Applications and Analogies

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Weak vs Strong Acids and Bases: The Football Analogy

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An important topic in any introductory chemistry course is that of acids and bases. Students generally have no trouble learning the Brønsted–Lowry definition of an acid as a proton donor and a base as a proton acceptor. The equation for the aqueous ionization of an acid (eq 1)

$$(H{-}B)$$
$$HB(aq) \rightleftharpoons H^{+}(aq) + B{:}^{-}(aq) \qquad (1)$$

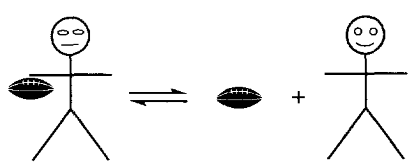
lends itself easily to showing that acids (HB) may dissolve in water to give the protonated, un-ionized form HB(aq), and they may then ionize into $H^+(aq)$ plus the deprotonated conjugate base form, $B{:}^-(aq)$. (Of course acids may be cationic, $HB^+(aq)$, in which case the conjugate base would be neutral, B:(aq).)

From this base of knowledge it is not hard to understand that in aqueous solution, an acid causes an increase in the free proton concentration, which lowers the pH below 7. Likewise, what makes a base a base is its ability to bond to $H^+(aq)$ and lower the proton concentration, thus raising the pH above 7.

Problems often arise, however, when chemistry teachers attempt to explain the difference between weak and strong acids, and between weak and strong bases. For acids in aqueous solution, we often speak of complete vs partial ionization, or $\approx$100% vs <<100% dissociation. This type of terminology works for those of us with a strong grasp of the equilibrium concept, but for many students it does not seem to do the trick. Partial ionization is a difficult concept for some to comprehend; the phrase may not evoke much in the mind of a "visual learner". Visual analogies are often helpful when difficulties like these arise (*1,2*). Although a few analogies relating to acid–base chemistry have been published in this *Journal* (*3, 4*), none address the distinction between strong and weak acids and bases. Accordingly, I have developed a football analogy for acids and bases.

In this analogy we liken an acid, which is a proton donor, to a quarterback: The quarterback is a football "donor", whose job is to deliver the ball by either passing it to a receiver or handing it off to a running back. The "chemical equation" for this process is depicted below (eq 2).

$$HB(aq) \rightleftharpoons H^{+}(aq) + B{:}^{-}(aq)$$



(2)

In this "equation", H^+ is visualized as the football, $B{:}^-$ as the quarterback, and HB as the quarterback holding onto the ball. The quarterback may be found in one of two "states": either holding onto the ball, immediately after the snap from the center (left side of eq 2), or without the ball, after successfully delivering a hand-off or pass (right side of eq 2). Similarly, acids may be found in either the protonated, acidic form (HB), or the deprotonated, conjugate base form ($B{:}^-$).

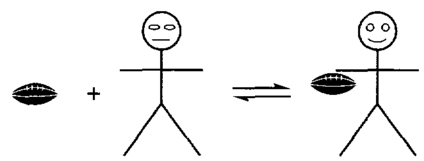
The difference between a strong acid and a weak acid is very much like the difference between an excellent quarterback and a terrible one. A good quarterback delivers the ball efficiently, by hand-off or pass. He (or she) gets rid of the ball, just as a strong acid gets rid of (donates) a proton. At the end of a play, nearly 100% of excellent quarterbacks will have delivered the ball, as represented by the right side of eq 2: ball (gone) plus quarterback. Almost none will still be left holding the ball. Likewise, at equilibrium, nearly 100% of strong acid molecules dissolved in water have delivered a proton, ionizing to $H^+(aq) + B{:}^-(aq)$; almost none remain as protonated/un-ionized HB(aq).

On the other hand, a bad quarterback delivers the ball inefficiently; he (or she) is indecisive and tends to hold onto the ball, just as a weak acid does not readily donate its proton. At the end of a play, far fewer than 100% of bad quarterbacks will have delivered the ball—perhaps only 1% (*really* bad!) or 10% (just plain bad). Using the latter figure, at the end of the play 10% of bad quarterbacks resemble the right side of eq 2: ball (gone) plus quarterback. The other 90% may be found at the end of the play still holding the ball, as on the left side of eq 2. Likewise, at equilibrium, only 10% of weak acid molecules dissolved in water may have delivered a proton, ionizing to $H^+(aq) + B{:}^-(aq)$; the other 90% of the molecules remain as protonated/un-ionized HB(aq).

A good quarterback is more productive in delivering the ball than a bad quarterback; similarly, a strong acid is more likely to deliver a free proton into aqueous solution than a weak acid. As a group, almost 100% of excellent quarterbacks will deliver the ball successfully; on the other hand, perhaps only 10% of bad quarterbacks will deliver the ball successfully. Similarly, one hundred aqueous strong acid molecules (HB(aq)) will successfully release one hundred H^+ ions into solution, whereas 100 aqueous weak acid molecules might only deliver 10 H^+ ions. In this way, a 1 M solution of a strong acid has $[H^+] \approx 1$ M, whereas a 1 M solution of a weak acid has $[H^+] << 1$ M.

A similar analogy may be drawn between a base and a wide receiver. Just as a base is a proton acceptor, a wide receiver is a football "acceptor", whose job is to catch the ball and hold onto it, no matter what. The "chemical equation" for this process is depicted below (eq 3).

$$H^+(aq) + B{:}^-(aq) \rightleftharpoons HB(aq)$$



(3)

Equation 3 is of course just the reverse of eq 2. Like the quarterback, a wide receiver may be found in one of two "states": either holding onto the ball after a successful catch (right side of eq 3), or without the ball, after a drop or fumble (left side of eq 3). Similarly, bases may be either protonated (HB) in their conjugate acid form, or deprotonated ($B{:}^-$) in their basic form. (For a neutral base, B:(aq), the conjugate acid will be cationic, HB^+(aq).)

The difference between a strong base and a weak base is very much like the difference between an excellent wide receiver and a terrible one. A good wide receiver catches the ball and holds onto it, just as a strong base holds onto (accepts) a proton. If the quarterback throws a good pass, nearly 100% of excellent wide receivers will catch the ball, as represented by the right side of eq 3; almost none will have dropped the ball. Likewise, at equilibrium, nearly 100% of strong base molecules dissolved in water have bonded to a proton, yielding protonated HB(aq); almost none remain as ionized H^+(aq) + $B{:}^-$(aq).

On the other hand, a bad wide receiver tends to drop the ball, just as a weak base does not readily bond to a proton. Even if the quarterback throws a good pass, far fewer than 100% of bad wide receivers will catch the ball; perhaps only 1% (*really* bad!) or 10% (just plain bad). Using the latter figure, at the end of the play 10% of bad wide receivers resemble the right side of eq 3: ball caught. The other 90% may be found at the end of the play to have dropped the ball, as on the left side of eq 3. Likewise, at equilibrium, only 10% of weak base molecules dissolved in water have bonded to a proton, yielding protonated HB(aq); the other 90% remain as ionized H^+(aq) + $B{:}^-$(aq).

In this way, viewing bases as H^+ "catchers", strong bases (i.e., excellent wide receivers) are much more likely to catch a proton than are weak bases. Molecule for molecule, they are better at lowering the [H^+] of an aqueous solution than are weak bases. A 1 M solution of a strong base will thus have a much lower [H^+] (i.e., a much higher pH) than a 1 M solution of a weak base.

When discussing acids and bases, it is crucial to stress the importance of water in the reactions involved. For aqueous acids, ionization of HB(aq) to yield H^+(aq) and $B{:}^-$(aq) (eq 1) only takes place after solvation of HB by water. This is sometimes depicted as in eq 1′ below, wherein a specific water molecule accepts the proton from HB.[1]

$$HB(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + B{:}^-(aq) \qquad (1')$$

In a similar fashion, bases may be depicted as in eq 3, bonding to aqueous protons in solution and in this manner lowering [H^+(aq)] and raising pH, or they may be depicted as stripping a proton from a specific water molecule to yield HB(aq) plus the hydroxide anion (eq 3′):

$$H_2O(l) + B{:}^-(aq) \rightleftharpoons HB(aq) + OH^-(aq) \qquad (3')$$

In this scenario, pH rises due to the increase in [OH^-(aq)]. Metal hydroxide salts that are strong bases (e.g., KOH, NaOH) dissolve completely in water, yielding free aqueous metal cations and hydroxide anions, thus causing a direct increase in hydroxide concentration; on the other hand, strong bases like sodium methoxide react completely with H_2O to give methanol plus hydroxide anion. Weak bases like ammonia react only incompletely with H_2O, giving a much lower equilibrium concentration of hydroxide anion relative to the initial ammonia concentration.

Chemists who prefer to use eq 3′ describe bases as reacting with water to remove a proton, creating hydroxide anion. In this scenario, H^+ is still the football, and H_2O, or HOH, may be seen as an offensive player (OH^-) running with the ball (H^+). This player is then tackled by a defensive player ($B{:}^-$), who tries to steal the ball. A strong tackler/ball stealer will frequently end up with the ball, leaving the offensive player (OH^-) bereft of the ball; at the end of the play we have OH^- + HB, just as we do when adding a strong base to water. A weak tackler/ball stealer usually will not end up with the ball, leaving it instead in the hands of the offensive player (OH^-); at the end of the play we have mainly HOH + $B{:}^-$ and only a small amount of OH^- + HB, just as we do when adding a weak base to water.

The great advantage of this analogy is its use of visual imagery to explain the difference between weak and strong acids and between weak and strong bases. The analogy has the rather formidable disadvantage of being based upon a sport about which more than half the population of the United States, and an even higher proportion of the world population, knows relatively little. Even so, the concept of throwing and catching a ball is quite basic, and easy to visualize and comprehend. I have found that the analogy can help even students unfamiliar with the mores of the gridiron to comprehend the mores of aqueous protons.

Note

1. Personally, I find this view both physically inaccurate and needlessly confusing. Single H^+ cations are not solvated by a single H_2O molecule to give H_3O^+; rather, H^+ cations are solvated by a hydration shell consisting of approximately four water molecules. It would thus be more physically accurate to write $H_9O_4^+$, rather than H_3O^+. On the other hand, the term Na^+(aq) is understood in a chemical equation to signify the Na^+ cation complexed to a hydration shell of several water molecules. It seems to me simpler, clearer, and less confusing to introductory chemistry students to understand H^+(aq) in exactly the same way, thus dispensing with the need to show a single H_2O molecule as a proton acceptor.

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