

## pH: Strong Acids.

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### A. Introduction

This handout follows Water. (Equations from the Water handout are indicated here with the suffix W.)

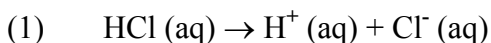
In pure (or neutral) water,  $[H^+] = [OH^-] = 1.0 \times 10^{-7}$  M. (Eq 2-W.) However, addition of some solutes to water causes these numbers to change. Substances that raise  $[H^+]$  are known as acids, substances that raise  $[OH^-]$  are bases. Recall that a change in one of these concentrations causes an inverse change in the other, according to Eq 3-W. And recall that pH is simply a way to denote  $[H^+]$ , Eq 4-W.

One group of acids and bases is particularly simple, the strong acids and bases. We start by looking at a particular example of a strong acid, and then we generalize.

Acids and bases that are not strong are called weak. The next handout is on weak acids and bases.

### B. Strong acids: an example

Consider a 0.10 M aqueous solution of hydrochloric acid.



In solution, HCl ionizes, to form  $H^+$  and  $Cl^-$  ions. In fact, this reaction is essentially 100% complete; the reverse reaction is negligible. (Note that Eq 1 has a unidirectional arrow  $\rightarrow$  rather than a double arrow  $\rightleftharpoons$ .) Further, the balanced equation shows that each mole of HCl produces 1 mol of  $H^+$ .

As a result,  $[H^+] =$  the acid concentration. In this case, the solution is 0.10 M HCl; therefore  $[H^+] = 0.10$  M.

### C. Strong acids: general

Strong acids are acids that are completely ionized. Therefore  $[H^+]$  is readily determined by writing the balanced equation showing ion formation, so you know how many  $H^+$  you get from each acid molecule. In the example of Sect B, the balanced equation (Eq 1) shows 1 mol acid  $\rightarrow$  1 mol  $H^+$ . Therefore  $[H^+] = [\text{acid}]$ .

How do you know which acids are strong? The bad news is that you just need to know the list of strong acids. The good news is that the list is very short, and some members of the list aren't even very important. Further, only one of the strong acids is more complicated than the example discussed above.

The strong acids... Three very common lab acids are strong acids: hydrochloric acid (HCl), nitric acid ( $HNO_3$ ) and sulfuric acid ( $H_2SO_4$ ). You really should know about these three common strong acids. The others: hydrobromic acid (HBr), hydroiodic acid (HI), chloric acid ( $HClO_3$ ), perchloric acid ( $HClO_4$ ).

Those looking for patterns in this list should specifically note that hydrofluoric acid (HF) is not strong. Other common acids that are not strong include phosphoric acid ( $H_3PO_4$ ) and acetic acid ( $CH_3COOH$ ). Some quibbling about whether certain acids belong on the list is possible. The basis of this will become clearer in the next handout.

What, then, should you really know? Well, in the real world you should know about the strength of the acids that you encounter. And you should be able to look up information about acid strength. In a class any statement that you "should know" a certain list is inevitably somewhat arbitrary. I would not emphasize any such list. If you are not sure whether an acid is strong, check the list above, or ask. However, we will talk primarily about "common" acids, and it is reasonable to know which of these are strong and which are not.

All but one of the strong acids give 1  $H^+$  per acid molecule; look at the molecular formulas. We say that they are monoprotic. Thus for all of these, the previous discussion of HCl holds entirely, and  $[H^+] =$  the acid concentration.

The special case is sulfuric acid, which gives off 2  $H^+$ ;  $H_2SO_4$  is diprotic. Unfortunately, this case is a little more complicated than just a factor of two. The first  $H^+$  comes off as a strong acid, but the second  $H^+$  comes off as a weak acid. As a result of this complication, we defer sulfuric acid until the weak acid handout.

### Problems

1. Given 0.01 M  $HNO_3$  (nitric acid), what is  $[H^+]$ ?
2. What is  $[H^+]$  in 1.5 M HCl (hydrochloric acid)?
3. What is the pH of  $10^{-3}$  M  $HClO_4$  (perchloric acid)?
4. What is the pH of 0.025 M HCl?
5. How much HCl (in M) do you need to get  $[H^+] = 0.5$  M?

6. If you want pH = 1.0, what concentration of nitric acid do you need?

#### D. Strong bases

Bases increase  $[\text{OH}^-]$ . Some bases contain the  $\text{OH}^-$  ion, and the ionization directly gives the increased  $[\text{OH}^-]$ . These are the hydroxides.

Strong bases are bases that ionize completely. As with acids, only a few bases are strong.

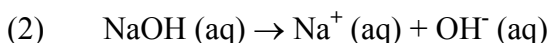
To a good approximation, the strong bases are the hydroxides of Group IA and IIA metals (alkali metals and alkaline earth metals). The most common strong bases in the lab are sodium hydroxide (NaOH) and potassium hydroxide (KOH). Magnesium hydroxide,  $\text{Mg}(\text{OH})_2$ , and calcium hydroxide,  $\text{Ca}(\text{OH})_2$ , are also of some interest.

Working with the strong bases is very similar to working with the strong acids, Sect C. And, of course, recall that  $[\text{H}^+]$ ,  $[\text{OH}^-]$ , and pH are all easily interconverted (Water handout).

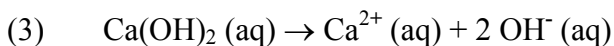
However, the Group IIA hydroxides do require some comment. First, some of them are not very soluble. The small amount that does dissolve ionizes completely (hence, they are strong), but you can't get much in solution. [Magnesium hydroxide is soluble to about  $10^{-4}$  M, calcium hydroxide to  $\sim 0.01$  M. Barium hydroxide,  $\text{Ba}(\text{OH})_2$ , is quite soluble, and strong -- but uncommon.]

More importantly (in terms of principles), the stoichiometry of the group IIA hydroxides is different. Just as you would expect, each formula unit of a group IIA hydroxide gives off two  $\text{OH}^-$  ions. Compare the following two equations with each other, and with Eq 1 for an acid, above.

Typical ionization of a strong base from Group IA:



Typical ionization of a strong base from Group IIA:



Pitfall. In working with Group IIA hydroxides, remember that you get two moles of  $\text{OH}^-$  ions per one mole of the compound.

#### Problems

7. Given 2.0 M NaOH (sodium hydroxide), what is  $[\text{OH}^-]$ ?
8. What is  $[\text{OH}^-]$  in 0.010 M  $\text{Ca}(\text{OH})_2$  (calcium hydroxide)?
9. How much KOH (potassium hydroxide) do you need to get  $[\text{OH}^-] = 2 \times 10^{-3}$  M.
10. Continuing... How much  $\text{Ca}(\text{OH})_2$ ?

\* 11. In the preceding two problems, you might have wondered whether “how much” meant g or mol, rather than M. In fact, you can’t calculate g or mol. Why not?

\* For the following solutions, calculate  $[H^+]$ ,  $[OH^-]$ , and pH.

12. 0.10 M NaOH

13. 0.038 M NaOH

14.  $3.5 \times 10^{-3}$  M  $Ca(OH)_2$

15. How much KOH do you need to reach pH = 12?

16. How much KOH do you need to reach pH = 12.5?

### E. More problems

The problems here generally tie together all the ideas in this handout and Water. You should be able to deal with strong acids and strong bases, and calculate  $[H^+]$  etc.

17. Calculate the pH of 0.10 M  $HNO_3$ .

18. Calculate  $[OH^-]$  for  $2.4 \times 10^{-3}$  M NaOH.

19. Calculate the  $[H^+]$  for 0.15 M NaOH.

20. What is the  $[OH^-]$  in 0.25 M HCl?

21. How much HCl do you need to reach pH = 3.0?

22. Continuing... pH = 2.5?

23. How much KOH do you need to reach pH 10.0?

24. Continuing... How much  $Ca(OH)_2$ ?

25. Continuing... What is  $[H^+]$ ?

26. In 0.040 M  $HNO_3$ , what is  $[OH^-]$ ?

27. What is the pH of 0.10 M  $Ba(OH)_2$  (barium hydroxide)?

28. Calculate the pH of 0.10 M LiOH (lithium hydroxide).

\* 29. If you need 0.31 moles of base, how many grams of KOH would you need?

\* 30. Continuing... How many g of  $Ca(OH)_2$ ?

Although our primary emphasis in this course is doing calculations, it’s also important to know when they are appropriate. Recall occasional “trick” questions in previous handouts, where you were asked to calculate something that shouldn’t be calculated. The following problems are of that type. In fact, the question is: why are the following not calculable (based on current material)?

31. How much HCl do you need to get to pH = 10?

32. What is  $[H^+]$  in 0.1 M acetic acid?

33. What is the pH of 0.1 M  $Al(OH)_3$ ?

34. What is the pH of 1.0 M  $Mg(OH)_2$ ?

### F. What if you have both? [optional]

You add 0.10 M HCl and 0.03 M NaOH to a solution. What is the pH?

0.10 M HCl gives  $[H^+] = 0.10$  M. 0.03 M NaOH gives  $[OH^-] = 0.03$  M. But it’s not possible to have  $[H^+] = 0.10$  M and  $[OH^-] = 0.03$  M. Why not? Recall Eq 3-W. These concentrations

violate that equation.

So what happens? The high concentrations of the two ions drive Eq 1-W backwards. That is, the  $H^+$  and  $OH^-$  ions combine to form water molecules -- until Eq 3-W is satisfied. To a very good approximation... Since we have much more  $H^+$  than  $OH^-$ , "all" the  $OH^-$  will be used up. That will leave 0.07 M  $H^+$  (original 0.10 M less the 0.03 M that was used to react with -- neutralize -- the  $OH^-$ ). Now we know  $[H^+]$ , and can calculate  $[OH^-]$  and pH by the normal procedures.

Caution... If the amount of  $H^+$  and  $OH^-$  are "equal", the approach above won't work. First, it's impossible to have either concentration go to exactly zero. [Why? It would violate Eq 3-W.] Second, it's likely that the difference between the two concentrations (= the remaining amount of the one present in larger quantity) is very small, and beyond the significance of the given concentrations. That may be a little abstract, so let's take an example. You add  $[H^+] = 0.10000$  M and  $[OH^-] = 0.09999$  M. The difference is 0.00001 M (excess  $H^+$ ), so the solution would be pH 5. But if the given concentrations had shown only the usual 2 or 3 SF, then this small difference would not be calculable.

The preceding discussion actually introduces the idea of titrations. In a titration you measure how much acid you have by seeing how much base it takes to neutralize it (or vice versa). "How much", of course, means moles. The fundamental relationship in a titration is moles acid = moles base.

### Problems

35. You add 0.2 M NaOH and 0.1 M HCl. What is the pH?
36. You add 0.03 M NaOH and 0.05 M HCl. What is the pH?
- \* 37. You mix equal volumes of pH 1 HCl and pH 12 NaOH. What is the final pH? Careful.

### G. Normality [optional]

Normality is a unit used to describe the concentration of solutions; it is very similar to molarity. See Sect E of Molarity for a brief introduction.

Normality is most commonly used for acids and bases. The idea is perhaps best introduced with an example. Say that you measure the amount of hydrogen ions in a solution, and find that  $[H^+] = 0.2$  M. (How could you measure this? By seeing how many moles of base, such as NaOH it takes to neutralize the acid.) But this could be due to a 0.2 M solution of HCl or to a 0.1 M solution of  $H_2SO_4$ . The method used to measure  $[H^+]$  does precisely that; in the absence of additional information, you can't tell which of the above possibilities applies.

Look at this another way. The two possibilities -- although quite different solutions -- are equivalent in one very meaningful way: they both provide the same amount of hydrogen ions.

The concentration unit normality (abbreviated N) deals with this situation. We can say that the acid solution is 2 N acid. And we can say that 0.2 M HCl and 0.1 M H<sub>2</sub>SO<sub>4</sub> are equivalent in that they have the same normality.

The form of normality is just as for molarity, except that the numerator is “equivalents”. That is, normality is equivalents of solute per liter of solution. Equivalents is most often used for acids and bases, and refers to the concentration of H<sup>+</sup> or OH<sup>-</sup>, without regard to whether these come one or two per molecule (e.g., HCl vs H<sub>2</sub>SO<sub>4</sub>).

Normality and molarity can be interconverted. All you need to know is the relationship between equivalents and moles. So long as you know the specific chemical, just look at the chemical formula.

Example. HCl contains one hydrogen ion per mole of hydrogen chloride. Thus there is 1 equivalent of H<sup>+</sup> per mole of HCl. And thus for HCl, normality and molarity are numerically equal.

In more detail... A lab bottle says 6 N HCl. What is the concentration in molarity?

$$\frac{6 \text{ equivalents}}{\text{liter}} \times \frac{1 \text{ mole HCl}}{\text{equivalent}} = \frac{6 \text{ moles HCl}}{\text{liter}} \quad \text{or } 6 \text{ M}$$

### Problems

38. Write a conversion factor between equivalents and moles H<sub>2</sub>SO<sub>4</sub>.
39. What is the molarity of 6 N H<sub>2</sub>SO<sub>4</sub>?
40. What is the normality of 6 M H<sub>2</sub>SO<sub>4</sub>?
41. What is the molarity of 0.02 N sodium hydroxide, NaOH?
42. What is the molarity of 0.02 N barium hydroxide, Ba(OH)<sub>2</sub>?
43. What is the normality of 0.1 M phosphoric acid, H<sub>3</sub>PO<sub>4</sub>?
- \* 44. Continuing... What H<sub>2</sub>SO<sub>4</sub> solution, in molarity, would have the same normality?

H. Answers

## Sect C

- |           |          |          |
|-----------|----------|----------|
| 1. 0.01 M | 2. 1.5 M | 3. 3     |
| 4. 1.60   | 5. 0.5 M | 6. 0.1 M |

## Sect D

- |                          |            |                         |
|--------------------------|------------|-------------------------|
| 7. 2.0 M                 | 8. 0.020 M | 9. $2 \times 10^{-3}$ M |
| 10. $1 \times 10^{-3}$ M |            |                         |
11. The only information in the problem is concentration ( $[\text{OH}^-]$ ). To calculate an actual amount, in g or mol, you would need to know the volume of the solution.
12.  $[\text{OH}^-] = 0.1 \text{ M}$ ;  $[\text{H}^+] = 10^{-13} \text{ M}$ ; pH = 13
13.  $[\text{OH}^-] = 0.038 \text{ M}$ ;  $[\text{H}^+] = 2.63 \times 10^{-13} \text{ M}$ ; pH = 12.58
14.  $[\text{OH}^-] = 7.0 \times 10^{-3} \text{ M}$ ;  $[\text{H}^+] = 1.43 \times 10^{-12} \text{ M}$ ; pH = 11.85
- |                         |             |
|-------------------------|-------------|
| 15. $10^{-2} \text{ M}$ | 16. 0.032 M |
|-------------------------|-------------|

## Sect E

- |                                     |                                    |                                     |
|-------------------------------------|------------------------------------|-------------------------------------|
| 17. 1                               | 18. $2.4 \times 10^{-3} \text{ M}$ | 19. $6.7 \times 10^{-14} \text{ M}$ |
| 20. $4.0 \times 10^{-14} \text{ M}$ | 21. $10^{-3} \text{ M}$            | 22. $3.2 \times 10^{-3} \text{ M}$  |
| 23. $10^{-4} \text{ M}$             | 24. $5 \times 10^{-5} \text{ M}$   | 25. $10^{-10} \text{ M}$            |
| 26. $2.5 \times 10^{-13} \text{ M}$ | 27. 13.30                          | 28. 13                              |
29. 17.4 g
30. 11.5 g (Did you get 23.0 g? Think about what the problem really means.)
31. Can't make a basic solution by adding an acid.
32. Acetic acid is not a strong acid; so far we know how to calculate  $[\text{H}^+]$  only for strong acids. (Actual pH is 2.9; next handout.)
33. Not a strong base. (Actually, it's very insoluble in water.)
34. It's a strong base, but quite insoluble. Not possible to have 1.0 M. (If you add 1 mol of magnesium hydroxide to a liter of water, actual pH is ~10.)

## Sect F

- |        |         |
|--------|---------|
| 35. 13 | 36. 1.7 |
|--------|---------|
37. 1.3 (Remember... the mixing reduces each given concentration 2-fold.)

## Sect G

- |  |           |            |
|--|-----------|------------|
| 38. 2 equivalents/mole $\text{H}_2\text{SO}_4$ |           |            |
| 39. 3 M  | 40. 12 N  | 41. 0.02 M |
| 42. 0.01 M                                     | 43. 0.3 N | 44. 0.15 M |