

Answers to Chemistry Self Assessment Test

1. Monobasic sodium phosphate is NaH_2PO_4 , which ionizes in solution to Na^+ and H_2PO_4^- (this form of phosphate is monobasic since it can accept one proton, H^+). Dibasic sodium phosphate is Na_2HPO_4 , which ionizes in solution to two Na^+ and one HPO_4^{2-} (this form of phosphate is dibasic since it can accept two protons).

2. The elements in question have the following molecular weights: Na 23, H 1, P 31 and O 16. Thus, NaH_2PO_4 has a molecular weight of 120, and Na_2HPO_4 has a molecular weight of 142. The molecular weights are in units of grams per mole.

Molar concentration of NaH_2PO_4 :

$$\left(\frac{2.4 \text{ gm}}{5 \text{ mL}}\right)\left(\frac{1 \text{ mole}}{120 \text{ gm}}\right)\left(\frac{1000 \text{ mL}}{1 \text{ L}}\right) = 4.0 \frac{\text{mole}}{\text{L}} = 4.0 \text{ M}$$

Molar concentration of Na_2HPO_4 :

$$\left(\frac{0.90 \text{ gm}}{5 \text{ mL}}\right)\left(\frac{1 \text{ mole}}{142 \text{ gm}}\right)\left(\frac{1000 \text{ mL}}{1 \text{ L}}\right) = 1.3 \frac{\text{mole}}{\text{L}} = 1.3 \text{ M}$$

3. The pH buffer system is formed by H_2PO_4^- acting as a weak acid (proton donor where $\text{H}_2\text{PO}_4^- \rightarrow \text{HPO}_4^{2-} + \text{H}^+$), and HPO_4^{2-} acting as the conjugate base (proton acceptor where $\text{HPO}_4^{2-} + \text{H}^+ \rightarrow \text{H}_2\text{PO}_4^-$).

4. Recall the formula for a hydrogen-ion ($\text{pH} = -\text{Log}[\text{H}^+]$) buffer system:

$$\text{pH} = \text{pK}_a + \text{Log} \frac{[\text{base}]}{[\text{acid}]},$$

so the pH of the solution is

$$\text{pH} = 6.9 + \text{Log} \frac{1.3}{4.0} = 6.4.$$

Thus, the solution is slightly acidic compared to extracellular body fluids; the solution is far less acidic than typical stomach contents, which can have $\text{pH} < 2$.

5. The total concentration of Na^+ in the solution is $4.0 + 2(1.3) = 6.6 \text{ M}$. Given the volume of the dose (3 tablespoons = 9 teaspoons = 45 mL), and the molecular weight of Na (23 grams per mole), we can compute the weight of ingested sodium:

$$\left(\frac{6.6 \text{ moles}}{\text{L}}\right)(45 \text{ mL})\left(\frac{1 \text{ L}}{1000 \text{ mL}}\right)\left(\frac{23 \text{ gm}}{\text{mole}}\right) = 6.8 \text{ gm}$$

6. In the ionized solution, the concentrations are as follows: 6.6 M for Na^+ , 1.3 M for HPO_4^- , and 4.0 M for H_2PO_4^- . Thus, the total osmolarity is $6.6 + 1.3 + 4.0 = 11.9 \text{ Osm}$ (osmoles per liter). The total number of osmoles (i.e., moles of osmotic particles) is:

$$\left(\frac{11.9 \text{ osmoles}}{\text{L}}\right)(45 \text{ mL})\left(\frac{1 \text{ L}}{1000 \text{ mL}}\right) = 0.54 \text{ osmoles}$$

7. We're asking the following: what volume of water will dilute 0.54 osmoles to produce a final osmolarity of $300 \text{ mOsm} = 0.3 \text{ Osm}$? Namely,

$$\left(\frac{0.54 \text{ osmoles}}{X \text{ L}}\right) = 0.3 \text{ Osm}$$

and the answer is $X = 1.8 \text{ L}$.

8. The half gallon (1.8 L) of body water is secreted into the intestines and lost in the toilet. Thus, it's reasonable to assume that this would make you feel quite thirsty.

9. If you made a solution exclusively out of NaH_2PO_4 , the solution would be quite acidic, owing to the reaction $\text{H}_2\text{PO}_4^- \rightleftharpoons \text{HPO}_4^{2-} + \text{H}^+$ (i.e., the monobasic phosphate is acting as a weak acid). However, sodium hydroxide (NaOH) is a strong base, and it can be used to raise the pH ("titrate the solution") to the pH of phospho soda. In so doing, some of the H_2PO_4^- will be converted to HPO_4^{2-} . When you get to a pH of 6.4 (the pH of phospho soda), you will have the correct proportions of the two different forms of phosphate.

The simplest way to construct the solution is to note the molar quantities of the different ionic species. The total amount of phosphate (both forms) is 5.3 M. Thus, you could start by constructing a 5.3 M solution of NaH_2PO_4 (i.e., weigh out $5.3 \times 120 \text{ gm} = 636 \text{ gm}$, place it in a large beaker, and add distilled water to just under 1 L total volume). You then note that the total molar concentration of Na^+ is 6.6 M—with 5.3 M already accounted for by the NaH_2PO_4 . The balance of Na^+ is made up by adding 1.3 moles of NaOH (i.e., add $1.3 \times 40 \text{ gm} = 52 \text{ gm}$ of NaOH , and then bring the final volume up to 1 L).

A better way to make the solution is as follows. Again, start with the 5.3 M solution of NaH_2PO_4 , where the final volume is just under 1 L. Then, insert a pH electrode and simply add NaOH until the pH rises to 6.4 (i.e., titrate the solution). This method ensures that the pH of the final phospho soda solution is correct.

Warning! Although laboratory reagents and chemicals are typically very pure, they are not intended for human consumption—the manufacturer is not required to follow FDA guidelines verifying that they are safe. Furthermore, how do you know that another laboratory worker had not inadvertently contaminated the chemical jar with a highly toxic substance (e.g., used a spatula to spoon out some sodium cyanide, then dip the same spatula into the jar of NaH_2PO_4)? You should never consume laboratory reagents (and this includes sweetening your coffee with laboratory sucrose, salting your sandwich with laboratory NaCl , or taking a drink from the distilled-water carboy)!