**Unit 14C:** Acids & Bases

### Skills:

* calculate *K*a and *K*b values
* compare acid & base strengths
* perform buffer pH calculations using the Henderson-Hasselbalch equation

###  Notes:

acid dissociation constant (Ka): is the equilibrium constant for the dissociation of an acid. For the “generic” acid HA:



 The greater the *Ka* value, the stronger the acid. (Remember your negative exponents! *E.g.,* 10−5 is *greater* than 10−7.)

p*K*a = −log *K*a  (analogous to pH). The lower (or more negative) the p*K*a, the stronger the acid.

Note that when the acid HA is exactly 50% neutralized, [H+] = [A−], and the above formula reduces to *K*a = [H+]. When this happens, pH = p*K*a.

strong acid: an acid with a pKa lower than that of H3O+ (1.0). Strong acids include HCl, HBr, HI, H2SO4 and HNO3.

 Strong acids dissociate completely into H+ and the corresponding anion. The dissociated H+ converts H2O molecules to H3O+ ions.

base dissociation constant (Kb) is the equilibrium constant for the dissociation of a base. For the “generic” base B:



strong base: a base whose conjugate acid is weaker than H2O (*i.e.,* whose conjugate acid has a pKa higher than 14). Hydroxides are strong bases because they release OH−. However, note that aqueous Mg(OH)2 *acts* more like a weak base because the limited solubility of Mg(OH)2 results in a concentration of OH− that is similar to that produced by a weak base.

 Strong bases either release OH− ions directly into solution, or form OH− ions by pulling H+ off of H2O molecules.

water dissociation constant (*K*w) is the equilibrium constant for the dissociation of H2O into H+ and OH−. At 25°C, *K*w = *K*a·*K*b = 1.008 × 10−14. As with any other equilibrium constant, the value of *Kw* is temperature dependent. Note that this means the pH of pure water is different at different temperatures!

|  |  |  |
| --- | --- | --- |
| Temp. (°C) | *Kw* | pH |
| 0 | 0.114 × 10−14 | 7.47 |
| 10 | 0.293 × 10−14 | 7.27 |
| 20 | 0.681 × 10−14 | 7.08 |
| 25 | 1.008 × 10−14 | 7.00 |
| 30 | 1.471 × 10−14 | 6.92 |
| 40 | 2.916 × 10−14 | 6.77 |
| 50 | 5.476 × 10−14 | 6.63 |
| 100 | 51.3 × 10−14 | 6.14 |

## Weak Acid & Base Solutions

The pH of a weak acid or base solution depends on the p*K*a of the acid, and on the ratio of the concentration of the acid and its conjugate base, which are in equilibrium. If you have more of the acid, the pH will be lower than the p*K*a. If you have more of the base, the pH will be higher than the p*K*a. The resulting pH is given by the Henderson-Hasselbalch equation:



The Henderson-Hasselbalch equation can be derived from the equilibrium expression and the definition of pH.

### Sample problem:

What is the pH of a solution that contains 0.20 *M* NH4+ and 0.10 *M* NH3?

### Solution:

The p*K*a of NH4+/NH3 is 9.25. Therefore, applying the Henderson-Hasselbalch equation:

Weak Acid-Weak Base Solutions

A solution of a weak acid *and* a weak base may be acidic or basic, depending on the relative strengths of the acid *vs.* the base. If the p*K­­a* of the weak acid is larger than the p*Kb* of the weak base, then there will be more of the acid in solution and the pH will be less than 7. If the p*Kb* of the weak base is larger than the p*K­­a* of the weak acid, then there will be more of the base in solution and the pH will be greater than 7. (Remember your negative exponents! 10−5 is *greater* than 10−7.)

Buffers

buffer: a weak acid or base that prevents the pH of a solution from changing drastically until the buffer has been neutralized.

For example, if you have a fish tank, you want to keep the pH from getting too low, you could add NaHCO3. The p*K*a of H2CO3 is approximately 6.4. This means that in the reaction:

H2CO3  H+ + HCO3−

[H+] = [HCO3−] at pH 6.4. As acid accumulates in your fish tank, it will react with the HCO3− ions, and the pH will remain above 6.4 until all of the HCO3− ions have been converted to H2CO3.

Buffers can work in either direction—to “absorb” H+ or OH− ions. If you use a combination of two buffers (one above and one below your desired pH), you can keep the pH within a narrow range.

buffer capacity: the amount of H+ or OH− that the buffer can absorb before it is completely neutralized. This is a stoichiometric calculation. Each mole of weak acid in the buffer can absorb one mole of OH−, and each mole of weak base in the buffer can absorb one mole of H+.

buffering region: the pH range in which the buffer can still absorb acid or base. In the graph below, this is the region between the points where the conjugate base and conjugate acid are exhausted.

 Notice that the p*K*a of the buffer is the point of inflection (as well as the midpoint) of the buffering region.



When choosing a buffer, it is most effective to have the desired pH in the middle of the buffering region. This means the best choice of buffer is the weak acid/base combination whose p*K*a is closest to the desired pH.

Once you have chosen a buffer system, use the desired concentration and the desired pH to calculate the amounts of the acid and base forms, using the Henderson-Hasselbalch equation.

### Sample problem:

You have a 5-person hot tub that holds 1,230 L of water, and you need to keep the pH buffered at 7.20, You want to use an 0.0100 *M* hypochlorous acid/hypochlorite buffer because it will also kill bacteria[[1]](#footnote-1) and keep the water fresh. Hypochlorous acid (HClO) has a p*K*a of 7.40 and a molar mass of . Sodium hypochlorite (NaClO) has a molar mass of . How many grams of hypochlorous acid and how many grams of sodium hypochlorite should you add to the water?

Solution:

We need [acid] + [base] = 0.0100 *M*. Therefore [base] = 0.0100 – [acid].

Using the Henderson-Hasselbalch equation:



Finally, to answer the original question, we need to calculate the number of grams of acid and base needed for 1,230 L of 0.0100 *M* solution:



1. “chlorine bleach” is an 0.0525% sodium hypochlorite solution. [↑](#footnote-ref-1)