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This is a clarification to a question asked in the tutorial yesterday. The question was, for assignment 2, question 10(b), why doesn't the sodium ions combine with the free hydroxide ions in the solution.

For assignment 2 question 10(b), we have the reactions:

$$Na(CH_3COO) \rightarrow Na^+ + CH_3COO^-$$
$$H_2O + CH_3COO^- \rightarrow H(CH_3COO) + OH^-$$

The first reaction is the complete dissociation (completely dissolving) of sodium acetate freeing up both sodium and acetate ions. Water will then free up protons *(H^+)* to be accepted by the acetate ions, creating acetic acid and hydroxide ions. So, now we're left with two ions in our solution, sodium *(Na^+)* and hydroxide *(OH^-)* ions. The question was, why won't these two ions combine to create sodium hydroxide *(NaOH)*?

Most metal hydroxides are insoluble, some such as *$Ca(OH)_2$*, *$Mg(OH)_2$*, *$Fe(OH)_2$*, *$Al(OH)_3$* etc are sparingly soluble. However, alkali metal hydroxides *CsOH*, *KOH*, and *NaOH* are very soluble, making them strong bases. When dissolved, these hydroxides are completely ionized. Since the hydroxide concentration, *$[OH^-]$*, is an integrated property of the solution, the solubility of metal hydroxide depends on pH, pOH or *$[OH^-]$*.

In other words, since sodium hydroxide is highly soluble, it will tend to completely dissociate in water. This also means that it is highly unlikely that free sodium ions and free hydroxide ions will combine to form sodium hydroxide in water. For this reason, we consider sodium hydroxide a strong base, because it will allow OH^- ions to exist freely in the solution, resulting in pH > 7.

Note that alkali metals are the metals that sit in the first column of the periodic table of elements (under hydrogen) and have a valency of +1.

I hope this explanation helps. Just remember that alkali metals, such as sodium and potassium tend to exist as free cations in water.

Good luck studying.

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Source: http://www.science.uwaterloo.ca/~cchieh/cact/c123/metalhyd.html