<u>Notes for Chapter 14/15 Section 1   AP Chemistry</u>

Most of this information can be found on p. 653-666 in your book (14.1-14.4)

Acid—has a sour taste
Base—has a bitter taste

3 Definitions of Acids and Bases:

1) **Arrhenius Acid/Base**
   Acid—a substance that produces hydrogen ions (H^+) in an aqueous solution.
   Base—a substance that produces hydroxide ions (OH^-) in an aqueous solution.
   This is the definition of Acid/Base that we have been limiting our identification of Acid/Base up until now.

   Example: (acid) **HCl** (aq) → $\mathbf{H^+}$(aq) + Cl^-(aq)
   (base) **NaOH** (aq) → Na^+(aq) + $\mathbf{OH^-}$(aq)

2) **Bronsted-Lowry Acid/Base** (Broader and **more inclusive** definition)
   Acid—a substance that is a proton **(H^+) donor** in an aqueous solution.
   Base—a substance that is a proton **(H^+) acceptor** in an aqueous solution.

   Hydrogen acids still "donate" or produce a H^+ ion just like the Arrhenius definition but Bronsted-Lowry uses water as the acceptor.
   Example: (acid) **HCl** + $\mathbf{H_2O}$ → $\mathbf{H_3O^+}$ + $\mathbf{Cl^-}$
   (acid) (base) (conj acid) (conj base)

   Note: What the acid "becomes" in the product is its "conjugate base".
   What the base "becomes" in the product is its "conjugate acid".

   Example: (base) $\mathbf{NH_3}$ + $\mathbf{H_2O}$ → $\mathbf{NH_4^+}$ + $\mathbf{OH^-}$
   (base) (acid) (conj acid) (conj base)

   Note: The Bronsred-Lowry **base** does not have to have a hydroxide, but it does produce a OH^- ion in the product.
   *Also, water **H_2O** can be **both** an **acid** or a **base** depending on whether it accepts or donates a proton. H_2O has 2 lone pair of electrons and 2 hydrogens so it can accept a proton on its lone pair or donate one of its protons and form OH^-.*

3) Lewis Acid/Base (Broadest and **most inclusive** definition) **We will not be using Lewis Acid/Bases.** I just want you to know what they are.
Acid—an electron pair **acceptor.** (H^+ ion has an empty orbital and accepts and e^- pair)
Base—an electron pair **donor.** (H_2O has 2 lone e^- pairs and can donate one)

Example: $H^+ + H_2O \rightarrow H_3O^+$ **or** $H^+ + OH^- \rightarrow H_2O$
H^+ (acid), H_2O (base); H^+ (acid), OH^- (base)

Even more different than you would expect:

$F^- + BF_3 \rightarrow BF_4^-$
F^- (base), BF_3 (acid)

F^- has 4 e^- pairs to donate. BF_3 , boron has an empty orbital to accept an e^- pair from F^-

This is kind of **"weird"** but it is an acid/base reaction. Lewis A/B reactions almost always have 2 substances becoming one. They are not used very much.
Bronsted-Lowry are the most important A/B reactions because they are more inclusive but easy to work with.

As previously stated, Bronsted-Lowry acids and bases always have their conjugate opposites on the product side of the reaction. (Refer back to the second definition examples.)

General Form: $HA + H_2O \leftrightarrow H_3O^+ + A^-$
HA (H-acid), H_2O (base), H_3O^+ (conj acid), A^- (conj base)

H_2O is always included in the reactants. These reactions are usually equilibrium reactions ($\leftarrow\rightarrow$) but not always.
(See p.655, Interactive 14.1)

(Refer to p.659-660)
Before proceeding, let's discuss the **properties of H_2O**.
H_2O can be **either** an **acid** or a **base** as seen in the Bronsted-Lowry examples.

H_2O is **self-ionizing** which means it reacts with itself to form H^+ and OH^- ions in pure H_2O. Water is both a proton donor and a proton acceptor.
(Remember we could write water as H—OH.) So the H can be donated and the e^- pairs on the OH can be H acceptors.
However, this self-ionizing is very little.

$$H_2O \quad + \quad H_2O \quad \leftarrow\rightarrow \quad H_3O^+ \quad + \quad OH^-$$

The K_C for water's self-ionization is **1.0×10^{-14}** and is known as the Equilibrium (Dissociation) Constant for water or **K_W**.

$$\text{Since} \quad K_W = \frac{[H_3O^+]\,[OH^-]}{\cancel{H_2O(l)}\ (1)} \qquad \text{Then} \quad \mathbf{K_W = [H_3O^+]\,[OH^-]}$$

H_3O^+ has a special name. We now will refer to it as the **Hydronium ion.**

K_W = $[H_3O^+]$ $[OH^-]$
$(1.0\times10^{-14}) = (1.0\times10^{-7})\ (1.0\times10^{-7})$ for Pure Water.

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pH is a measure of **H^+ (H_3O^+)** in a solution. **pOH** is a measure of **OH^-** in solution.

pH = - log (H_3O^+) For pure water, H_3O^+ was found to be 1.0×10^{-7} M.
So the **p H** of pure water is – log (1.0×10^{-7}) or **7** (neutral).

pOH = - log (OH^-) For pure water, the OH^- was found to be 1.0×10^{-7} M.
So the **p OH** of pure water is –log (1.0×10^{-7}) or **7** (neutral).

pK_W = - log (K_W) = -log $(.10\times10^{-14)}$ or **14** (The measure of the pH or pOH scale, 1—14).

These equations are very handy for finding the pH, $[H_3O^+]$, pOH, $[OH^-]$.
Remember, pH + pOH = 14.
So pH = 14 – pOH, and pOH = 14 – pH
This will give a process to follow to find any of the values we need.

The following can be put on and index card:

$[H_3O^+]$ → pH → pOH → $[OH^-]$
($-\log[H_3O^+]$) (14-pH) (inverse log –pOH)

$[H_3O^+]$ ← pH ← pOH ← $[OH^-]$
(inverse log –pH) (14-pOH) ($-\log[OH^-]$)

So given any single value in the above process, you can find the other 3.
(For **inverse log**, hit 2nd then log on the calculator.)

<u>Strong Acids and Bases</u>

Strong acids and strong bases are ones that dissociate or ionize (break apart) completely and do not reverse themselves (equilibrium reactions). There are relatively few strong acids and bases. Since they break apart completely, the [strong acid] = $[H_3O^+]$ and the [strong base]x(subscript of OH) = $[OH^-]$.

Example: Concentration of HCl, a strong acid, is 0.10M so the $[H_3O^+]$ is 0.10M.
Concentration of NaOH, a strong base, is 0.20M so the $[OH^-]$ is 0.20M.
Concentration of $Sr(OH)_2$, another strong base is 0.20M, the $[OH^-]$ is 0.40M.
(I had to multiply the concentration by 2 since the subscript of OH is 2.)

Strong Acids:
HCl, HBr, HI, H_2SO_4, HNO_3, $HClO_4$, some books include $HClO_3$

(All other acids are weak and are in an equilibrium expression. We will deal with that later.)

Strong Bases:
LiOH, NaOH, KOH, RbOH, CsOH, $Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$.

(All other bases are weak and are in an equilibrium expression.)

So for [HCl]= 0.10M then $[H_3O^+]$ is 0.10M. **pH** = -log(0.10) = 1
For [NaOH] = 0.20M then $[OH^-]$ is 0.20M. **pOH** = -log(0.20) = .70
pH = 14-pOH= 14-.7 = 13.3 (We usually want the pH)
For $[Ca(OH)_2$ = 0.20M, then $[OH^-]$ is 0.40M. **pOH** = -log(.40) = .40
pH = 14-pOH= 14-0.40 = 13.6

pH < 7 acidic
pH > 7 basic pH = 7 (neutral)

We will do an exercise together converting Stong Acids and Bases.

[HCl] = 0.050M   Find $[H_3O^+]$, pH, pOH, and $[OH^-]$

pH of HNO_3 is 3.5   Find $[H_3O^+]$, $[HNO_3]$, pOH, $[OH^-]$

pOH of NaOH is 12.9   Find $[OH^-]$, [NaOH], pH, $[H_3O^+]$

pOH of $Ca(OH)_2$ is 11.8   Find $[OH^-]$, $[Ca(OH)_2]$, pH, $[H_3O^+]$