The rest of the year:

Chapter: Topic

17: Acids & Bases: Equilibria & Reactivity

18: Buffers and Solubility: Equilibria & Reactivity

19: Thermodynamics: $\Delta G^o = -RT\ln K$, $\Delta G = \Delta H + T\Delta S$

20: Electron Transfer Reactions: Equilibria & Reactivity

The month ahead:

Topic: Chapter 17 & Chapter 18

17: Acids & Bases

18: Buffers and Solubility

Exam 2, on Chapter 15(end) 16, 17 & 18(start)
: **Nov 9, 6 PM**

***The* critical concept to understand is the reaction equilibrium:**

$HA + B \rightleftharpoons A^- + BH^+ \rightarrow K = [BH^+][A^-]/[HA][B]$

Chapter 17 key terms to know
(end of Ch 17 in KT)

17.1 – Water ionization constant (Kw)

17.2 – pH

17.3 – pOH

17.4 – pKw

Acids & Bases

Ch 17 GOALS:

- Use Brønsted-Lowry and Lewis concepts of bases
- Apply principles of chemical equilibrium to acids and bases
- Predict acid-base reaction outcomes
- Understand the influence of structure on acid-base properties

Today

- What are Brønsted acids & bases?
  - What is a conjugate acid-base pair?
- What is the ionization constant?
  - K_A (and K_B)
- What are the special properties of water?
  - K_W
- What are pH, pOH and p*K*?
- Manipulating K_A & K_B

Definition

- What is a Brønsted acid?
  A. A compound that can donate a proton
  ~~B. A compound that can accept a proton.~~
- A Brønsted base is a compound that can accept a proton

Ch 17 — Part 1

1

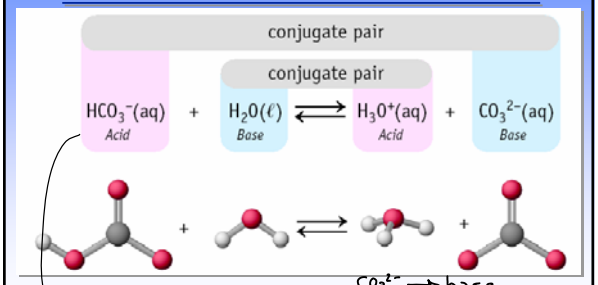
ACID-BASE THEORIES

- The most general theory for common aqueous acids and bases is the **BRØNSTED - LOWRY** theory
- **ACIDS DONATE H^+ IONS**
- **BASES ACCEPT H^+ IONS**

© 2006 Brooks/Cole - Thomson

2

Conjugate Pairs



CO_3^{2-} → base

amphiprotic → HCO_3^-

H_2CO_3 → acid

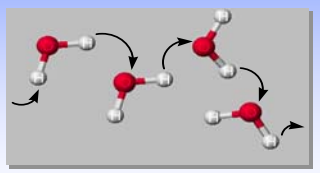
© 2006 Brooks/Cole - Thomson

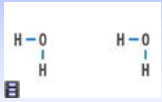
3

More About Water

H_2O can function as both an ACID and a BASE.

In pure water there can be **AUTOIONIZATION**





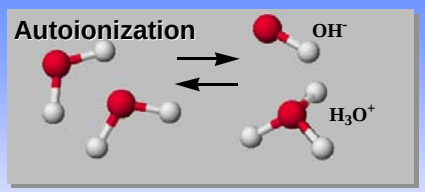
Equilibrium constant for autoion = K_w

$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ at 25 °C

© 2006 Brooks/Cole - Thomson

4

More About Water



$K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ at 25 °C

In a **neutral** solution $[H_3O^+] = [OH^-]$

so $K_w = [H_3O^+]^2 = [OH^-]^2$

and so $[H_3O^+] = [OH^-] = 1.00 \times 10^{-7}$ M

© 2006 Brooks/Cole - Thomson

5

$[H_3O^+]$, $[OH^-]$ and pH

A common way to express acidity and basicity is with pH

$$pH = \log(1/[H_3O^+])$$
$$= -\log[H_3O^+]$$

In a neutral solution,

$[H_3O^+] = [OH^-] = 1.00 \times 10^{-7}$ at 25 °C

$pH = -\log(1.00 \times 10^{-7})$
$= -(-7) = 7$

at 50°C

$K_w = 5.48 \cdot 10^{-14}$

pH of neutral solution = 6.6

© 2006 Brooks/Cole - Thomson

6

Calculating $[H_3O^+]$ & $[OH^-]$

You add 0.0010 mol of NaOH to 1.0 L of pure water. Calculate $[H_3O^+]$ and $[OH^-]$.

Solution

$2\,H_2O(liq) \leftrightarrows H_3O^+(aq) + OH^-(aq)$

Le Chatelier predicts equilibrium shifts to the left.

At the equilibrium, the contribution of $[H_3O^+]$ or $[OH^-]$ from the autoionization of water is $< 10^{-7}$ M.

$NaOH \rightleftharpoons Na^+ + OH^-$

© 2006 Brooks/Cole - Thomson

Ch 17 — Part 1

7

Calculating $[H_3O^+]$ & $[OH^-]$

You add 0.0010 mol of NaOH to 1.0 L of pure water. Calculate $[H_3O^+]$ and $[OH^-]$.

Solution

The initial concentration of OH^- is 0.0010 M

$K_w = 1.0\ 10^{-14} = [H_3O^+]\ [OH^-] = [H_3O^+]\ 0.0010$

$[H_3O+] = 1.0\ 10^{-14}/0.0010 = 1.0\ 10^{-11}$ M

Why we didn'T Take inTo account The H_3O^+ and OH^- ions produced by water autoionization?

© 2006 Brooks/Cole - Thomson

8

Calculating $[H_3O^+]$ & $[OH^-]$

You add 0.0010 mol of NaOH to 1.0 L of pure water. Calculate $[H_3O^+]$ and $[OH^-]$.

Solution

$[H_3O^+] = x$

$[OH^-] = 0{,}0010\ M + x$

$x < 10^{-7} \Rightarrow$ insignificant compared with 0,0010

Then $[OH^-] = 0{,}0010$ M

© 2006 Brooks/Cole - Thomson

9

$[H_3O^+]$, $[OH^-]$ and pH

What is the pH of the 0.0010 M NaOH solution?

$[H_3O^+] = 1.0 \times 10^{-11}$ M

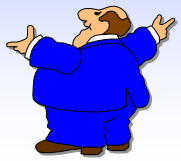
pH = - log (1.0×10^{-11}) = 11.00

General conclusion —

Basic solution pH > 7

Neutral pH = 7

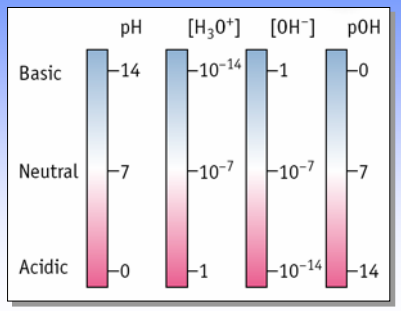
Acidic solution pH < 7



© 2006 Brooks/Cole - Thomson

10

The pH Scale



Active Figure 17.2

© 2006 Brooks/Cole - Thomson

11

Other pX Scales

In general pX = -log X

and so pOH = - log $[OH^-]$

$K_w = [H_3O^+]\ [OH^-] = 1.00 \times 10^{-14}$ at 25 °C

Take the log of both sides

$-\log(10^{-14}) = -\log[H_3O^+] + (-\log[OH^-])$

$pK_w = 14 = pH + pOH$

© 2006 Brooks/Cole - Thomson