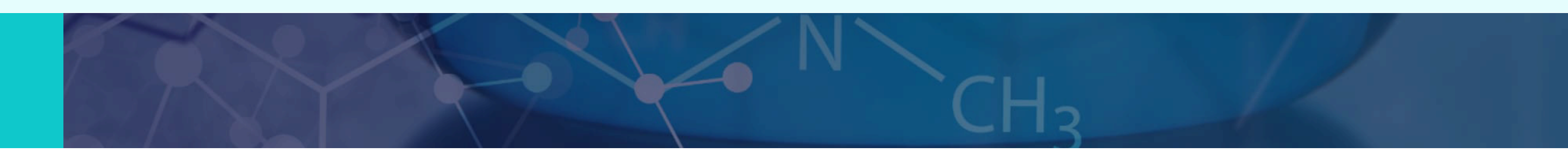


Table of Contents

9. Weak Acids and Bases

9.1

Introduction to Acids and Bases

9.1 Definitions of Acids and Bases

9.1.1

Definitions of Acids and Bases

Arrhenius definition: compounds identified as **producing H^+** or **OH^- ions:**

- **Acids** produce H_3O^+ aka hydronium ions:

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

- **Bases** produce OH^- aka hydroxide ions:

$$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$$

Bronsted - Lowry definition: compounds identified as **H^+ acceptors or donors**

- **Acids** donate H^+ ions

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

- **Bases** accept H^+ ions

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

Lewis definition: compounds identified as **electron acceptors or electron donors**

- **Acids:** electron acceptors (anions/cations: cations), and we call them (nucleophiles/electrophiles): electrophiles
- **Bases:** electron donors and we call them nucleophiles/electrophiles: nucleophiles
  Examples: ligands, --:NH_2

WIZE CONCEPT

Recall: Together, a Lewis acid and base can form a specific type of bond (when the Lewis base donates a pair of electrons to the Lewis acid to form a bond): coordinate covalent bond

Example: In the following reaction between NH_3 and BH_3, which is acting as the Lewis acid and which is acting as the Lewis base?

NH_3 is donate its pair of electron to BF .
NH_3 is acting as the Lewis Base(donate a pair of electrons)
BF_3 is acting as the Lewis Acid(accepting a pair of electrons)

WIZE TIP

If we remember the definitions of acids and bases in alphabetical order, we can remember that the definitions go from simplest to most complex:

Arrhenius definition-acids and bases produce H+ or OH- ions

Bronsted-Lowry definition-acids and bases are H+ donors or acceptors

Lewis definition-acids and bases are electron pair acceptors or donators.

WIZE TIP

Think of the acid HCl to remember the Bronsted-Lowry definition of an acid. Acids are proton (H+) donors.

Then, to remember the Lewis acid definition, remember that since acids donate a proton they do they opposite for electrons...they accept electrons!

Bases on the other hand are proton acceptors and since they accept protons, they do the opposite for electrons...they donate them!

9.2 Conjugate Acid-Base Pairs

9.2.1

Identifying Conjugate Acid and Base Pairs

Two substances that **differ from each other only by one proton (H+)** are referred to as a **conjugate acid-base pair**.

Identify the acid, base, conjugate acid, and conjugate base in the following reactions:

1) $HNO_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NO_2^-(aq)$

Acid Base Conj. Acid Conj. Base

Write the Ka or Kb expression for this reaction:

- If we were to increase Ka, would that lead to more/less dissociation into ions? More
- Would this mean we have a stronger/weaker acid? Stronger

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

Write the Ka or Kb expression for this reaction:

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

- If we were to increase Kb, would that lead to more/less OH- ions? more
- Would we have a stronger/weaker base? stronger

9.2.2

Practice: Conjugate Acid-Base Pairs

In the reaction HSO_4^-(aq) + OH^-(aq) ⇌ SO_4^{2-}(aq) + H_2O(l), the conjugate acid-base pairs are:

- HSO_4^- and OH^- ; SO_4^{2-} and H_3O^+
- HSO_4^- and SO_4^{2-} ; H_2O and OH^- ✓
- HSO_4^- and OH^-; SO_4^{2-} and H_2O.
- HSO_4^- and H_2O ; OH and SO_4^{2-}
- HSO_4^- and H_3O^+ ; SO_4^{2-} and OH^-

9.3 Acid Equations (Ka, pKa)

9.3.1

Acids and their Equations

Strong Acids

Strong acids **dissociate completely**

Example:

$HCl_{(aq)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$

WIZE TIP

The following are strong acids that you should memorize!

- **HCl (hydrochloric acid)**
- **H_2SO_4 (sulfuric acid)**
- **HNO_3 (nitric acid)**
- **$HClO_4$ (perchloric acid)**
- **HBr (hydrobromic acid)**
- **HI (hydroiodic acid)**

Weak Acids

Weak acids dissociate incompletely, so they have a K_a value (acid dissociation constant)

Ka is defined as the equilibrium constant for the reaction below:

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

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$pK_a = -\log K_a \; and \; K_a = 10^{-pK_a}$$

Acid strength follows the following trends:

- A weaker acid would have a higher/lower Ka value? lower
- A weaker acid would have a higher/lower pKa value? higher

WIZE TIP

Ka and pKa are like opposite trends.
The higher Ka goes, the lower pKa will go,
and the lower Ka goes, the higher pKa will go!

- A weaker acid will have a higher/lower pH value: higher
- A stronger acid will have a higher/lower Ka value? higher
- A stronger acid will have a higher/lower pKa value? Lower
- A stronger acid will have a higher/lower pH value? Lower

In general:

- If the $K_a > 1$, $pK_a < 0$ the acid is strong/weak: strong and reactants/products: products are favored
- If the $K_a < 1$, $pK_a > 0$ the acid is strong/weak: weak and reactants/products: reactants are favored

Practice: Strong Acids

Which one of the following statements about strong acids is true?

- All strong acids have H atoms bonded to electronegative oxygen atoms. X
- Strong acids are 100% ionized in water. ✓
- The conjugate base of a strong acid is itself a strong base. X strong acid → weak conj. base
- Strong acids are very concentrated acids. X
- Strong acids produce solutions with a higher pH than weak acids. X

Acids With Different Numbers of Protons

Monoprotic Acids

- These are acids that only have **1 acidic proton (H+)** that they can lose in a reaction
- *think: mono means 1

  Examples: HCl, HBr, HI, HNO_3, $HClO_4$, etc

Diprotic Acids

- These are acids that have **2 acidic protons**
- *think: di means 2

  Example: H_2SO_4

Triprotic Acids

- These are acids that have **3 acidic protons**
- *think: tri means 3

  Example: H_3PO_4

Polyprotic Acids

- This term is just referring to any **acid that has more than 1 proton**
- *think: poly means many
  Example: H_2SO_4, H_3BO_3, $H_2C_2O_4$

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

$$H_2PO_4^-\ (aq)\ +\ H_2O\ (l)\ \leftrightharpoons\ HPO_4^{2-}\ (aq)\ +\ H_3O^+\ (aq)$$

$$HPO_4^{2-}\ (aq)\ +\ H_2O\ (l)\ \leftrightharpoons\ PO_4^{3-}\ (aq)\ +\ H_3O^+\ (aq)$$

Polyprotic Acids

Name	Formula	K_{a1}	K_{a2}	K_{a3}
Phosphoric acid	H_3PO_4	7.6×10^{-3}	6.3×10^{-8}	4.2×10^{-13}
Sulfuric acid	H_2SO_4	VERY LARGE	1.2×10^{-2}	
Sulfurous acid	H_2SO_3	1.7×10^{-2}	6.5×10^{-8}	
Carbonic acid	HOCOOH	4.3×10^{-7}	5.6×10^{-11}	
Oxalic acid	HOOCCOOH	5.9×10^{-2}	6.5×10^{-5}	
Citric acid	$C_3H_5O(COOH)_3$	7.4×10^{-4}	1.7×10^{-5}	4.0×10^{-7}

WIZE CONCEPT

The main concept to take away from this table is just that each ionization reaction has its own Ka value and **Ka_1>Ka_2>Ka_3.**

9.4 Base Equations (Kb, pKb)

9.4.1

Bases and their Equations

Strong Bases

Strong bases **dissociate completely**

Example:

$NaOH_{(aq)} \rightarrow Na^{+}_{(aq)} + OH^{-}_{(aq)}$

WIZE TIP

The following are strong bases that you should memorize!

- **All Group 1 Hydroxides (ex. NaOH, KOH, RbOH, CsOH)**
- **3 Group 2 Hydroxides ($Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$)**
- **All Group 1 Oxides (ex. Li_2O, Na_2O, K_2O)**
- **Metal amides (ex. $M\text{-}NH_2$ where M is a metal)**

Weak Bases

Weak bases **dissociate incompletely** so they have a **K_b value** (base dissociation constant)

K_b is defined as the equilibrium constant for the reaction below:

$$B + H_2O \rightleftharpoons BH^+ + OH^-$$

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

$$pK_b = -\log K_b \ and \ K_b = 10^{-pK_b}$$

WIZE CONCEPT

Weak acids and weak bases both reach an equilibrium between the undissociated and dissociated ions.

Base strength follows the following trends:

- A weaker base will have a higher/lower Kb value? lower
- A weaker base will have a higher/lower pKb value? higher
- A weaker base will have a higher/lower pH? lower
- A stronger base will have a higher/lower Kb value? higher
- A stronger base will have a higher/lower pKb value? lower
- A stronger base will have a higher/lower pH value? higher

WIZE TIP

If you see an acid or base that is not listed above (above we listed the strong acids and bases), assume you are dealing with a weak acid/base!

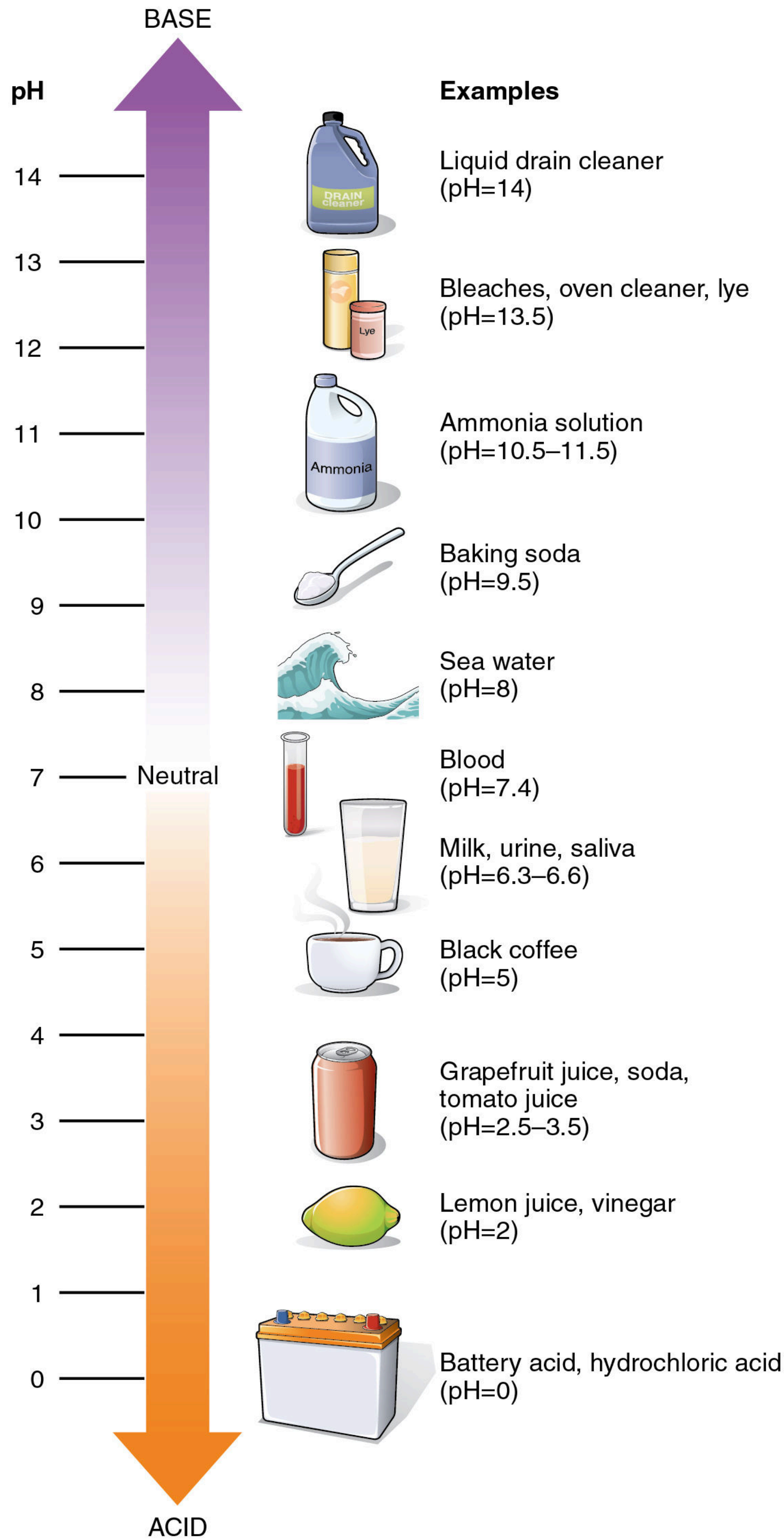
9.5 pH

9.5.1

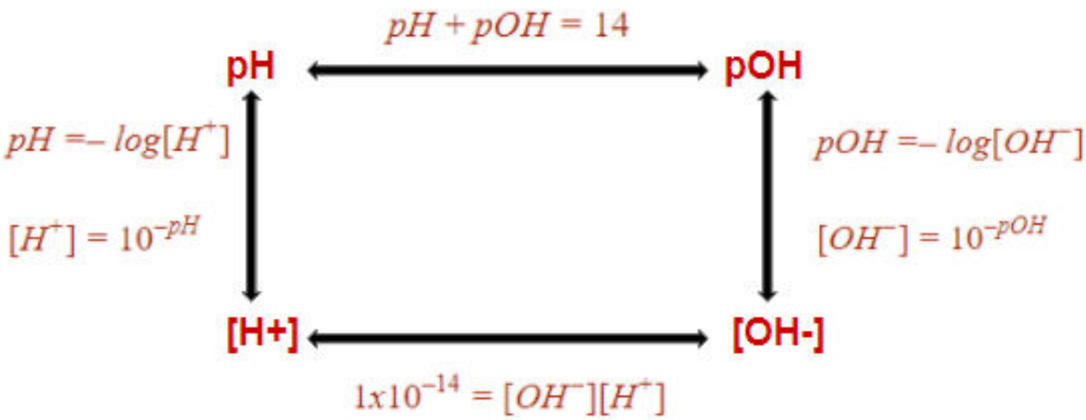
pH Scale

The **pH scale** is a logarithmic scale used to measure the **concentration of [H+]** in a solution, since this value can vary over many orders of magnitude. It ranges from 0 to 14, with 7 representing neutral pH

Note: There are some exceptions to this. For example, an acid could have a negative pH! These aren't commonly seen.


BASE
pH
Examples
14
Liquid drain cleaner
(pH=14)
DRAIN cleaner
13
Bleaches, oven cleaner, lye
(pH=13.5)
Lye
12
11
Ammonia solution
(pH=10.5–11.5)
Ammonia
10
Baking soda
(pH=9.5)
9
Sea water
(pH=8)
8
7
Neutral
Blood
(pH=7.4)
6
Milk, urine, saliva
(pH=6.3–6.6)
5
Black coffee
(pH=5)
4
Grapefruit juice, soda,
tomato juice
(pH=2.5–3.5)
3
2
Lemon juice, vinegar
(pH=2)
1
0
Battery acid, hydrochloric acid
(pH=0)
ACID

A lower pH indicates that a substance is more (acidic/basic) acidic and has a (higher/lower) higher [H+]



We also have Ka and pKa equations:

$$pK_a = -\log K_a \; and \; K_a = 10^{-pK_a}$$

As well as Kb and pKb equations:

$$pK_b = -\log K_b \; and \; K_b = 10^{-pK_b}$$

What does it mean to be a logarithmic scale?

How does the [H+] change as the pH goes from 1 to 2 or 1 to 3??

pH 1

$\downarrow$

pH 2

$\downarrow$

pH 3

WIZE CONCEPT

Since pH is a logarithmic scale, each 1 unit change in pH corresponds to a 10-fold change in [H+]!

9.5.2

Example: pH Problem

If 0.2 mols of $\mathrm{HCl}_{(s)}$ are dissolved in 1.8 L of water, what is the pH?

$$HCl \rightarrow H^+ + Cl^-$$

$$[H^+] = \frac{0.2\,mol}{1.8\,L} = 0.11\,mol/L$$

$$pH = -\log[H^+] = 0.95$$

Example: pH Problem

$$C_6H_5COOH \rightleftharpoons H^+ + C_6H_5COO^-$$

Ka of $C_6H_5COOH = 6.5 \times 10^{-7}$ and the concentration of C_6H_5COOH is 0.01 M, what is the pH of the solution?

	C_6H_5COOH	H	C_6H_5COO
I	0.01	0	0
C	$-x$	$+x$	$+x$
E	$0.01-x$	x	x

$$Ka = \frac{x^2}{(0.01-x)}$$

$$x^2 = 0.01\,Ka$$

$$x^2 = 6.5 \times 10^{-9}$$

$$x = 8.06 \times 10^{-5}$$

$$pH = -\log(8.06 \times 10^{-5})$$
$$= 4.09$$

9.5.4

Practice: Solve for [H+]

A solution of HBr has a pOH of 12.2, what is the $[H^+]$ concentration?

- 2.4×10^{-3} M
- 6.3×10^{-13} M
- 12.2 M
- 1.6×10^{-2} M
- We can not find the $[H^+]$ from the pOH

$$pH + pOH = 14$$

$$pH = 14 - 12.2 = 1.8$$

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

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

$$= 10^{-1.8} \approx 1.58 \times 10^{-2}\ M$$

9.6 Percent Dissociation

9.6.1

Percent Ionization or Dissociation

$$p = \frac{(concentration\ of\ acid\ ionized)}{(concentration\ of\ acid\ solute)} x100\%$$

p=percent ionization

For the general weak acid equation:

HA(aq) ⇌ H^{+}(aq) + A^{-}(aq)

$$\boxed{p = \frac{[H+]}{[HA]} x100}$$

- Do you think a weak acid would have a high or low percent ionization? ________
- Do you think a strong acid would have a high or low percent ionization? ________

9.6.2

Example: Percent Dissociation

A 0.150 M solution of acetic acid (shorthand formula = HAc) is found to be 1.086% dissociated. What is the Ka?

9.7 Salts

9.7.1

Acidic, Basic, and Neutral Salts

Salts are ionic compounds that dissociate (incompletely/completely) ____________ in water!

Salts are formed in **neutralization reactions** (reaction where an acid and base react together)

Example:

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

These reactions are exothermic (and produce a salt and water!)

The **resulting pH of a solution** will depend on whether an **acidic, basic, or neutral salt is made**

When deciding if a salt is acidic, basic, or neutral, consider the ions that make up the salt:

If we have the conjugate acid of a strong base, would this be a very weak or decent acid?

- It would be very ________
- So ________ that we can ignore it!!

If we had the conjugate base of a strong acid, would this be a very weak or decent base?

- Very ________
- So ________ that we can ignore it!

If we had the conjugate acid of a weak base, would that be a very weak or decent acid?

- Would be a ________ acid

If we had the conjugate base of a weak acid, would that be a very weak or decent base?

- Would be a ________ base

Examples:

1) CH_3COONa

- **Consider the anion** we get from this salt: __________
    - This is a conjugate _______ (acid/base) of a _______ (weak/strong) _______ (acid/base) which is: __________
    - Does this contribute to pH? If so, how?
        - ________________
- **Consider the cation** we get from this salt: ______
    - This is a conjugate ______ of a _______ ______ (NaOH)
    - Does this contribute to pH? If so, how?
        - _____________

Therefore, this is a(n) _______ salt since the resulting pH is _______.

WIZE TIP

Being asked if a salt is acidic, basic, or neutral is commonly seen on exams. :)

2) $NaNO_3$

- **Consider the cation** we get from this salt: ________
    - We already found this does not contribute to pH

- **Consider the anion** we get from this salt: ________
    - This is a conjugate ________ of a ________ ________ which is: ________
    - Does this contribute to pH? If so, how?
        - ________________

Therefore, this is a(n) __________ salt since the resulting pH is __________.

3) NH_4Cl

- **Consider the anion** we get from this salt: ______
  - This is a conjugate ________ (acid/base) of a __________ (weak/strong) ________ (acid/base) which is: _______
  - Does it contribute to pH? If so, how?
    - ________________
- **Consider the cation** we get from this salt: _________
  - This is a conjugate ________ (acid/base) of a _________ (weak/strong) ________ (acid/base) which is: ________
  - Does it contribute to pH? If so, how?
    - ______________________

Therefore, this is a(n) _________ salt since the resulting pH is _________.

Salts (Cntd.)

Another way of being presented this info is through equations:

1) Weak acid with a strong base:

$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

Could be asked about the resulting pH, we already found that it would be ________.

2) Strong acid with a strong base:

$HNO_3 + NaOH \rightarrow H_2O + NaNO_3$

We found that the resulting pH would be ________.

3) Strong acid with a weak base:

$HCl + NH_3 \rightarrow NH_4Cl + H_2O$

We found that the resulting pH would be ________.

4) There is one last scenario where we could have a weak acid with a weak base

To determine whether the resulting pH is acidic, neutral, or basic, we need to figure out if the weak acid or weak base is stronger!

- If the Ka **of the weak acid** > Kb **of the weak base**, the resulting pH is: __________
    - pH would be (</=/>) _____ 7

- If the Kb **of the weak base** > Ka **of the weak acid**, the resulting pH is: __________
    - pH would be (</=/>) _____ 7

- Finally, if the Ka **of the weak acid** = Kb **of the weak base**, the resulting pH is: ___________

Example: Acidic, Basic, or Neutral Solution

Will a solution of ammonium nitrite (NH_4NO_2) be acidic, basic or neutral?

$K_a = 5.6 \times 10^{-10}$ for NH_4^+

$Kb = 1.7 \times 10^{-11}$ for NO_2^-

9.7.4

Example: Salts

Given the following reactions, is $NaHCO_3$ an acidic, basic, or neutral salt?

$$H_2CO_3 \rightleftharpoons H^+ + HCO_3^- \quad Ka=4.3x10^{-7}$$

$$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-} \quad Ka=4.3x10^{-10}$$

From our salt, the **cation** is: ________ and the **anion** is: ________

- The **cation** is a conjugate (acid/base): ________ of a strong/weak: ________ acid/base: ________
  - Does it contribute to pH? ______

Look at the above equations for HCO_3-.

In equation 1, HCO_3- is acting as an (acid/base): ________

In equation 2, HCO_3- is acting as an (acid/base): ________

To figure out whether HCO_3- will cause the pH to be more acidic or basic, we will need to **compare the Ka value for HCO_3- (when it acts as an acid) to the Kb value for HCO_3- (when it acts as a base)**

The larger value (Ka or Kb) will help us know how the pH would be affected!

$$H_2CO_3 \rightleftharpoons H^+ + HCO_3^- \quad Ka=4.3x10^{-7}$$

$$HCO_3^- \rightleftharpoons H^+ + CO_3^{2-} \quad Ka=4.3x10^{-10}$$

Ka of HCO_3- = ____________

Kb=?

- Which is larger, Ka or Kb? ______
- How would HCO_3- affect the pH? (lower/neutral/raise): ________ pH
- Therefore the salt is (acidic/neutral/basic): ________

Practice: Salts

Is an aqueous solution of each of these salts acidic, basic or neutral?

Part 1

Is KBr an acidic, basic, or neutral salt?

- Acidic salt
- Basic salt
- Neutral salt
- There is not enough information provided to determine this

Practice: Salts

Is an aqueous solution of each of these salts acidic, basic or neutral?

Part 2

Is NH_4Cl an acidic, basic, or neutral salt?

- Acidic salt
- Basic salt
- Neutral salt
- There is not enough information provided to determine this

Practice: Salts

Is an aqueous solution of each of these salts acidic, basic or neutral?

Part 3

Is KCN an acidic. basic, or neutral salt?

- Acidic salt
- Basic salt
- Neutral salt
- There is not enough information provided to determine this

Salts Cheatsheet

Some salts that dissociate completely in water exhibit aicd/base properties. In order to determine if the solution will be acidic or basic, first separate the substance into it's positive and negative ions.

If neither the cation nor the anion can affect the pH, the solution should be neutral

$$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$$

If only the cation of the salt is acidic, the solution will be acidic (NH_4Cl)

$$NH_4Cl(s) \rightleftharpoons NH_4^{+}(aq) + Cl^{-}(aq)$$

If only the anion of the salt is basic, the solution will be basic (NaCN)

$$NaCN(s) \rightleftharpoons Na^{+}(aq) + CN^{-}(aq)$$

If a salt has a cation that is acidic and an anion that is basic:

The pH of the solution is determined by the relative strengths of the acid and the base, based on the K_a and K_b of the ions.

If $K_a > K_b$, it will be acidic.

If $K_b > K_a$, it will be basic.

9.8 Applications of Acid and Base Equations

9.8.1

Example: Ka and pKa Values

What is the strongest acid below? What is the weakest acid?

Strongest acid: ______________

Weakest acid: ________

<u>Standard K_a values</u>

Formula	K_a	pK_a
HCN	4.0×10^{-10}	9.40
HClO	3.5×10^{-8}	7.46
CH_3COOH	1.8×10^{-5}	4.74
Benzoic acid	6.3×10^{-5}	4.20
HCOOH	1.9×10^{-4}	3.72
HNO_2	4.5×10^{-4}	3.35
HF	6.7×10^{-4}	3.17
Cl_3CCOOH	2.0×10^{-1}	0.70

What is the strongest base below? What is the weakest base?

Strongest base: ________________

Weakest base: __________

Standard K_b values

Formula	K_b	pK_b
Aniline	7.4×10^{-10}	9.13
Pyridine	1.5×10^{-9}	8.82
Ammonia	1.8×10^{-5}	4.74
Trimethylamine	7.4×10^{-5}	4.13
Ethylamine	4.3×10^{-4}	3.37
Methylamine	6.4×10^{-4}	3.19
Dimethylamine	7.4×10^{-4}	3.13

9.8.2

Example: pH of Acids and Bases

Part 1: What is the pH of a 0.5 M HCl solution?

HCl is a (strong/weak acid) ________

It dissociates (completely/incompletely) ____________

Part 2: What is the pH of a 1.0 M CH_3COOH solution? Ka = 1.8×10^{-5}

This is an (acid/base) ________

It is (weak/strong) ________

Does this type of acid dissociate into ions completely/incompletely? ____________

Part 3: What is the pH of a solution prepared by dissolving 0.10 mol of $Ba(OH)_2$ in 1.0 L of pure water?

Is this an acid or base? ________

Is this a strong one or weak one? __________

Complete/incomplete dissociation? ____________

Part 4: Many liquid household cleaners contain ammonia. The concentration of ammonia (NH_3) in these products is usually around 5.0 M. What is the pH of such a cleaning solution? (K_b NH_3 = 1.8×10^{-5})

Is this an acid or a base? ________

Is it strong or weak? ________

Complete or incomplete dissociation? ____________

9.8.3

Types of Acid or Base Reactions and pH Calculations Summary

1. Strong Acid Alone

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

$$[H^+_{(aq)}] = [Cl^-_{(aq)}] = [HCl]_{(initial)}$$

$$pH = -\log[H^+_{(aq)}]$$

2. Strong Base Alone

$$NaOH \rightarrow Na^+_{(aq)} + OH^-_{(aq)}$$

$$[OH^-_{(aq)}] = [Na^+_{(aq)}] = [NaOH]_{initial}$$

$$pOH = -\log[OH^-_{(aq)}]$$

$$pH = 14 - pOH$$

3. Weak Acid Alone

Example: 2.0 mols of weak acid, HA, with a Ka = 2 x 10^{-4} in 1.0 L of solution.
Step 1: Make an ICE table

$$HA_{(aq)} + H_2O_{(l)} \leftrightharpoons H_3O^+_{(aq)} + A^-_{(aq)}$$

I	$2.0M$	n/a	0	0
C	$-x$	n/a	$+x$	$+x$
E	$2.0-x$	n/a	x	x

Step 2: Set up K_a equation and solve for $[H_3O^+]$

$$Ka = \frac{[H_3O^+_{(aq)}][A^-_{(aq)}]}{[HA_{(aq)}]} = \frac{x^2}{2.0-x} \approx \frac{x^2}{2.0}$$

$$x = \sqrt{(2.0)K_a} = 0.02M = [H_3O^+_{(aq)}]$$

$$pH = -\log[H_3O^+_{(aq)}] = 1.70$$

4. Weak Base Alone

Example: 2.0 moles of weak base, B, with a Kb of 7 x 10^{-7} in 1.0 L of solution.
Step 1: Make an ICE table

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

I	$2.0M$	n/a	0	0
C	$-x$	n/a	$+x$	$+x$
E	$2.0-x$	n/a	x	x

Step 2: Set up Kb equation and solve for [OH-]

$$K_b = \frac{[OH^-_{(aq)}][BH^+_{(aq)}]}{[B_{(aq)}]} = \frac{x^2}{2.0 - x} \approx \frac{x^2}{2.0}$$

$$x = \sqrt{(2.0)K_b} = 0.00118M = [OH^-_{(aq)}]$$

$$pOH = -\log[OH^-_{(aq)}] = 2.93$$

$$pH = 14 - pOH = 11.07$$

When can we make a simplifying assumption about x?

In K= x^2/(y-x)
When **y/K > 400** you can simplify and ignore the "-x"

9.9 Autoionization of Water (Kw)

9.9.1

Autoionization of Water (Kw)

Two water molecules can react with themselves in an **autoionization** reaction as follows:

$$2H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$

$$\boxed{K_w = [H_3O^+_{(aq)}][OH^-_{(aq)}] = 1 \times 10^{-14}\ (25^oC)}$$

Here is the reaction in action!

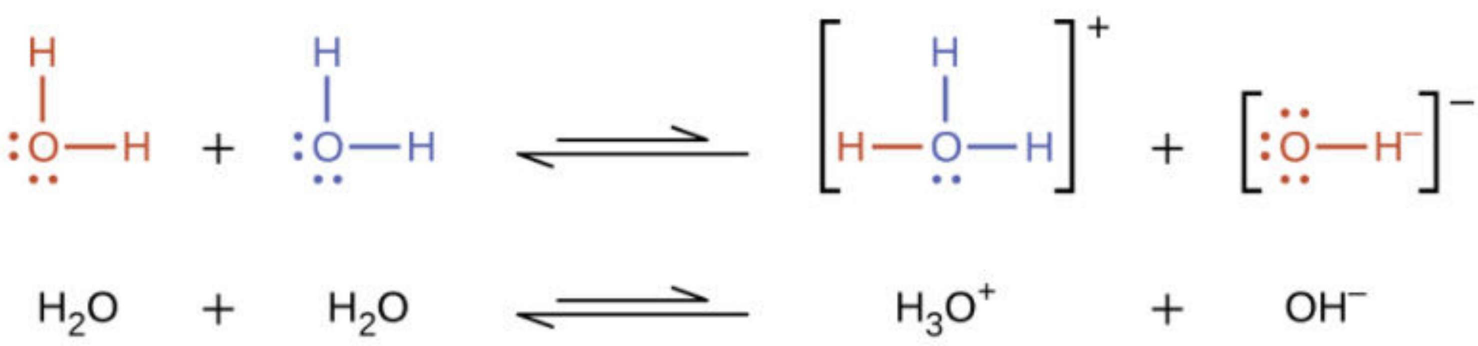


Photo by Rice University / CC BY

Label whether each reactant and product in the above reaction is acting as an acid/base or conjugate acid/conjugate base.

Water is **amphoteric:** it can act as an acid or a base!

Other Key Relationships

Kw = $[H^+][OH^-] = 1x10^{-14}$, $[H^+]$= $1x10^{-7}$M and $[OH^-]$= $1x10^{-7}$M at 25^{o}C

This is why the pH (and pOH) of neutral water is 7

$$\boxed{K_w = [H_3O^+_{(aq)}][OH^-_{(aq)}] = (Ka)(Kb) = 1 \times 10^{-14}\ (25^oC)}$$

*Ka and Kb have to be the K values for an acid and its conjugate base (or a base and its conjugate acid)

If [H+] goes higher, what do you think would happen to [OH-]? Goes higher/lower: ________

9.10 Acid Base Equations Summary

9.10.1

Important Relationships:

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

$$pOH = -\log OH^- and \; OH^- = 10^{-pOH}$$

$$2H_2O \rightleftharpoons H_3O^+ + OH^-$$

In pure water, at 25 °C:

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

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

$$pH + pOH = pK_w = 14.00 \; at \; 25^\circ C$$

K_a is the acid dissociation constant and is defined as the equilibrium constant for the reaction below:

$$HA(aq) + H_2O(l) \rightleftharpoons A^-(aq) + H_3O^+(aq)$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$pK_a = -\log K_a \; and \; K_a = 10^{-pK_a}$$

Strong acids dissociate completely.

$HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$ **(not an equilibrium, K_a is very large)**

Weak acids partially dissociate to give H^+ or H_3O^+

$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$ **(equilibrium, K_a)**

K_b is the base dissociation constant and is defined as the equilibrium constant for the reaction below:

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

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

$$pK_b = -\log K_b \; and \; K_b = 10^{-pK_b}$$

Strong bases dissociate completely.

$LiOH(s) \rightarrow Li^+(aq) + OH^-(aq)$ **(not an equilibrium, Kb is very large)**

Weak bases partially dissociate (they accept protons to give OH-, but the reaction is not complete)

$$CH_3NH_2(aq) + H_2O(l) \rightleftharpoons CH_3NH_3^+(aq) + OH^-(aq)$$

For a given conjugate acid-base pair:

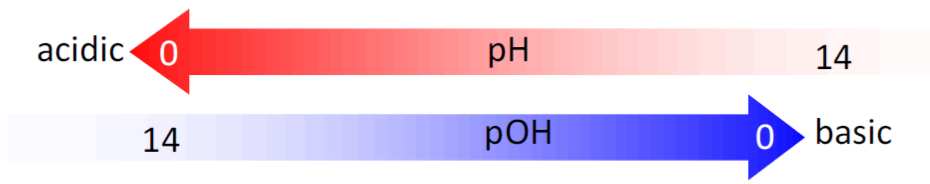
$$\begin{array}{ll} HA + H_2O \rightleftharpoons A^- + H_3O^+ & K_a \\ \underline{A^- + H_2O \rightleftharpoons HA + OH^-} & \underline{K_b} \\ 2H_2O \rightleftharpoons H_3O^+ + OH^- & K_w \end{array}$$

$$K_a \times K_b = K_w = 1 \times 10^{-14}\ and\ pK_a + pK_b = pK_w = 14.00\ (at\ 25^\circ C)$$

9.10.2

Acid/Base Important Relationships Cheatsheet

pH = -log[H_3O^+]	pOH = -log[OH^-]	pH + pOH = 14
[H_3O^+] = 10^{-pH}	[OH^-] = 10^{-pOH}	



Acidic solution	**Neutral solution**	**Basic solution**
[H^+] > [OH^-]	[H^+] = [OH^-]	[H^+] < [OH^-]
pH < 7 (at 25 °C)	pH = 7 (at 25 °C)	pH > 7 (at 25 °C)

	Strong acid	Weak acid	Strong base	weak base
pH (at 25 °C)	pH < 7		pH > 7	
pOH (at 25 °C)	pOH > 7		pOH < 7	
K_a or K_b	huge K_a	smaller K_a	huge K_b	smaller K_b
pK_a or pK_b	tiny pK_a	larger pK_a	tiny pK_b	larger pK_b
[H_3O^+] or [OH^-]	dissociate completely [H_3O^+] = [acid]	doesn't dissociate completely, need ICE table	dissociate completely [OH^-] = [base]	doesn't dissociate completely, need ICE table
conjugate acid/base	very weak conjugate base (~neutral)	stronger conjugate base	very weak conjugate acid (~neutral)	stronger conjugate acid