

The $[H^+]$ of a solution with pH=2 is how many times different than the $[H^+]$ of a solution with pH=3?	10 times greater. Remember that $[H^+]$ of pH=2 is <u>greater</u> than pH=3. Also recall that pH is a log scale; every change in pH of 1x represents a change in $[H^+]$ of 10x.
11.2, 11.7, 11.9	
Give the Arrhenius definition of acids and bases.	Acids release $H^+$ in solution. Bases release $OH^-$ in solution.
Distinguish between dissociation and ionization. Give an example of each.	Dissociation refers to the separation of ions (e.g. NaCl is an ionic solid. When dissolved in water its ions come apart). Ionization refers to the creation of ions where none existed before (e.g. HCl is not an ionic compound, yet it reacts with water to form ions: $HCl + H_2O \rightarrow H_3O^+ + Cl^-$ ).
Give the Bronsted definition of acids and bases.	Acids are $H^+$ donors. Bases are $H^+$ acceptors.
Bronsted definitions lead to what idea in acid-base equilibria?	The idea of conjugate acid-base pairs (since if a $H^+$ is transferred in a reaction, it must be transferred back in the reverse reaction).
Identify conjugate acid and base pairs in the reaction: $HCN + H_2O \leftrightarrow H_3O^+ + CN^-$	On the left, HCN donates $H^+$ to $H_2O$ , thus HCN is a Bronsted acid and $H_2O$ is a base. On the right, $H_3O^+$ donates $H^+$ to $CN^-$ , thus $H_3O^+$ is a Bronsted acid (the conjugate of $H_2O$ ) and $CN^-$ is a bronsted base (the conjugate of HCN).
Give the Lewis definition of acids and bases.	Acids are electron pair acceptors. Bases are electron pair donors.

### 15.3

What does diprotic and triprotic mean?	Di- means two, -protic means proton ( $H^+$ ). A diprotic acid (such as $H_2SO_4$ ) is capable of giving off $2H^+$ . $H_3PO_4$ is triprotic. (In $K_a$ or $K_b$ problems we will consider only the first $H^+$ released).
Write the chemical equation for a weak acid in its most general form.	$HA(aq) \leftrightarrow H^+(aq) + A^-(aq)$ , where A represents the anion part of the acid. (This is simplified from $HA + H_2O \leftrightarrow H_3O^+ + A^-$ ).
What is meant by $K_a$ ? Give the equilibrium law for $K_a$ .	$K_a$ is the acid ionization constant. $K_a = \frac{[H^+][A^-]}{[HA]}$
How are $K_a$ problems solved?	Same as other equilibrium problems (using a RICE chart).
How are $K_a$ problems often different than other equilibrium problems?	They involve $H^+$ , thus they often require calculating pH (or they may require calculating $[H^+]$ from pH for use in a RICE chart).
How is acid strength defined according to $K_a$ .	$K_a < 10^{-3}$ Weak acid, $10^{-3} < K_a < 1$ Moderate acid, $K_a > 1$ Strong acid.
Distinguish between strength and concentration (use 0.0001 M HCl as an example).	Strength refers to the degree of ionization. HCl is a strong acid because all HCl molecules dissociate (into $H^+$ , $Cl^-$ ). Concentration refers to moles/volume (described as dilute or concentrated). Thus, 0.0001 M HCl is a dilute solution of a strong acid.

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Write the chemical equation for a weak base in its most general form.	$B + H_2O \leftrightarrow BH^+ + OH^-$ , where B represents the base. (Note: B often has a -ve charge. Thus, $BH^+$ would have a net charge of 0)
What is meant by $K_b$ ? Give the equilibrium law for $K_b$ .	$K_b$ is the base ionization constant. $K_b = \frac{[BH^+][OH^-]}{[B]}$
How are $K_b$ problems solved?	Same as other equilibrium problems (using a RICE chart).
How are $K_b$ problems often different than other equilibrium problems?	They involve $OH^-$ , thus they often require conversions between $H^+$ , pH, $OH^-$ and pOH.
Which bases are the easiest to identify as being strong bases?	All Group IA and IIA hydroxides (e.g. NaOH and $Ca(OH)_2$ ). It is difficult to predict the strength of other metal hydroxides.