Chem With Com

Acids, Bases, and pH

- Chapter 14.1 (Bronsted-Lowry Acids and Bases)
- Chapter 14.2 (pH and pOH)

Adapted From: Paul Flowers, Edward J. Neth, William R. Robinson, PhD, Klaus Theopold, Richard Langley; Chemistry: Atoms First 2e; OpenStax; Feb 14, 2019; Houston Texas; https://openstax.org/books/chemistry-atoms-first-2e/pages/1-introduction

Acids and Bases

- Chapter 14.1 (Bronsted-Lowry Acids and Bases)

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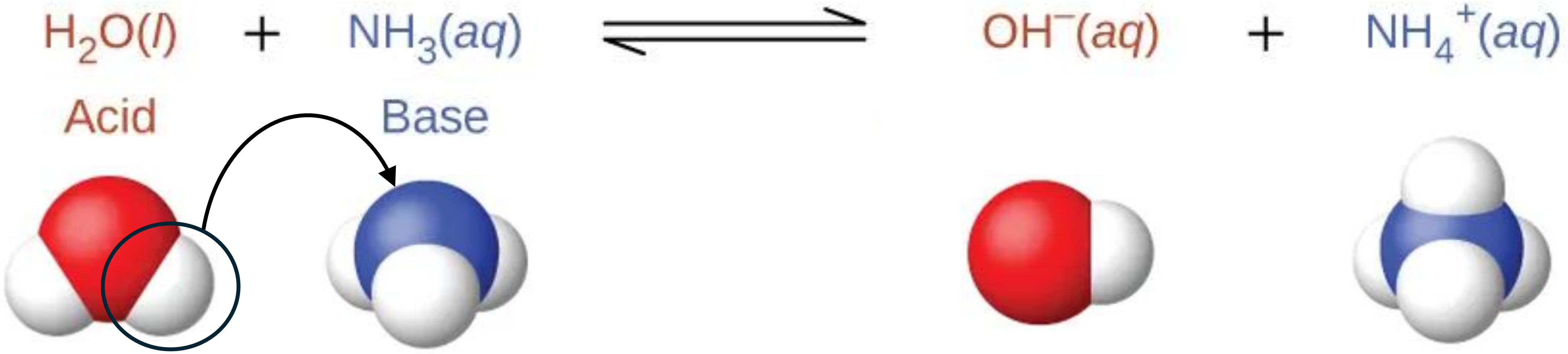
Acids donate H^{1+} and bases receive H^{1+} ions (protons)



ACIDS
donates H^{1+}

BASES
receives H^{1+}

Notes:
- H^{1+} are called a proton
- H^{1+} is a shortcut for H_3O^{1+}
- Acid-Base reactions usually occur in water

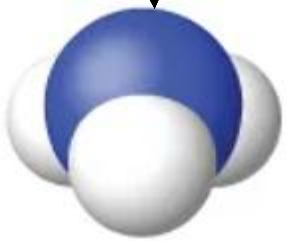
Acid-Base Conjugate Pairs

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

Acid | Base | Conjugate base | Conjugate acid

Conjugate Base:
What the acid
turned into after it
lost it’s proton

ACIDS
donates H^{1+}



Conjugate Acid:
What the base
turned into after it
gained it’s proton

BASES
receives H^{1+}

Common Acids and Bases

6 Strong Acids		6 Strong Bases	
$HClO_4$	perchloric acid	LiOH	lithium hydroxide
HCl	hydrochloric acid	NaOH	sodium hydroxide
HBr	hydrobromic acid	KOH	potassium hydroxide
HI	hydroiodic acid	$Ca(OH)_2$	calcium hydroxide
HNO_3	nitric acid	$Sr(OH)_2$	strontium hydroxide
H_2SO_4	sulfuric acid	$Ba(OH)_2$	barium hydroxide

CO_3^{2-} is amphoteric, and it can donate or receive a proton

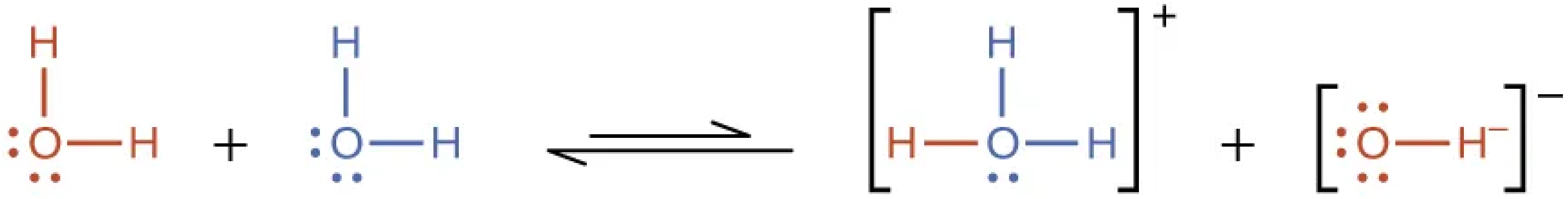
$$HCO_3^- \,(aq) + H_2O\,(l) \rightleftharpoons CO_3^{2-}\,(aq) + H_3O^+\,(aq)$$

$$HCO_3^- \,(aq) + H_2O\,(l) \rightleftharpoons H_2CO_3\,(aq) + OH^-\,(aq)$$

Another common moderately strong base:

Na_2CO_3 sodium carbonate (CO_3^{2-} accepts the protons)

Water autoionizes (turns into ions naturally)



$$H_2O + H_2O \rightleftharpoons H_3O^{+} + OH^{-}$$

When water autoionizes the concentration of each product (H_3O^{1+} and OH^{1-}) is very low at 1×10^{-7} M @25 °C

		hydronium	hydroxide
The ion-product constant of water is	$K_W =$	$[H_3O^{1+}]$	$[OH^{1-}]$
		1×10^{-7} M	1×10^{-7} M

$$\boxed{K_W = 1 \times 10^{-14}\ \text{M}}$$

The ion-product constant of water allows us to calculate $[H^{1+}]$ or $[OH^{1-}]$ if we know one of them

$$K_W = 1 \times 10^{-14} \qquad K_W = [H^{1+}]\,[OH^{1-}]$$

$$1.0 \times 10^{-14} = [H^{1+}]\,[OH^{1-}]$$

Brackets mean molarity [M]

Use this if you know [**OH$^-$**] and want to calculate [**H$^+$**]

$$[H^+] = \frac{1 \times 10^{-14}}{[OH^-]}$$

Use this if you know [**H$^+$**] and want to calculate [**OH$^-$**]

$$[OH^-] = \frac{1 \times 10^{-14}}{[H^+]}$$

$$H^+ \underset{[H^+] = \frac{1 \times 10^{-14}}{[OH^-]}}{\overset{[OH^-] = \frac{1 \times 10^{-14}}{[H^+]}}{\rightleftarrows}} OH^-$$

EXAMPLE 14.2

The Inverse Relation between $[H_3O^+]$ and $[OH^-]$

A solution of an acid in water has a hydronium ion concentration of 2.0×10^{-6} *M*. What is the concentration of hydroxide ion at 25 °C?

$$1.0 \times 10^{-14} = [H^{1+}]\,[OH^{1-}]$$

$$H^{+} \underset{[H^{+}] = \frac{1 \times 10^{-14}}{[OH^{-}]}}{\overset{[OH^{-}] = \frac{1 \times 10^{-14}}{[H^{+}]}}{\rightleftarrows}} OH^{-}$$

Check Your Learning

What is the hydronium ion concentration in an aqueous solution with a hydroxide ion concentration of 0.001 *M* at 25 °C?

$$1.0 \times 10^{-14} = [H^{1+}]\,[OH^{1-}]$$

$$[OH^-] = \frac{1 \times 10^{-14}}{[H^+]}$$

$$H^+ \rightleftharpoons OH^-$$

$$[H^+] = \frac{1 \times 10^{-14}}{[OH^-]}$$

pH and pOH

- Chapter 14.2 (pH and pOH)

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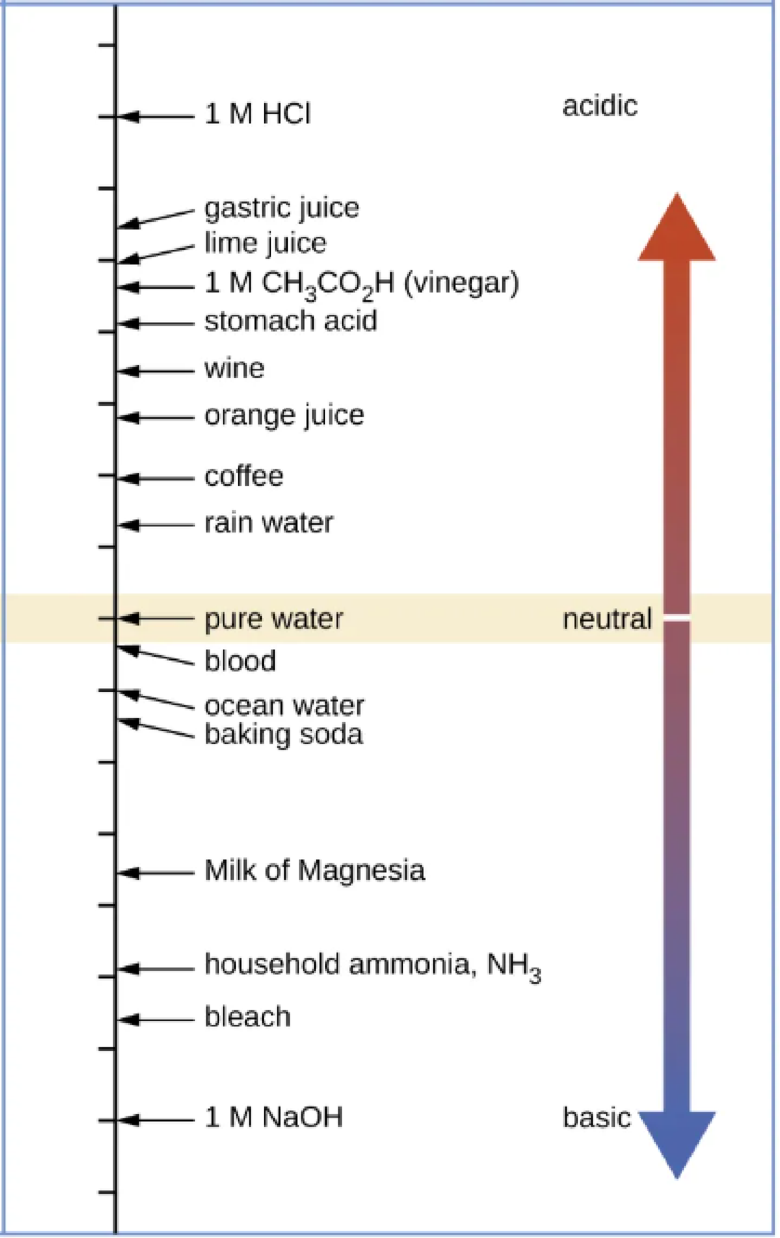
pH and pOH Scale

$[H_3O^+]$ (M)	$[OH^-]$ (M)	pH	pOH
10^{1}	10^{-15}	–1	15
10^{0} or 1	10^{-14}	0	14
10^{-1}	10^{-13}	1	13
10^{-2}	10^{-12}	2	12
10^{-3}	10^{-11}	3	11
10^{-4}	10^{-10}	4	10
10^{-5}	10^{-9}	5	9
10^{-6}	10^{-8}	6	8
10^{-7}	10^{-7}	7	7
10^{-8}	10^{-6}	8	6
10^{-9}	10^{-5}	9	5
10^{-10}	10^{-4}	10	4
10^{-11}	10^{-3}	11	3
10^{-12}	10^{-2}	12	2
10^{-13}	10^{-1}	13	1
10^{-14}	10^{0} or 1	14	0
10^{-15}	10^{1}	15	–1

Sample Solution



Chapter 14.2

Figure 14.2 The pH and pOH scales represent concentrations of H_3O^+ and OH^-, respectively. The pH and pOH values of some common substances at 25 °C are shown in this chart.

"p" means -log

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

pH is the negative log of the molarity of H^+

$$[H^+] = 1 \times 10^{-2}$$

$$pH = -\log[1\times10^{-2}]$$

$$pH = 2.0$$

pH	$[H_3O^+]$ (M)
–1	10^{1}
0	10^{0} or 1
1	10^{-1}
$\boxed{2}$	$\boxed{10^{-2}}$
3	10^{-3}

pOH	$[OH^-]$ (M)
15	10^{-15}
$\boxed{14}$	$\boxed{10^{-14}}$
13	10^{-13}
12	10^{-12}
11	10^{-11}

$$pOH = -\log[OH^-]$$

pOH is the negative log of the molarity of OH^-

$$[OH^-] = 1 \times 10^{-14}$$

$$pOH = -\log[1\times10^{-14}]$$

$$pOH = 14.0$$

Practice on your calculator!

You may have to use the "2nd" button to get to the 10^x which will allow you to input exponents. Another option is the ^ symbol

"log" button

Sometimes you may have to press "2nd" or "alpha" or "shift" to get to the function

On some calculators, the negative sign and minus signs are different

If you aren't sure when to use brackets, use them on all numbers with exponents like (5.1×10^{-8})

Chapter 14.2, https://ti84calc.com/ti84calc

Finding the $[H^+]$ or $[OH^-]$ from pH and pOH

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

$[OH^-] = 10^{-pOH}$

pH = 3.0

$[H^+] = 10^{-3.0}$

$[H^+] = 1 \times 10^{-3}$

pH	$[H_3O^+]$ (M)
–1	10^1
0	10^0 or 1
1	10^{-1}
2	10^{-2}
3	10^{-3}

pOH	$[OH^-]$ (M)
15	10^{-15}
14	10^{-14}
13	10^{-13}
12	10^{-12}
11	10^{-11}

pOH = 13.00

$[OH^-] = 10^{-13.00}$

$[OH^-] = 1.0 \times 10^{-13}$

SigFigs and Logarithms

Rules for Significant Figures in Logarithms and pH

Logarithm. When you take the logarithm of a number, keep as many significant figures to the right of the decimal point as there are significant figures in the original number. For example, log 4.000 (4 s.f.) = 0.6021(4 s.f. to right of the decimal point).	The pH of a solution with $H+ \ = 3.44M$: $\mathrm{pH} = \ \text{-}\log H+ \ = \ \text{-}\log 3.44 = 3.463$	3.44 has three s.f., so 3.463 is reported to three decimal places.
Antilogarithms; pH. Conversely, when you take the antilogarithm of a number, the result should have the same number of significant figures as the number of significant decimal places in the principal value. For example, antilog 0.0334 (3 s.f. to the right of the decimal) = 1.08 (3 s.f.).	The $H+$ in a solution with a pH of 2.55: $H+ \ = 10^{-2.55} \ = 2.8 \times 10^{-3}$	2.55 has two decimal places, so 2.8 is reported to two significant figures.

https://cxp.cengage.com/contentservice/assets/owms01h/references/significantfigures/index.html

Calculate pH if the $[H^+]$ is 1.34×10^{-2}

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

$[H^+] = 1.34 \times 10^{-2}$

$\mathrm{pH} = -\log[1.34 \times 10^{-2}]$

$\mathrm{pH} = 1.873$

3 significant figures,
3 decimals

Calculate $[OH^-]$ if the pOH is 13.00

$[OH^-] = 10^{-\mathrm{pOH}}$

$\mathrm{pOH} = 13.35$

$[OH^-] = 10^{-13.35}$

$[OH^-] = 4.5 \times 10^{-14}$

2 decimals,
2 sigfigs

pH to pOH is the simplest of them all!

Finally, the relation between these two ion concentration expressed as p-functions is easily derived from the K_w expression:

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

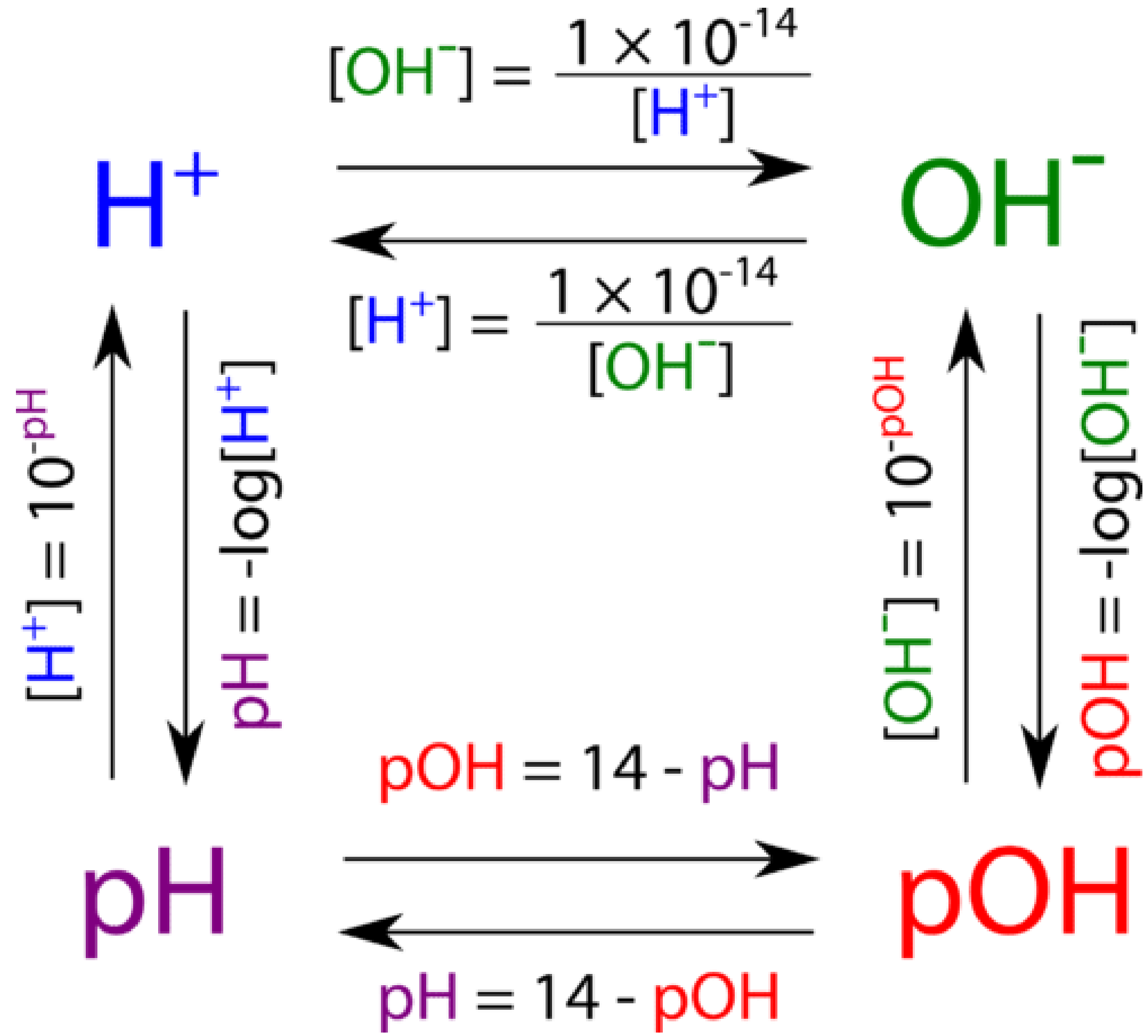
$$-\log K_w = -\log\left([H_3O^+]\,[OH^-]\right) = -\log[H_3O^+] + -\log[OH^-]$$

$$pK_w = pH + pOH$$

At 25 °C, the value of K_w is 1.0×10^{-14}, and so:

$$\boxed{14.00 = pH + pOH}$$

pH, pOH, $[H^+]$, and $[OH^-]$ conversions

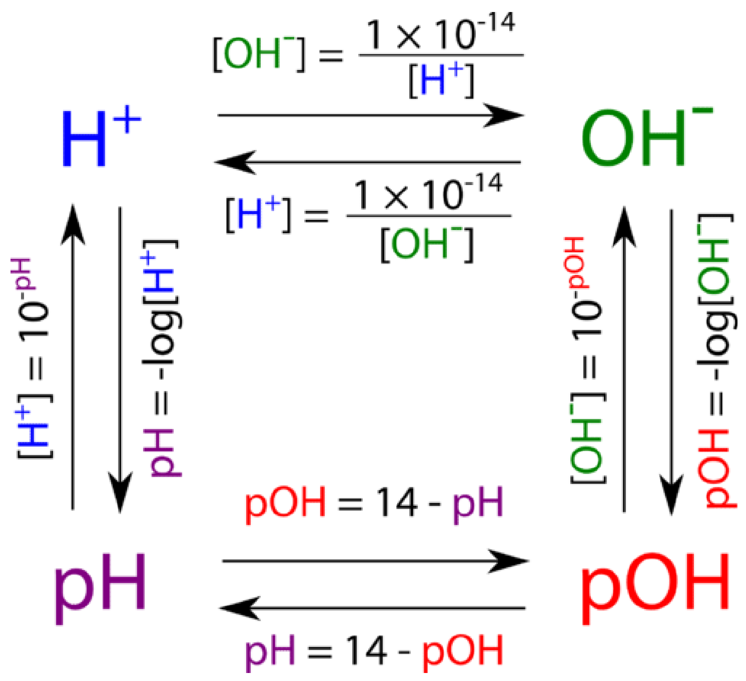


Check Your Learning

Water exposed to air contains carbonic acid, H_2CO_3, due to the reaction between carbon dioxide and water:

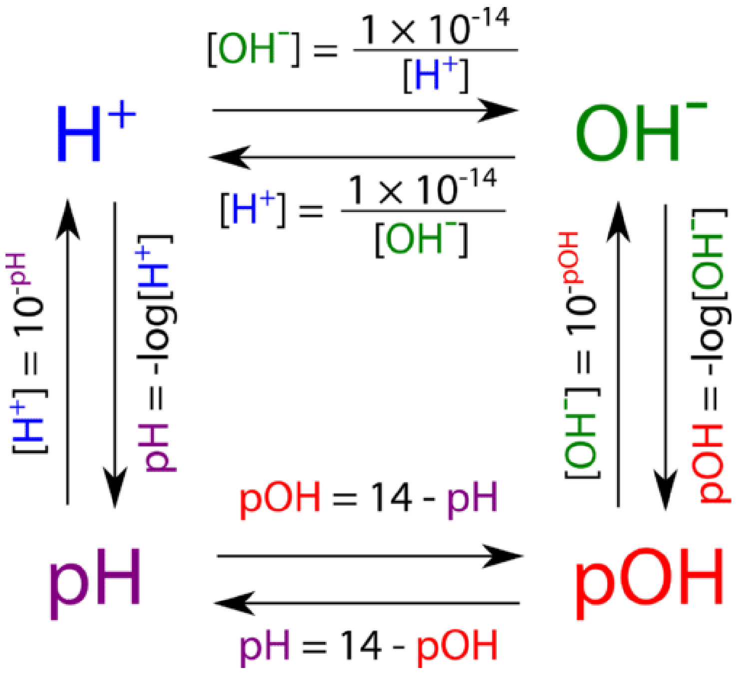
$$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

Air-saturated water has a hydronium ion concentration caused by the dissolved CO_2 of 2.0×10^{-6} *M*, about 20-times larger than that of pure water. Calculate the pH of the solution at 25 °C.



Check Your Learning

Calculate the hydronium ion concentration of a solution with a pH of −1.07.



Check Your Learning

The hydronium ion concentration of vinegar is approximately 4×10^{-3} *M*. What are the corresponding values of pOH and pH?