# *Ka,* p*K*a, and Buffers

**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 (\_\_\_): is the equilibrium constant for the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ of an acid. For the “generic” acid HA:

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

p*K*a = \_\_\_\_\_\_\_\_\_\_\_\_\_  (analogous to pH). The \_\_\_\_\_\_\_\_\_ (or more negative) the p*K*a, the stronger the acid.

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

strong acid: an acid with a pKa lower than that of H3O+ (1.0). Strong acids include \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

Strong acids \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ into H+ and the corresponding anion. The dissociated H+ converts H2O molecules to \_\_\_\_\_\_\_\_ ions.

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

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

water dissociation constant (*K*w) is the equilibrium constant for the dissociation of H2O into \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. At 25°C, *K*w = *K*a·*K*b = 1.008 × 10−14. As with any other equilibrium constant, the value of *Kw* is \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ 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 \_\_\_\_ 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 \_\_\_\_\_\_, the pH will be lower than the p*K*a. If you have more of the \_\_\_\_\_, 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:

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:

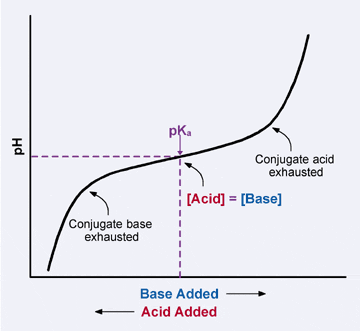
[H+] = [HCO3−] at pH 6.4. As acid accumulates in your fish tank, it will react with the \_\_\_\_\_\_\_\_ ions, and the pH will remain above 6.4 until all of the HCO3− ions have been converted to \_\_\_\_\_\_\_\_\_\_\_.

Buffers can work in either direction—to “\_\_\_\_\_\_\_\_\_” 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 \_\_\_\_\_\_\_\_\_ of H+ or OH− that the buffer can absorb before it is completely \_\_\_\_\_\_\_\_\_\_\_\_\_\_. This is a stoichiometric calculation. Each mole of weak acid in the buffer can absorb one mole of \_\_\_, and each mole of weak base in the buffer can absorb one mole of \_\_\_.

buffering region: the \_\_\_\_\_\_\_\_\_\_\_\_\_ 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 \_\_\_\_\_\_\_\_\_\_\_\_\_\_.

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



When choosing a buffer, it is most effective to have the \_\_\_\_\_\_\_\_\_\_\_\_\_\_ in the middle of the \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. This means the best choice of buffer is the weak acid/base combination whose \_\_\_\_\_ 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)