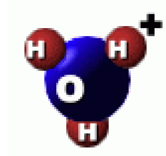
Chem Ch. 18 Notes: ACIDS & BASES

NOTE: Vocabulary terms are in **<u>boldface and underlined</u>**. Supporting details are in *italics.*

18.1 Notes



I. Properties and Examples of Acids and Bases

A. **Acids** *produce hydrogen ions (H^+) when dissolved in water*
… more accurately, they produce **hydronium ions = H_3O^+**

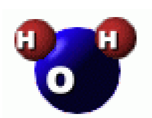
1) properties: tart, sour, form electrolytic solutions of ions
2) examples: citric acid, vinegar, hydrochloric acid, sulfuric acid
3) *usually have H in front of the formula* or COOH at the end: HCl, CH_3COOH
4) turns litmus paper red
5) pH less than 7

B. **Bases** *produce hydroxide ions (OH^-) when dissolved in water*

1) properties: bitter, slippery, form electrolytic solutions of ions
2) examples: lye/soap, ammonia, sodium hydroxide, other metal hydroxides
3) *usually have OH at the end of the formula*: KOH, NaOH, $Ca(OH)_2$
   - ammonia (NH_3) is a base, even though for formula doesn't look like one—it forms NH_4OH in water
4) turns litmus paper blue
5) pH greater than 7

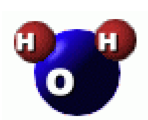
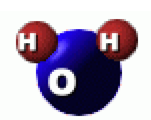
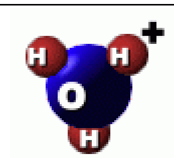
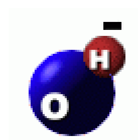
C. self-ionization of water:

NEUTRALIZATION OF WATER

H_2O → H^+ + OH^-

WATER → HYDROGEN ION + HYDROXIDE ION

<u>2</u> H_2O → H_3O^+ + OH^-

2 WATER molecules → HYDRONIUM ION + HYDROXIDE ION

Images from www.worsleyschool.net/science/files/pH/page.html

II. Models of Acid-Base Behavior

A. Arrhenius model (Svante Arrhenius, 1859-1927)

1) *Arrhenius acids*
   a) *produce hydrogen ions (H^+) when dissolved in water*
   b) acidic hydrogen—hydrogen atoms that will be given up by acids as hydrogen ions

$HA (aq) \rightarrow H^+ (aq) + A^- (aq)$

2) *Arrhenius bases—bases that produce hydroxide ions when dissolved in water*

$BOH (aq) \rightarrow B^+ (aq) + OH^- (aq)$

B. Brönsted-Lowry model
(Johannes Brönsted, 1879-1947) and (Thomas Lowry, 1843-1909)

1) *acid— hydrogen ion donor*
2) *base— hydrogen-ion acceptor*
3) conjugate base— what the acid becomes after it donating hydrogen ion
4) conjugate acid – what the base becomes after accepting hydrogen ion
5) *water can function as an acid or a base*
6) examples

E1)

$HF (aq)$	$+ H_2O (l)$	$\rightarrow H_3O^+ (aq)$	$+ F^- (aq)$
acid	base	conjugate acid	conjugate base

E2)

$NH_3 (aq)$	$+ H_2O (l)$	$\rightarrow NH_4^+ (aq)$	$+ OH^- (aq)$
base	acid	conjugate acid	conjugate base

C. "-protic" model

1) **monoprotic acids** donate 1 H^+ to the solution (HCl, HNO_3)

$HCl (aq) \rightarrow H^+ (aq) + Cl^- (aq)$
$HNO_3 (aq) \rightarrow H^+ (aq) + (NO_3)^- (aq)$
$HC_2H_3O_2 (aq) \rightarrow H^+ (aq) + (C_2H_3O_2)^- (aq)$

2) **diprotic acids** donate 2 H^+ to the solution (H_2S, H_2SO_4)

$H_2S(aq) \rightarrow \underline{2} H^+ (aq) + S^{2-} (aq)$
$H_2SO_4 (aq) \rightarrow \underline{2} H^+ (aq) + (SO_4)^{2-} (aq)$

3) **triprotic acids** donate 3 H^+ to the solution (H_3BO_3, H_3PO_4)

$H_3BO_3 (aq) \rightarrow \underline{3} H^+ (aq) + (BO_3)^- (aq)$
$H_3PO_4 (aq) \rightarrow \underline{3} H^+ (aq) + (PO_4)^{3-} (aq)$

4) **polyprotic acids** donate more than 1 H^+ to the solution (di- or tri-protic)

D. Lewis model (Gilbert Lewis, 1875-1946)

1) *Lewis acid— electron-pair acceptor*
2) *Lewis base— electron pair donor*

E. Anhydrides

1) **acidic anhydrides**—nonmetal oxides which react with water to form acids

$CO_2 + H_2O \rightarrow H_2CO_3$
$SO_3 + H_2O \rightarrow H_2SO_4$

2) **basic anhydrides**—metal oxides which react with water to form bases

$Na_2O + H_2O \rightarrow \underline{2} NaOH$
$ZnO + H_2O \rightarrow Zn(OH)_2$

18.2 Notes

III. Strengths of Acids and Bases

A. acid strength (see diagrams below and on the next page)

1) **strong acids**

a) *completely dissociate into ions*

b) *common examples: HCl, HNO_3, H_2SO_4*

2) **weak acids** *partially dissociate (not all come apart) into ions*

B. base strength

1) **strong bases**—*completely dissociate (come apart) into ions*

2) **weak bases** *partially dissociate (not all come apart) into ions*

C. strength vs. concentration

1) weak and strong refer to dissociation only

2) concentrated vs. dilute

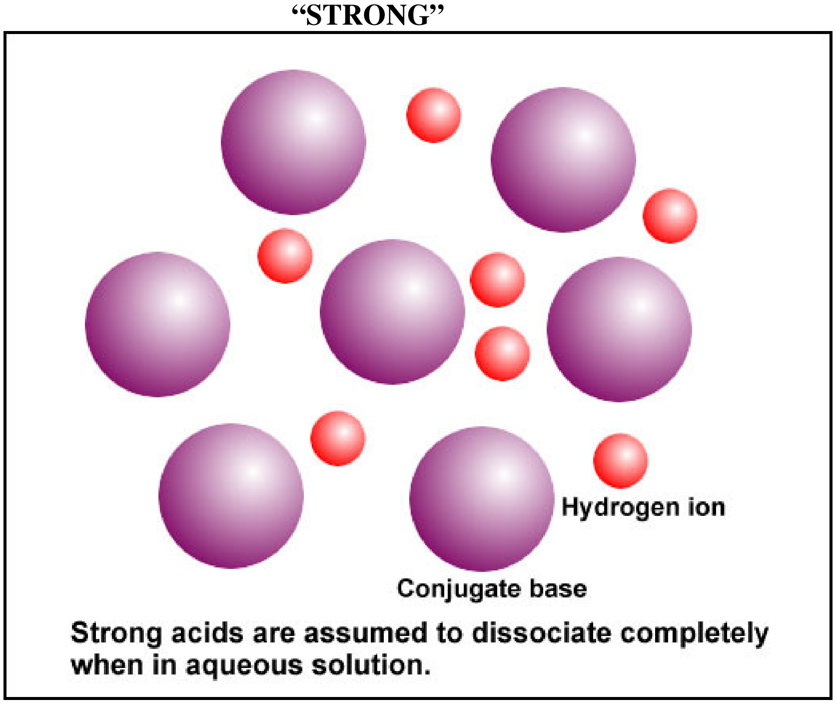
a) amount of particles in the solution

b) **molarity**—(M); *a measure of solution concentration in mol/L*

3) application

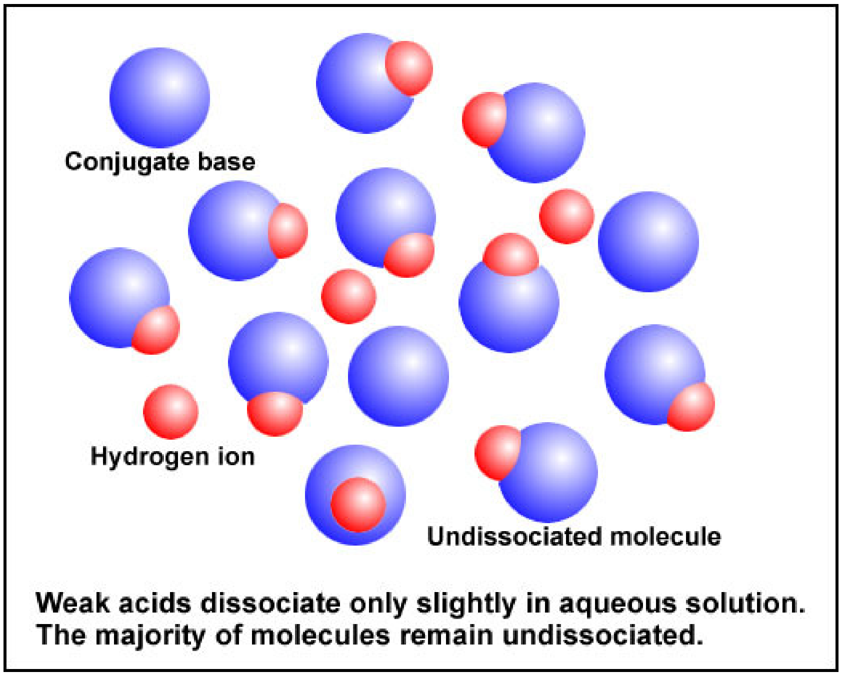
SAMPLE 1: a 0.10 M solution of H_2SO_4	SAMPLE 2: a 1.00 M solution of HF
Which is more concentrated?	(HF, because the molarity is higher)
Which is the stronger acid?	(H_2SO_4, because it is listed as a strong acid and HF isn't)

"STRONG"



www.chem.ubc.ca/.../pH/section1/content.html

"WEAK"



www.chem.ubc.ca/.../pH/section1/content.html

18.3 Notes

IV. pH (the power of Hydrogen)

A. neutrality of water

1) *Water is mostly neutral* $[H^+] = [OH^-]$

$[H^+] = 10^{-7}$ *M* and $[OH^-] = 10^{-7}$ *M*

2) Ion product constant for water = K_w

$$K_w = [H^+][OH^-] = 10^{-14} \text{ M}$$

3) Acidic solutions: $[H^+] > [OH^-]$

4) Basic (alkaline) solutions: $[OH^-] > [H^+]$

B. **<u>pH</u>** = the negative logarithm of the hydrogen ion concentration

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

C. *pH is a measure of the acidity or basic quality (alkalinity) of a substance*

D. pH values

1) *acid pH < 7*

2) *base pH > 7*

3) **<u>neutral</u>** *pH = 7*

0	**7**	**14**
strongest acid	**neutral**	**strongest base**

E. other important equations

$$pOH = -\log[OH^-]$$
$$pH + pOH = 14$$

F. examples

E3) If [H+] of a solution = 1.0×10^{-11} M…

a) Find pH. b) Find pOH. c) Find $[OH^-]$. d) Acid, base, or neutral?

a) $[H^+] = 1.0 \times 10^{-11}$ M so **pH = 11**

b) pH + pOH = 14 11 + pOH = 14 **pOH = 3**

c) $[H^+][OH^-] = 10^{-14}$ M $[10^{-11}][OH^-] = 10^{-14}$ M **$[OH^-] = 10^{-3}$ M**

d) base (pH = 11)

E4) If $[OH^-]$ of a solution = 1.0×10^{-9} M…

a) Find pOH. b) Find pH. c) Find $[H^+]$. d) Acid, base, or neutral?

a) $[OH^-] = 1.0 \times 10^{-9}$ M so **pOH = 9**

b) pH + pOH = 14 pH + 9 = 14 **pH = 5**

c) $[H^+][OH^-] = 10^{-14}$ M $[H^+][10^{-9}] = 10^{-14}$ M **$[H^+] = 10^{-5}$ M**

d) acid (pH = 5)

18.4 notes

V. Neutralization reactions

ACID + BASE → WATER + SALT

A. **neutralization**—when acid and base "cancel each other out"

B. *acid-base neutralization net ionic equation*: $H^+ + OH^- \rightarrow H_2O$

C. common acids

1) **hydrochloric acid = HCl**
2) **acetic acid = $HC_2H_3O_2$ or CH_3COOH**
3) **nitric acid = HNO_3**
4) **sulfuric acid = H_2SO_4**
5) **phosphoric acid = H_3PO_4**
6) **carbonic acid = H_2CO_3**

D. classic *double displacement reactions* **AB + CD → AD + CB**

1) You will have to write and balance these double displacement reactions.
2) If the formula is not provided, you must "crisscross" to get it.
3) Remember, to get the products, you must "un-crisscross" and "re-crisscross" the reactant ions.
4) If you have trouble balancing, keep water as H(OH) to make it easier.
5) Practice naming the salt that is formed.

E5) hydrochloric acid + strontium hydroxide → _____ + _____

$H^+ Cl^-$		$Sr^{2+} (OH)^-$		$H^+ (OH)^-$	$Sr^{2+} Cl^-$
A B	+	C D	→	A D +	C B
Acid	+	Base	→	Water +	Salt
2 HCl	+	$Sr(OH)_2$	→	H_2O +	$SrCl_2$
				salt = strontium chloride	

E6) phosphoric acid + magnesium hydroxide → _______ + _______

H^+ $(PO_4)^{3-}$ Mg^{2+} $(OH)^-$ H^+ $(OH)^-$ Mg^{2+} $(PO_4)^{3-}$

A B + C D → A D + C B

H_3PO_4 + $Mg(OH)_2$ → H_2O + $Mg_3(PO_4)_2$

$\underline{2}\ H_3PO_4$ + $\underline{3}\ Mg(OH)_2$ → $\underline{6}\ H_2O$ + $Mg_3(PO_4)_2$

salt = magnesium phosphate

VI. Titration

A. **titration**—adding a specific amount of a solution of known concentration to a solution of unknown concentration, to calculate the molarity (M) of the unknown solution

B. *standard solution*—the solution of known concentration

C. End point of the titration: when $[H^+] = [OH^-]$

D. **indicators**

a) a *dye* which is a different color in an acid vs. a base

b) phenolphthalein (PHTH) = clear in acid, "funky fuchsia" in base

c) other indicator dyes: methyl red, bromothymol blue, Orange IV…

TITRATION LAB SETUP