

Acid-Base Reactions

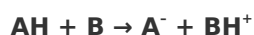
[See online here](#)

Acids and bases are groups of compounds that play major roles in different facets of society. They find applications in the food industry and pharmaceutical industries, as well as in human metabolism. This article tackles the basics of acids and bases. Topics discussed include the operational definitions of acids and bases; classification of acids and bases in relation to their strengths; different concepts in acid-base calculations like pH, pOH and equilibrium constants; and buffers and their applications.

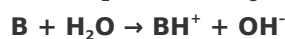
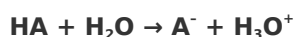


Acidity

The acidity of a solution is the concentration of H^+ (protons) [or H_3O^+] in water. An acid is a substance that can lose one more H^+ . A base is a substance that can pick up H^+ .



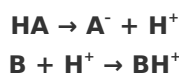
When an acid (AH) reacts with a base (B), H^+ is transferred. Water can behave as an acid or a base – it is amphoteric.



It can also ionize itself:



As most acid/base reactions take place in water, it is common not to include it in the equations:

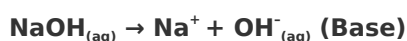
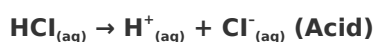


Acid-Base Definitions

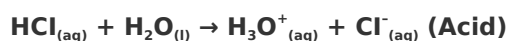
There are three operational definitions of acids and bases, developed because of the relative limitations of each one of them. These three definitions are the **Arrhenius definition**, the **Bronsted-Lowry definition**, and the **Lewis definition**.

Arrhenius definition

For **Arrhenius**, acids are compounds that dissociate in water to produce an H^+ , while bases are compounds that dissociate in water to produce OH^- ions. Examples of compounds that fall under the scope of the Arrhenius definition are the hydrohalic acids and the metal hydroxides. When these two groups of compounds dissociate in water they produce an H^+ or an OH^- ion, depending on whether the compound is a base or an acid.



Ammonia, NH_3 , is a basic compound but the Arrhenius definition fails to define it. NH_3 cannot dissociate in water to produce an OH^- because it does not contain an $-\text{OH}$ group in its structure. It can produce an OH^- by reacting itself with water. One of the failures of the Arrhenius definition is that it does not consider the solvent as a reactant.



Bronsted-Lowry definition

Bronsted-Lowry defines an acid as a proton donor, while a base as a proton acceptor. Water molecules, having the necessary requirements to donate and accept a proton, may now not only participate as a solvent but also as a reactant.

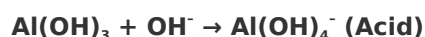
Lewis definition

The problem with the **Bronsted-Lowry definition** is that it is not able to explain the acidity of metal cations. Because of this, the **Lewis definition** was developed, which focused on the electrons of atoms involved in a reaction.

For Lewis, an acid is an electron pair acceptor while a base is an electron pair donor. The Lewis definition of acids correctly encompasses the acidity of cations. Since cations are positively charged, they have a higher tendency to accept an electron pair. The Lewis definition is considered the most general of the three as it is more inclusive.

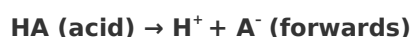
A good example of this is $\text{Al}(\text{OH})_3$ which is considered an amphoteric substance. Amphoteric substances are compounds that can act as either acids or bases. The

aluminum atom in the structure, having only six electrons, can still accommodate an electron pair so it can act as an acid. The oxygen atom has electron pairs to donate, and so it can act as a base in a reaction.



Equilibrium

Many reactions can go backward as well as forwards, e.g.:

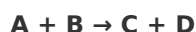


This is shortened:



After a certain time, the speed of the forwards' reaction will be the same as the backward reaction. The concentrations of the reactants and the products will remain stable - the system is said to be at **equilibrium**.

So for the reaction:



at equilibrium:

$$K = [\text{C}] [\text{D}] / [\text{A}] [\text{B}]$$

Note: Where [X] is the concentration of X, **K** is the **equilibrium constant**

If K is large the reaction will give mostly products with only small amounts of reactants remaining. If K is small there will only be small amounts of the products. In the case of acid dissociation in water:

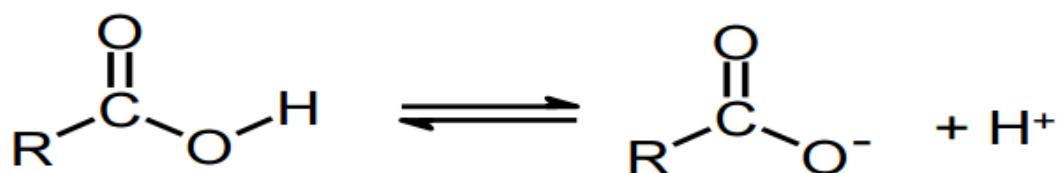


$$K_a = [\text{H}^+] [\text{A}^-] / [\text{HA}]$$

The acid dissociation constant is K_a

If K is large then HA will be almost completely dissociated giving a large $[\text{H}^+]$. Such acids are described as **strong acids**, e.g. H_2SO_4 , **HCl**, HNO_3 .

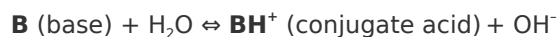
Acids which only dissociate to a small extent are **weak acids** (small K_a), e.g. Carboxyl acids:



e.g. ethanoic acid $\text{CH}_3\text{CO}_2\text{H}$, $K_a = 1.8 \times 10^{-5}$

Bases

Bases can accept protons:



As with acids, the strength of the base can be shown by the equilibrium constant (K_b) for the above reaction. This allows $\text{p}K_b$ to also be found. More commonly base strength is represented by $\text{p}K_a$ by considering the reaction of the conjugate acid.



$$\mathbf{p}K_a = -\log \frac{[\mathbf{B}][\mathbf{H}^+]}{[\mathbf{BH}^+]}$$

- weak bases have low $\text{p}K_a$ values.
- strong bases have high $\text{p}K_a$ values.

For any acid or base: $\mathbf{p}K_a + \mathbf{p}K_b = 14.0$

Strengths of Acids and Bases

Acids and bases can be classified as **strong** or **weak** depending on the degree of dissociation in water. Strong acids and bases completely dissociate in water, producing stoichiometric amounts of the H^+ and OH^- ions, respectively. On the other hand, weak acids and bases produce the same ions but do not completely dissociate in water. The following list provides an easy way of classifying compounds as strong or weak acids and bases.

Strong Acids

The hydrohalic acids HCl, HBr, and HI. Oxoacids in which the number of O atoms exceeds the number of ionizable protons by two or more, such as H_2SO_4 , HNO_3 , and HClO_4 .

Weak Acids

- The hydrohalic acid HF
- Acids in which H does not bond to O or a halogen, such as HCN and H_2S .
- Oxoacids in which the number of O atoms equals or exceeds by one the number of ionizable protons, such as HClO, HNO_2 , and H_3PO_4 .
- Carboxylic acids

Strong Bases

Water-soluble compounds containing O^{2-} or OH^- ions are strong bases. The cations are usually those of the most active metals:

- M_2O or MOH , where M = Group 1A metal (Li, K, Na, Rb, Cs)
- MO or $\text{M}(\text{OH})_2$, where M = Group 2A metal (Ca, Sr, Ba)

Weak Bases

Many compounds with an electron-rich nitrogen atom are weak bases (none are Arrhenius bases). The common structural feature is an N-atom with a lone electron pair:

- Ammonia

- Amines (general formula RNH_2 , R_2NH , or R_3N)

Acids in Water

When an acid (or a base) is added to water $[\text{H}^+]$ will change. The acidity of a solution is conveniently expressed by pH: **$\text{pH} = -\log [\text{H}^+]$** Pure water has a pH of 7.

Note: Acidic solutions will have greater $[\text{H}^+]$ and $\text{pH} < 7$. Basic solutions will have smaller $[\text{H}^+]$ and $\text{pH} > 7$.

Acid-Base Calculation

Important quantitative concepts in acid-base calculations are pH, pOH, K_w , K_a , and K_b . The following equations will be useful in the acid-base calculation:

For strong acids and bases, since there is complete dissociation, the $[\text{H}^+]$ and $[\text{OH}^-]$ is substituted into the equations for pH and pOH. For weak acids and bases, incomplete dissociation is achieved because of the reversible nature of the reaction. For example, the dissociation of the weak base ammonia is described by the reaction below:



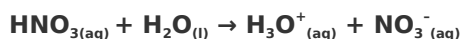
This means that not all of the NH_3 molecules convert to the hydroxide ion so the concentration of the hydroxide cannot be assumed equal to the concentration of the base. To get the hydroxide ion concentration, the equilibrium constant for the reaction needs to be accounted for. The following equations should be considered in the calculation.

The value obtained is valid if the percent of dissociation of the compound is $\leq 5\%$.

Sample Problems

Strong acid problem

What are the pH and pOH of a 0.15 M HNO_3 at 25°C?



Note: HNO_3 is a strong acid and so complete ionization occurs. This means for every mole of HNO_3 , one mole of H_3O^+ is produced.

$$\begin{aligned} [\text{H}_3\text{O}^+] &= [\text{HNO}_3] = 0.15 \text{ M} \\ \text{pH} &= -\log [\text{H}_3\text{O}^+] = -\log [0.15] = 0.82 \\ \text{pOH} &= 14 - \text{pH} = 14 - 0.82 = 13.18 \end{aligned}$$

Strong base problem

What are the pH and pOH of a 0.15 M $\text{Ca}(\text{OH})_2$ at 25°C?



Note: $\text{Ca}(\text{OH})_2$ is a strong base and so complete ionization occurs. This means for every mole of $\text{Ca}(\text{OH})_2$, 2 moles of OH^- is produced.

$$\begin{aligned} [\text{OH}^-] &= 2 [\text{Ca}(\text{OH})_2] = 2 [0.15 \text{ M}] = 0.30 \text{ M} \\ \text{pOH} &= -\log [\text{OH}^-] = -\log [0.30] = 0.52 \\ \text{pH} &= 14 - \text{pOH} = 14 - 0.52 = 13.48 \end{aligned}$$

Weak acid problem

What are the $[H^+]$ and $[OH^-]$ of a 0.15 M CH_3COOH at $25^\circ C$? What is the percent of dissociation of the solution ($K_a = 1.8 \times 10^{-5}$)?



Note: CH_3COOH is a weak acid, and so incomplete ionization occurs. This means not all of CH_3COOH is converted to its conjugate base. Since a reversible reaction is involved, equilibrium may be established, and so the ICE method may be used.



I	0.15 M	0	0
C	-x	+x	+x
E	0.15 - x	x	x

Note: Use the quadratic equation to solve for x.

$$\begin{aligned}x &= [H_3O^+] = 1.64 \times 10^{-3} \text{ M} \\pH &= -\log [H_3O^+] = -\log [1.64 \times 10^{-3}] = 2.78 \\pH &= 14 - pOH = 14 - 2.78 = 11.22\end{aligned}$$

Buffer Solutions

Buffers are solutions that can resist pH changes. This is possible because the buffer solution is saturated with acidic and basic ions that can readily react with an added base or acid, respectively. A primary requirement to produce a buffer solution is that the acidic and basic components of the buffer system must not be able to neutralize each other. This can only be achieved using a weak acid or base with its conjugate. For example, if we consider an acetic acid-acetate buffer, the two components are not able to neutralize each other.



In the process of "**neutralization**" seen above, the products are still the conjugate acid-base pair. For the acetic acid-acetate buffer system, when small amounts of acid are added, they will react with the acetate component, which has basic qualities, producing its conjugate acid. On the other hand, if small amounts of a base are added to the system, the acetic acid component will react to produce its conjugate base.

Henderson-Hasselbalch Equation

The **Henderson-Hasselbalch equation** is used to solve for the pH of buffer solutions. This equation is derived from the equilibrium constant expression for the acid's dissociation in water. For example, for the acetic acid-acetate buffer, the dissociation reaction of acetic acid is:



Getting the negative log of this, we will get the equation:

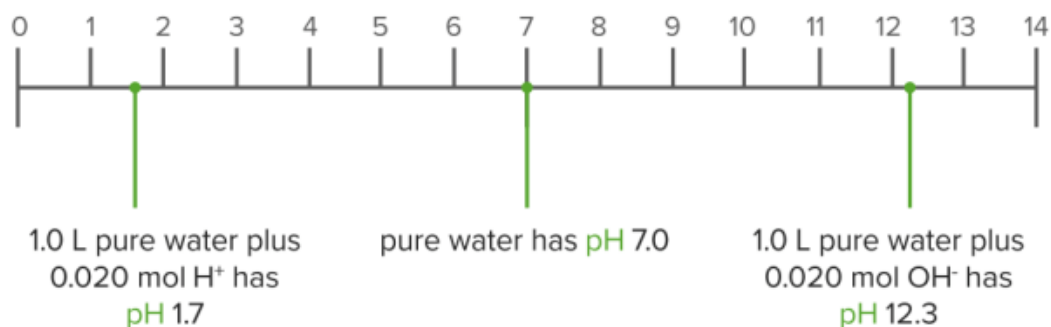
This means that if 1 L of a buffer solution contains 0.10 M CH_3COOH and 0.01 M $NaCH_3COO$, the pH of the solution will be equal to 3.74.

A buffer's efficiency to resist pH changes is called buffer capacity. It is formally defined as the amount of strong acid or base, in gram-equivalents, that must be added to 1 liter of a solution to change its pH by one unit.

One common **buffer solution is human blood**. [Human blood](#) samples maintain a pH of 7.4. Significant changes in the pH of blood may cause conditions like acidosis (pH < 7.4) or alkalosis (pH > 7.4). Buffer solutions are important in humans because some reactions necessary for metabolism will only occur at certain pH values of the solution.

Water as buffer

Water has no buffering capacity:



A buffered solution would be able to maintain its pH within a few hundredths of a pH unit after the addition of similar amounts of acid or base.

References

Burdge, J., & Overby, J. (2012). *Chemistry: atoms first*. New York, NY: McGraw-Hill Education.

Petrucci, R. H. (2017). *General chemistry: principles and modern applications*. Toronto: Pearson Canada.

Chang, R. (2010). *Chemistry* (10th ed.). Boston: McGraw-Hill.

Whitten, K. W., Davis, R. E., Peck, M. L., & Stanley, G. G. (2014). *Chemistry* (10th ed.). Belmont, CA: Brooks/Cole Cengage Learning.

Legal Note: Unless otherwise stated, all rights reserved by Lecturio GmbH. For further legal regulations see our [legal information page](#).