

3/4/19

Autonization of Water

- Water is amphoteric
- In pure water, a few molecules act as bases and a few act as acids
- This is referred to as autonization



Equilibrium expression of autonization of water:

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

ion product constant for water

@ 25°C, $K_w = 1.0 \times 10^{-14}$

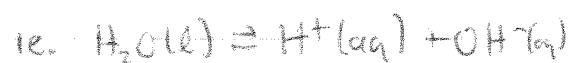
Ion Product Constant

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ @ } 25^\circ\text{C}$$

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

$$[\text{H}^+] > [\text{OH}^-] \Rightarrow \text{acidic}$$

$$[\text{H}^+] < [\text{OH}^-] \Rightarrow \text{basic}$$



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

$$[\text{OH}^-] = ? \text{ @ } 25^\circ\text{C}$$

$$[\text{OH}^-] = 1 \times 10^{-2} = 1 \times 10^{-12}$$

$$[\text{OH}^-] = 1 \times 10^{-12}$$

pH Scale

- In pure water @ 25°C

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1 \times 10^{-7} \text{ M}$$

- pH scale is a convenient way to report $[\text{H}_3\text{O}^+]$ in soln.

$$\text{pH} = -\log [\text{H}^+] \text{ @ } 25^\circ\text{C}$$

- neutral, $\text{pH} = 7$

- acidic, $\text{pH} < 7$

- basic, $\text{pH} > 7$

$$\text{ie. } [\text{H}^+] = 1 \times 10^{-9} \text{ M}$$

$$\Downarrow$$

$$\text{pH} = 9$$

pOH scale @ 25°C

- Neutral, $\text{pOH} = 7$

- Acidic, $\text{pOH} > 7$

- Basic, $\text{pOH} < 7$

$$\text{pH} + \text{pOH} = 14$$

Relationship pH and pOH

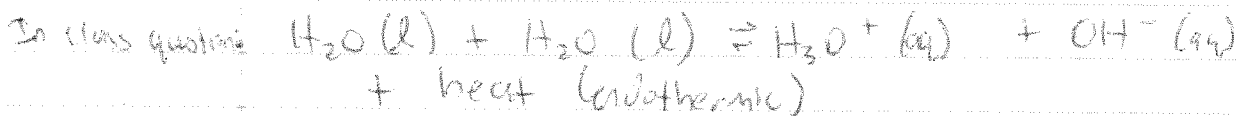
$$[\text{H}_3\text{O}^+][\text{OH}^-] = K_w = 1 \times 10^{-14}$$

- Take $-\log$ of everything

$$-\log [\text{H}_3\text{O}^+] + -\log [\text{OH}^-] = -\log K_w = 14.00$$

- Result @ 25°C

$$\text{pH} + \text{pOH} = \text{p}K_w = 14$$



increases temperature, shift right, so K_w INCREASE

K_w depends on temperature

In class question: pH goes down as temperature increases

Measuring pH

- pH meters

- Indicators: molecules that are one color in acidic form and another color in basic form

Strong Acids (SA)

- 7 strong acids: HCl , HBr , HI , HNO_3 , H_2SO_4 , HClO_3 , HClO_4

- Strong acids virtually completely dissociate in (aq) solution



- for monoprotic strong acids, assume

$$[\text{H}_3\text{O}^+] = [\text{acid}]_0$$

Calculating pH for SA

In class question: 1 drop of 2M HCl is added to 100 mL of water. pH?

$$\rightarrow 0.05 \times 10^{-3} \text{ L} \times 2 \text{ M} = 1 \times 10^{-4} \text{ mol HCl} = \text{mol H}_3\text{O}^+$$

$$0.02 \text{ drop} \approx 0.05 \text{ mL}$$

$$\frac{1 \times 10^{-4} \text{ mol H}_3\text{O}^+}{0.1 \text{ L solution}} = 1 \times 10^{-3} \text{ M}$$

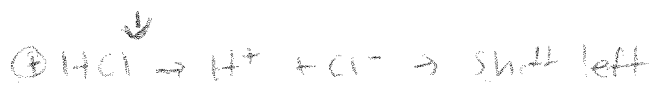
$$\text{pH} = 3$$



$$\downarrow$$

$$1 \times 10^{-7} \text{ M} @ 25^\circ\text{C}$$

Ignore autoionization of water
most of the time for SA



problems

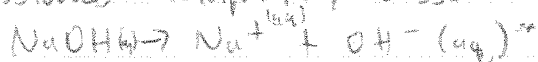
Add HCl to beaker of water so that $[\text{HCl}] = 1 \times 10^{-9} \text{ M}$. pH?

$1 \times 10^{-9} \text{ M H}^+$ from HCl
 $1 \times 10^{-7} \text{ M H}^+$ from water } about $1 \times 10^{-7} \text{ mol H}^+ \Rightarrow \text{pH} \approx 7$

Strong Bases (SB)

- Strong bases are the soluble hydroxides, which are the alkali metal hydroxides (ie. NaOH) and the heavier alkaline earth metal hydroxides (ie. $\text{Ca}(\text{OH})_2$)

- these substances completely dissociate in water (aq)



ex) add 40g of NaOH to water so final volume is 10 L. pH?



$$[\text{NaOH}]_0 = [\text{OH}^-]_{\text{aq}}$$

$$4.0 \text{ g NaOH} \times \frac{1 \text{ mol}}{40 \text{ g}} = 0.1 \text{ mol in 1 L of soln.}$$

$$\frac{0.1 \text{ mol NaOH}}{1 \text{ L}} = [\text{OH}^-] = 1 \times 10^{-1} \text{ M} \rightarrow \text{pOH} = 1$$

$$\text{pH} = 13$$

Comparing SA and Weak Acids (WA)

- SA completely dissociate to ions

- WA only partially dissociate to ions

Comparing SB and Weak Bases (WB)

- SB completely dissociate to ions

- WB only partially dissociate to ions

$$K_a = [\text{H}_3\text{O}^+][\text{A}^-] / [\text{HA}]$$

How strong is a WA?

HA(aq) For a WA, equation for dissociation is



- The equilibrium constant for this rxn is K_a , acid dissociation constant