There are several ways to define acids and bases:

- Arrhenius (narrowest/most common definition)
  - acids – produce H^+ ions in water
  - bases – produce OH^- ions in water

- Bronsted-Lowry (broader definition)
  - acids – proton donors
  - bases – proton acceptors

- Lewis (broadest definition)
  - acids – electron acceptors
  - bases – electron donors

Arrhenius acids:

Any substance that causes hydrogen ions to be formed in an aqueous solution.

Consider hydrogen chloride gas, dissolved in water:

$$HCl(g) \xrightarrow{H_2O} \underset{\text{(a proton)}}{H^+}(aq) + Cl^-(aq)$$

A H^+ ion (a proton) is very reactive and does not actually exist in water. Instead, when the ion is formed it reacts with a water molecule to form a hydronium ion, H_3O^+.

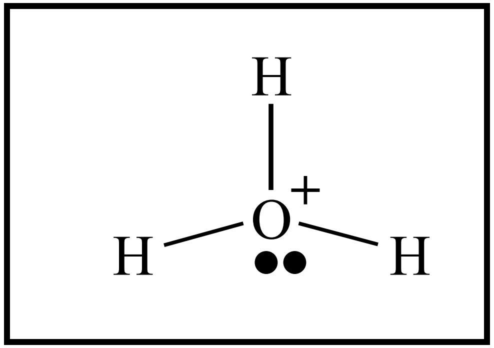
$$H^+ (aq) + H_2O \rightarrow H_3O^+ (aq)$$

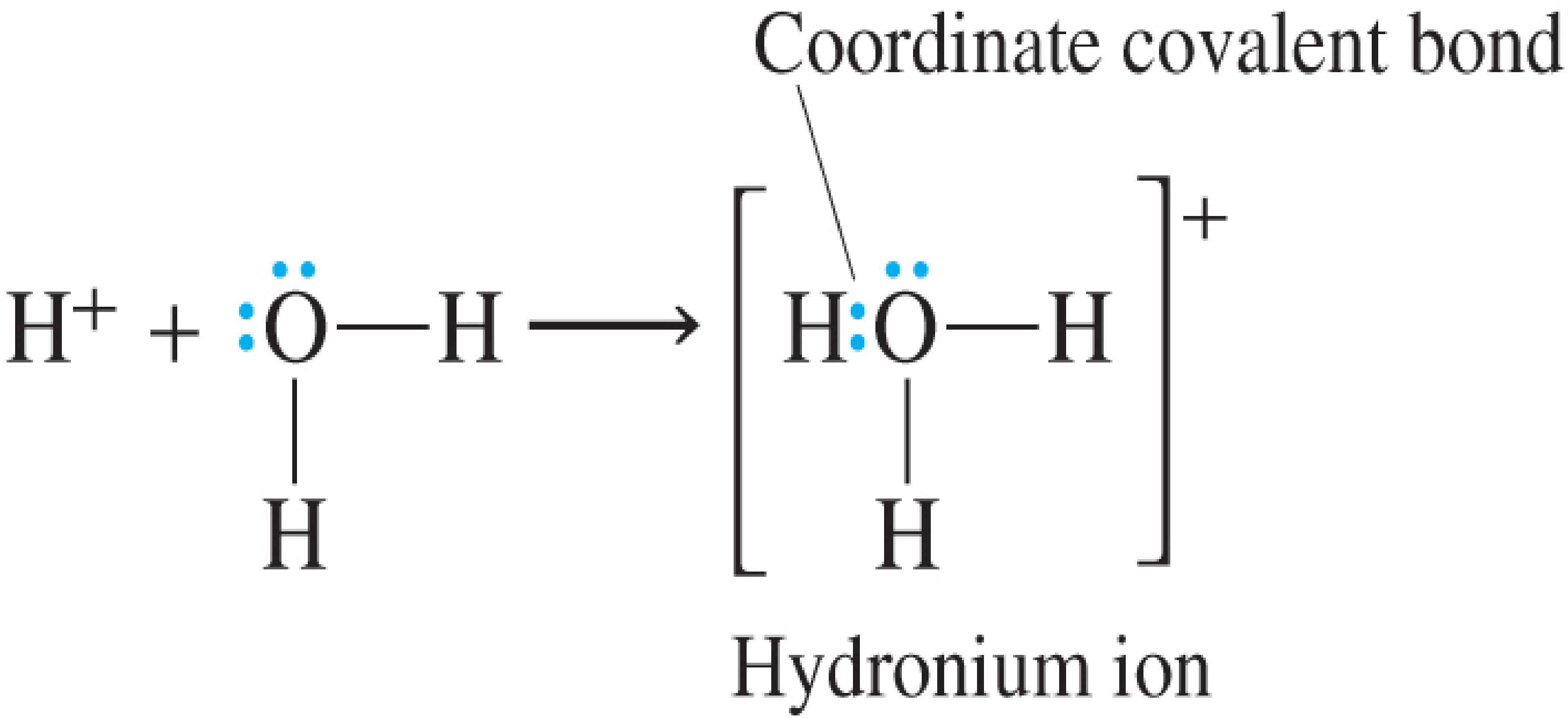
The chloride ion (Cl^-) remains unchanged, but is surrounded by water molecules to keep it dissolved.

The hydronium ion is often abbreviated:

$H_3O^+ (aq) = H^+ (aq)$

and in many cases, aqueous is assumed, so the (aq) is often not added.




Coordinate covalent bond
Hydronium ion

Therefore, the same reaction between gaseous HCl and water could be written in two different ways:

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

(a proton)

or:

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

hydronium ion

The Arrhenius definition for a base is:

Any substance that causes hydroxide ions (OH^-) to be produced in aqueous solution.

$$NaOH\ (s) + H_2O \rightarrow Na^+\ (aq) + OH^-\ (aq)$$

Notice that when written in this manner, the equation cannot be balanced. In this case water is the solvent, not a reactant, and more properly written above the arrow to indicate reaction conditions:

$$\underset{\text{Sodium hydroxide}}{NaOH(s)} \xrightarrow{H_2O} \underset{\text{Sodium ion}}{Na^+(aq)} + \underset{\text{Hydroxide ion}}{OH^-(aq)}$$

Calcium hydroxide produces 2 equivalents of OH^-:

$$Ca(OH)_2(s) \xrightarrow{H_2O} Ca^{2+}(aq) + 2OH^-(aq)$$

What about ammonia (NH_3)? It is a base, but has no OH group.

$$NH_3(aq) + H_2O(l) \longrightarrow NH_4OH(aq) \text{ ammonium hydroxide}$$

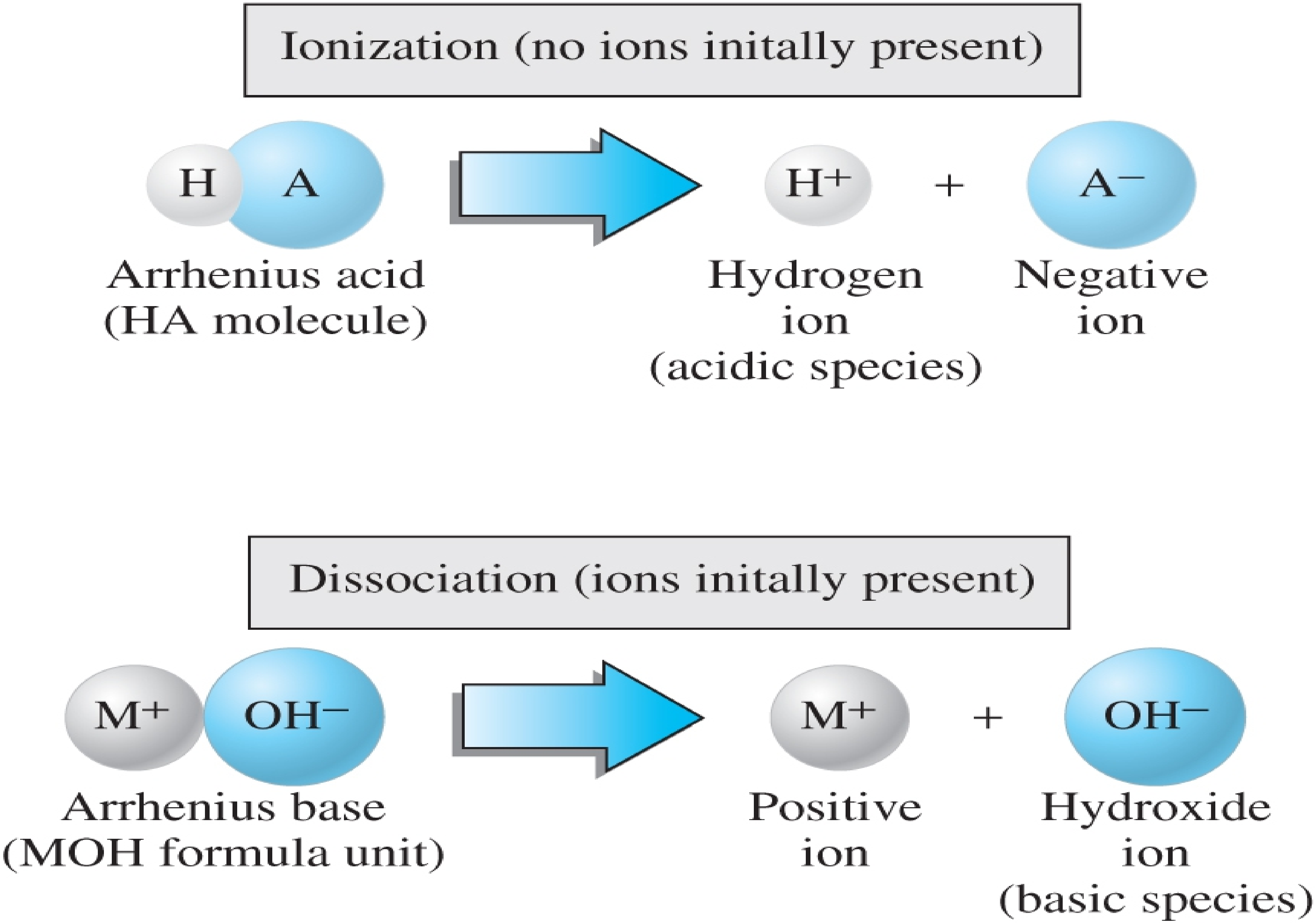
$$NH_4OH(aq) \xrightarrow{H_2O} NH_4^+(aq) + OH^-(aq)$$

- Ionization
  - When a molecule splits apart to form ions
  - Most acids ionize in water:

$$HNO_3\ (aq) \rightarrow H^+\ (aq) + NO_3^-\ (aq)$$

- Dissociation
  - When an ionic compound separates into its individual ions
  - Most bases dissociate in water:

$$Ca(OH)_2\ (aq) \rightarrow Ca^{+2} + 2\ OH^-\ (aq)$$


Ionization (no ions initally present)
H A
Arrhenius acid
(HA molecule)
H+
+
A−
Hydrogen
ion
(acidic species)
Negative
ion
Dissociation (ions initally present)
M+ OH−
Arrhenius base
(MOH formula unit)
M+
+
OH−
Positive
ion
Hydroxide
ion
(basic species)

Acid/Base Strength

The chemistry definition of strength refers only to how easily the acid or base forms ions in water:

- Strong acids/bases completely form ions
- Weak acids/bases only partially form ions, so stay as the initial compound
- Strength has nothing to do with concentration or reactivity

- Strong acids: fully ionize to form H^+ ions:

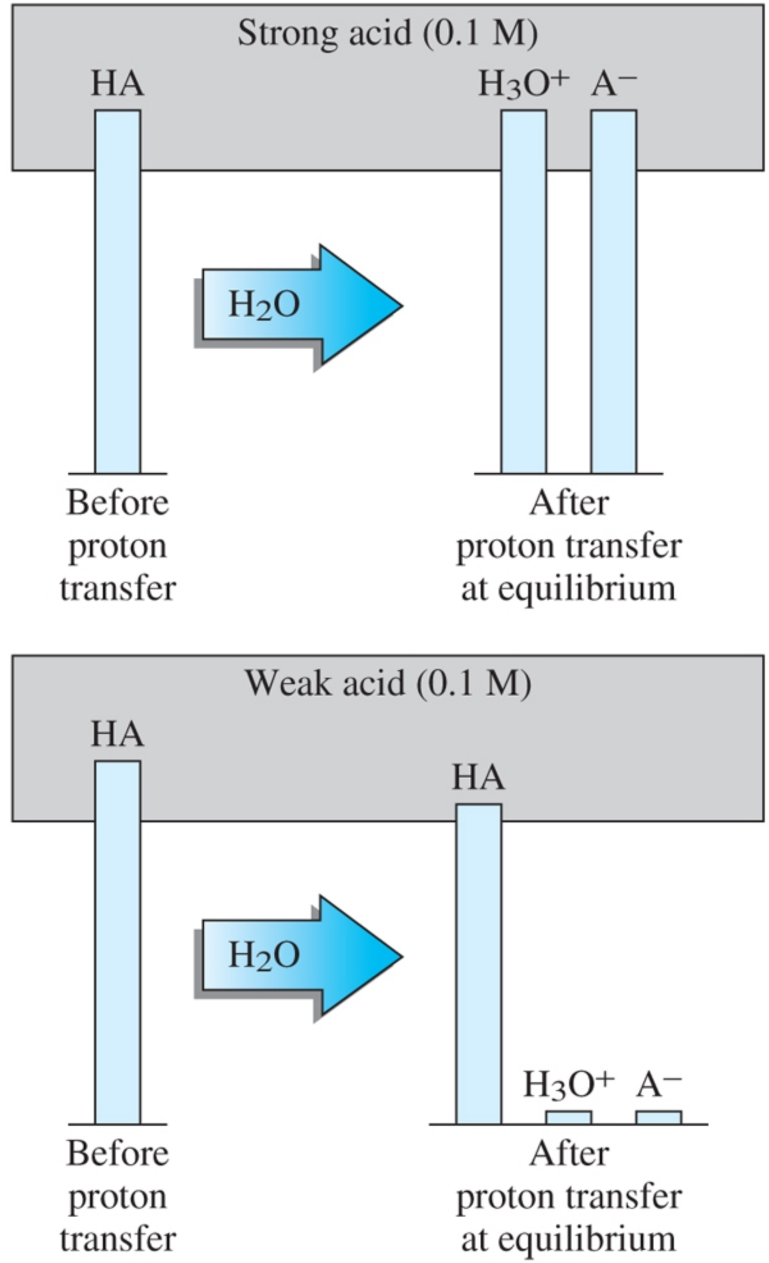
$$HCl\ (aq) \rightarrow H^{+1}\ (aq) + Cl^{-1}\ (aq)$$

99.9% of HCl molecules split apart

- Weak acids: only a small percentage of the molecules form hydronium ions:

$$HF\ (aq) \leftrightarrow H^{+1}\ (aq) + F^{-1}\ (aq)$$

less than 1% of HF molecules split apart



Strong acid (0.1 M)
HA
H3O+ A−
H2O
Before proton transfer
After proton transfer at equilibrium
Weak acid (0.1 M)
HA
HA
H2O
H3O+ A−
Before proton transfer
After proton transfer at equilibrium

- Strong acids:
  - HCl – hydrochloric acid
  - HBr – hydrobromic acid
  - HNO_3 – nitric acid
  - H_2SO_4 – sulfuric acid
  - $HClO_3$ – chloric acid
- Weak acids:
  - CH_3COOH – acetic acid
  - HF – hydrofluoric acid
  - H_3PO_4 – phosphoric acid

- Soluble ionic compounds completely split apart (dissociate) into individual ions in water.

- Therefore, soluble metal hydroxides fully split apart and are strong bases:

$$NaOH\ (aq) \rightarrow Na^{+1}\ (aq) + OH^{-1}\ (aq)$$

- For 'insoluble' metal hydroxides, only a very small amount of ions are formed:

$$FeOH\ (s) \leftrightarrow Fe^{+1}\ (aq) + OH^{-1}\ (aq)$$

most of the compound remains as a solid

- Strong/weak – refers to % ionization

- Concentrated/dilute – refers to the amount of acid/base per liter of solution

- Reactivity – depends on the acid or base, and with what it is reacting

- HCl
  - Strong acid
  - Concentrated in your stomach
  - Does not react readily with stomach lining
  - Does react to digest food particles

- HF
  - Weak acid
  - Extremely reactive with most substances
  - Even dilute solutions eat through glass

- $HC_2H_3O_2$ – acetic acid
  - Weak acid
  - Dilute in vinegar
    - used to make pickles, salad dressing...
  - Concentrated solutions would be hazardous to ingest.

When acids and bases react with each other, we call this a **neutralization reaction**.

$$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$$

$$\textit{acid} + \textit{base} \longrightarrow \text{salt} + \text{water}$$

In **neutralization** reactions, hydrogen ions from an acid combine with the hydroxide ions from a base to form molecules of water (hydrogen hydroxide).

The other product is a salt (an ionic compound).

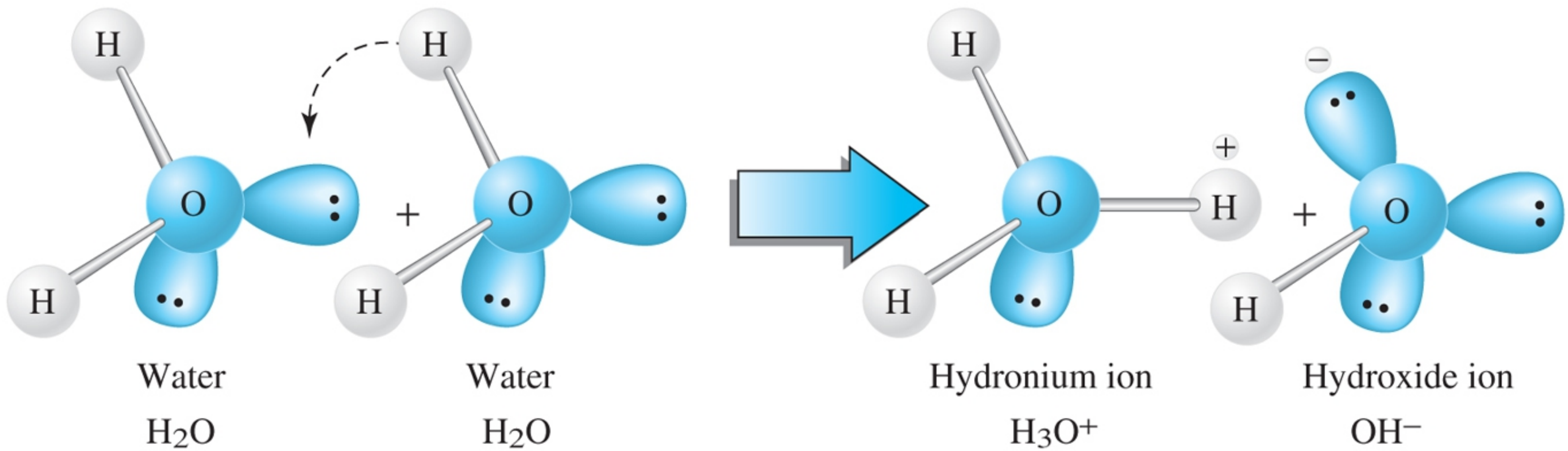
Amphoteric compounds can act as a Bronsted-Lowry acid or base (donate or accept a proton):

$$HCl + H_2O \rightarrow Cl^{-1} + H_3O^{+1}$$

here water acts as a base

$$NH_3 + H_2O \rightarrow NH_4^{+1} + OH^{-1}$$

here water acts as an acid


H
H
O
H
H
O
H
H
O
H
H
O
Water
H_2O
Water
H_2O
Hydronium ion
H_3O^+
Hydroxide ion
OH^-

The hydroxide and hydronium concentrations in water are related by the expression:

$$K_w = [H^+][OH^-] = 1 \text{ x } 10^{-14} \quad (\text{at } 25\ ^oC)$$

where K_w is the ion-product constant for water, and remember that [x] means molarity (moles/liters) of x.

Since K_w is a constant, as the hydronium concentration is increased, the hydroxide concentration decreases, and vice versa.

Knowing the hydroxide ion concentration, we can calculate the $[H^+]$ and vice versa.

$$K_w = [H^+][OH^-] = 1 \times 10^{-14} \quad (\text{at } 25\ {}^oC)$$

This expression shows that both ions, H^+ and OH^-, are always present in water.

For acidic solutions, the hydronium concentration is greater.

For basic solutions, the hydroxide concentration is greater.

For neutral solutions, the two concentrations are equal (10^{-7} M), not zero.

pH is used to compare the concentration of the H^+ ions present in aqueous solutions (easier to compare than using exponential values).

The mathematical expression for pH is a log-based scale and is represented as:

$$\boxed{pH = -\log[H^+]}$$

So for a solution with a

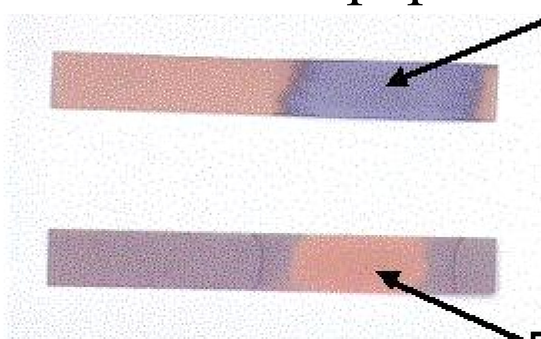
$[H^+] = 1.0 \times 10^{-3}$ M, the $pH = -\log(1.0 \times 10^{-3})$, or $-(-3.0) = 3.0$

Since pH is a log scale based on 10, a pH change of 1 unit represents a power of 10 change in $[H^+]$.

That is, a solution with a pH of 2 has a $[H^+]$ ten times that of a solution with a pH of 3.

Measuring pH with a pH meter

Or with litmus paper



Base applied to red litmus

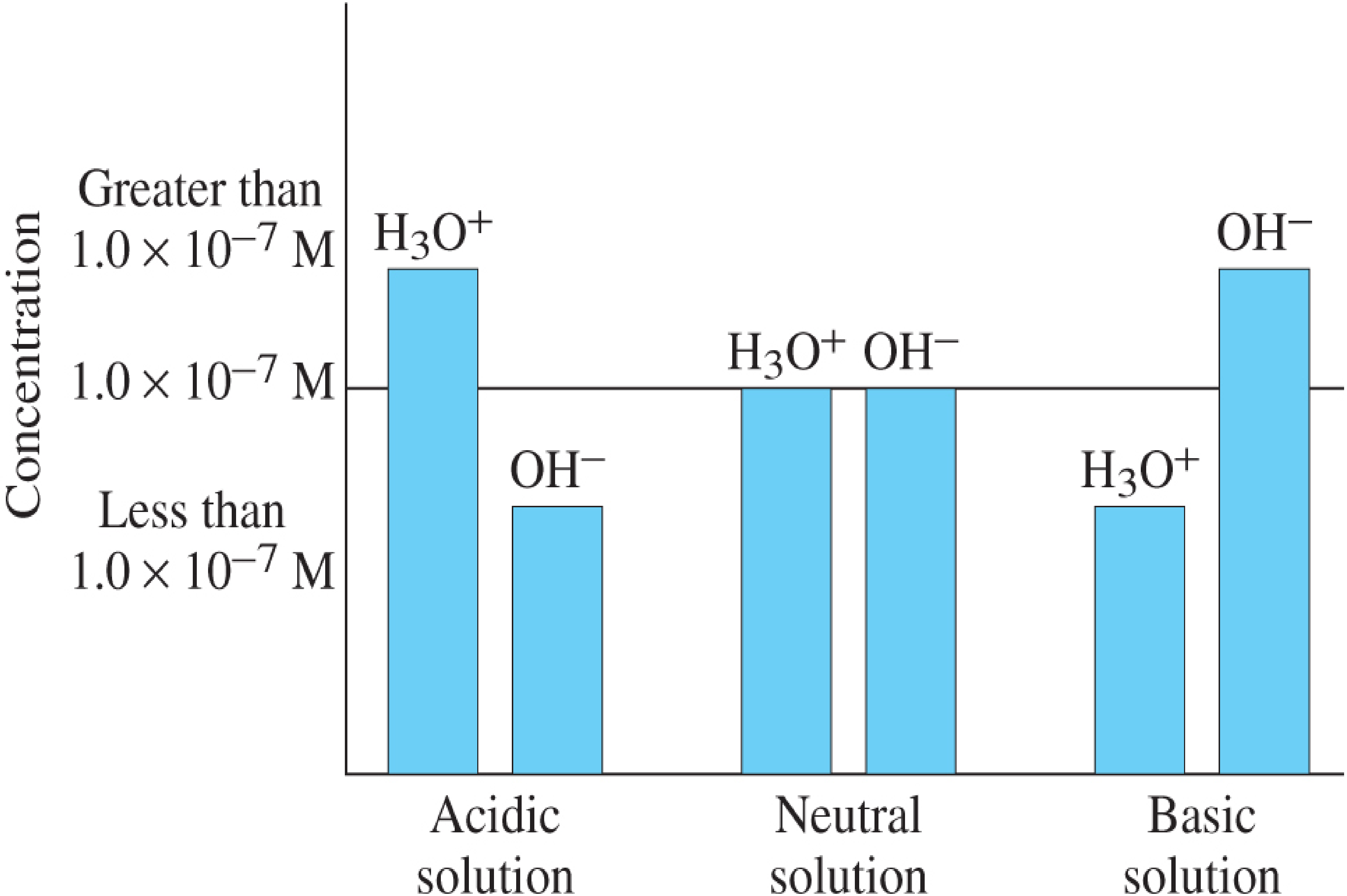
Acid applied to blue litmus

To measure the pH in basic solutions, we make use of the expression below to calculate $[H^+]$ from $[OH^-]$.

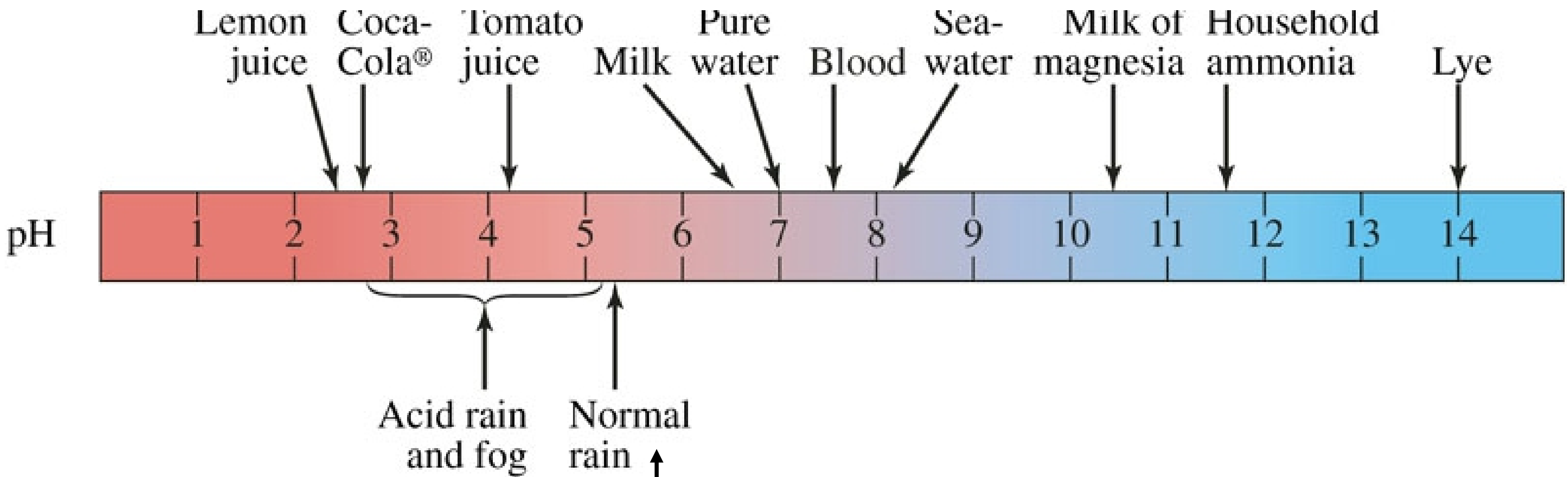
$$K_w = [H^+][OH^-] = 1 \text{ x } 10^{-14} \quad (\text{at } 25\ ^{o}\text{C})$$

The three possible aqueous solution situations are:

$[H^+] = [OH^-]$ a neutral solution (pH = 7)
$[H^+] > [OH^-]$ an acidic solution (pH < 7)
$[H^+] < [OH^-]$ a basic solution (pH > 7)


Concentration
Greater than
1.0×10^{-7} M
1.0×10^{-7} M
Less than
1.0×10^{-7} M
H_3O^+
OH^-
H_3O^+
OH^-
H_3O^+
OH^-
Acidic
solution
Neutral
solution
Basic
solution

Common substances and their pH values



Note that "normal" rain is slightly acidic.

Rain water is naturally slightly acidic due to dissolved carbon dioxide:

$$CO_2\ (g) \leftrightarrow CO_2\ (aq)$$

$$CO_2\ (aq) + H_2O\ (l) \leftrightarrow H_2CO_3\ (aq)$$

$$H_2CO_3\ (aq) \leftrightarrow H^{+1}\ (aq) + HCO_3^{-1}\ (aq)$$

In fact any pure water left exposed to air will become slightly acidic due to small amounts of dissolved carbon dioxide

Monoprotic acids produce one proton per molecule:

HCl, $HC_2H_3O_2$

Diprotic acids can produce two protons per molecule:

H_2SO_4, H_2CO_3

Triprotic acids can produce three protons per molecule:

H_3PO_4

Conjugate acid/base pairs (Bronsted-Lowry)

The conjugate base of an acid has one less proton and also a charge that is one less than that of the acid.

The conjugate acid of a base has one more proton and a charge that is one greater than that of the base

$$\underset{\text{Acid}}{HA} + \underset{\text{Base}}{B} \rightleftharpoons \underset{\text{Conjugate acid}}{HB^+} + \underset{\text{Conjugate base}}{A^-}$$

Conjugate pair

$$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$$

Acid	Base	Acid	Base

Conjugate pair

$$\underbrace{\overbrace{HBr(aq) + \underbrace{H_2O(l) \rightleftharpoons H_3O^+}_{\text{Conjugate pair}}(aq) + Br^-}^{\text{Conjugate pair}}(aq)}$$

A very strong acid will have a very weak conjugate base, and vice versa.

Some compounds/ions may be both a conjugate acid or base, depending on the reaction.

Conjugate pair

$$HBr(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Br^-(aq)$$

Conjugate pair

Conjugate pair

$$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$$

Acid Base Acid Base

Conjugate pair

Conjugate acid–base pair

$$CN^- + H_2O \longrightarrow HCN + OH^-$$

CN^-	H_2O	HCN	OH^-
Proton acceptor	Proton donor	Weak acid	Makes solution basic

Conjugate acid–base pair

$$F^- + H_2O \longrightarrow HF + OH^-$$

F^-	H_2O	HF	OH^-
Proton acceptor	Proton donor	Weak acid	Makes solution basic