

CHAPTER 15 ACIDS & BASES

Acid Base Definitions - Arrhenius vs Bronsted-Lowry

Arrhenius – Aqueous solutions only

Acid: Substance contains hydrogen and dissociates to form H_3O^+ in water

Base: Substance contains hydroxide group and dissociates to form OH^- in water

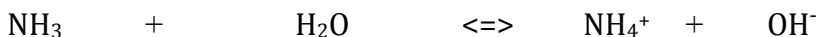
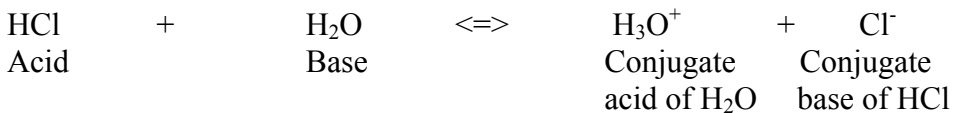
Bronsted-Lowry – Any solvent, even gas

Acid: Proton (H^+) donor

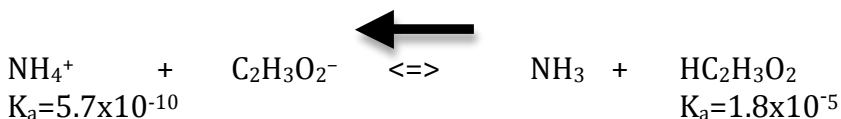
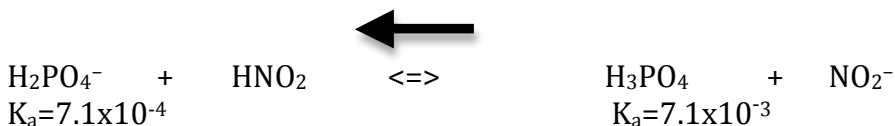
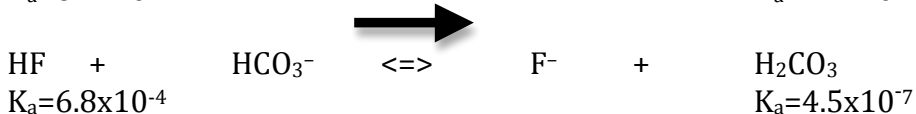
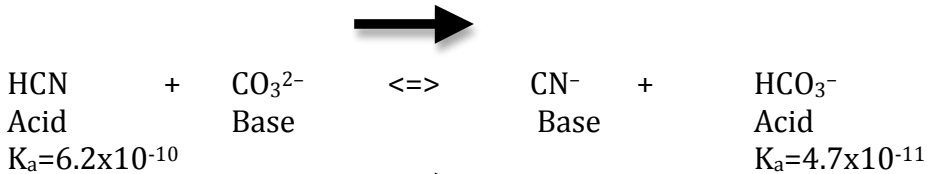
Base: Proton acceptor

Any substance that is an Arrhenius acid is an Bronsted-Lowry acid. Likewise, any Arrhenius base is a Bronsted-Lowry base.

Much Broader Range



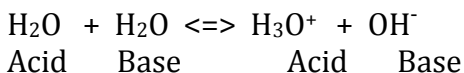
Label conjugate acid-base pairs and predict direction of reaction (that is, which way equilibrium lies)



The stronger the acid, the weaker is the conjugate base and vice-versa.

Dissociation of Water

Water acts as both Bronsted-Lowry acid & base



But water is pure liquid so $[\text{H}_2\text{O}]^2$ is combined with K_c yielding K_w .

$$\begin{aligned} K_c [\text{H}_2\text{O}]^2 &= [\text{H}_3\text{O}^+][\text{OH}^-] \\ K_w &= [\text{H}_3\text{O}^+][\text{OH}^-] \end{aligned}$$

At 25°C $K_w = 1.0 \times 10^{-14}$

So in neutral water $[\text{H}_3\text{O}^+] = [\text{OH}^-] = x$ (same)

$$[\text{H}_3\text{O}^+][\text{OH}^-] = K_w$$

$$(x)(x) = 1.0 \times 10^{-14}$$

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

$$x = [\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$$

Definitions

Neutral $\Rightarrow [\text{H}_3\text{O}^+] = [\text{OH}^-]$

Acid $\Rightarrow [\text{H}_3\text{O}^+] > [\text{OH}^-]$

Base $\Rightarrow [\text{H}_3\text{O}^+] < [\text{OH}^-]$

$\text{pH} = -\log[\text{H}_3\text{O}^+]$ $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$

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

$\text{p}K_w = -\log K_w$ $K_w = 10^{-\text{p}K_w}$

At 25°C, neutral $\text{pH} = -\log(1.0 \times 10^{-7}\text{M}) = 7.00$

$\text{pOH} = -\log(1.0 \times 10^{-7}\text{M}) = 7.00$

Relationships between pH and pOH

$$[\text{H}_3\text{O}^+][\text{OH}^-] = K_w$$

$$\log[\text{H}_3\text{O}^+] + \log[\text{OH}^-] = \log K_w$$

$$-\log[\text{H}_3\text{O}^+] + -\log[\text{OH}^-] = -\log K_w$$

$$\text{pH} + \text{pOH} = \text{p}K_w$$

At 25°C, $K_w = 1.0 \times 10^{-14}$

so $\text{p}K_w = -\log(1.0 \times 10^{-14}) = 14.00$

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

Is pure water always 7? K_w values change with temperature.

At 100°C, $K_w = 5.5 \times 10^{-13}$

So if pH scale had been set at 100 °C:

$$[\text{H}_3\text{O}^+][\text{OH}^-] = K_w = 5.5 \times 10^{-13}$$

At 100°C, $K_w = 5.5 \times 10^{-13}$

So if pH scale had been set at 100 °C:

$$[\text{H}_3\text{O}^+][\text{OH}^-] = K_w = 5.5 \times 10^{-13}$$

Neutral pH (x) (x) = 5.5×10^{-13}

$$x = [\text{H}_3\text{O}^+] = [\text{OH}^-] = 7.4 \times 10^{-7}\text{M}$$

$$\text{pH} = \text{pOH} = -\log(7.4 \times 10^{-7}\text{M}) = 6.13$$

At 100°C, $\text{pH} + \text{pOH} = \text{p}K_w = 12.26$

Strong Acids and Bases

Strong Acids have large K_a .

Include HCl , HBr , HI , HNO_3 , HClO_4 , and H_2SO_4 (first H^+)

Strong bases are soluble hydroxides of groups 1A and 2A.

Include LiOH , NaOH , KOH , LiOH , RbOH , CsOH , $\text{Ca}(\text{OH})_2$, $\text{Ba}(\text{OH})_2$, and $\text{Sr}(\text{OH})_2$

Calculation of pH of strong acids and bases

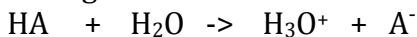
What is the pH of the following solutions?

0.15M HBr , 0.20M KOH , and 0.015M $\text{Ba}(\text{OH})_2$

For strong acids,

$[\text{H}_3\text{O}^+] = [\text{acid}]$ since 100% ionized.

Strong acid 100%

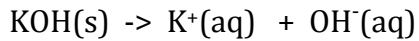


Note: $\text{HA} = \text{acid}$

0.15M $\text{HBr} \Rightarrow [\text{H}_3\text{O}^+] = 0.15 \text{ M}$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(0.15) = 0.82$$

Likewise for strong bases, 100% dissociation.



$$0.20\text{M KOH} \Rightarrow [\text{OH}^-] = 0.20\text{ M}$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log(0.20) = 0.70$$

$$\text{At } 25^{\circ}\text{C, } \text{pH} + \text{pOH} = 14.00$$

$$\text{At } 25^{\circ}\text{C, } \text{pH} + \text{pOH} = 14.00$$

$$\text{so } \text{pH} = 14.00 - \text{pOH} = 14.00 - 0.70 = 13.30$$

What's different about the 0.015M Ba(OH)₂ solution?

$$[\text{OH}^-] = 2 \times 0.015\text{ M} = 0.030\text{ M}$$

$$\text{pOH} = -\log(0.030) = 1.52$$

$$\text{so } \text{pH} = 14.00 - \text{pOH} = 14.00 - 1.52 = 12.48$$