Review:

1. Using your knowledge of the Brønsted-Lowry theory of acids and bases, write equations for the following acid-base reactions and indicate each conjugate acid-base pair.

a) $HNO_3 + OH^- \rightarrow$

b) $CH_3NH_2 + H_2O \rightarrow$

c) $OH^- + HPO_4^{-2} \rightarrow$

Lesson 6.9 review + hwk answers

Lesson 6.9: pH and pOH Calculations of Strong Acids and Bases

Recall that water is amphoteric, meaning that it will behave like an acid or base, and will to a slight extent dissociate.

$$H_2O(l) + H_2O(l) \leftrightarrow H_3O^+(aq) + OH^-(aq)$$

The concentration of hydronium and hydroxide ions present from the dissociation of pure water at a given temperature always multiply to give you a constant. This constant is given the symbol K_w and is often called the **ion product constant.** K_w does not have a given unit.

$$K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$$

This relationship till holds true even if additional hydronium or hydroxide ion is present from the ionization or dissociation of an acid or base.

If $[H_3O^+] > [OH^-]$, the solution is acidic. If $[OH^-] > [H_3O^+]$, the solution is basic. If they are equal (that is, both are 1.0×10^{-7} M – the solution is neutral. A = acidic B = basic N = neutral

Review:

1. Find [H+] for solutions having the following [OH-] value in molarity:

A a) $[OH^-] = 1 \times 10^{-13}$ $\frac{K_w}{[OH^-]} = [H^+]$ $\frac{1.0 \times 10^{-14}}{1 \times 10^{-13}} = 0.1$ M

B b) $[OH^-] = 2.7 \times 10^{-4}$ $\frac{K_w}{[OH^-]} = [H^+] = \frac{1.0 \times 10^{-14}}{2.7 \times 10^{-4}} = 3.7 \times 10^{-11}$ M

B c) $[OH^-] = 1 \times 10^{-3}$ $\frac{K_w}{[OH^-]} = [H^+] = \frac{1.0 \times 10^{-14}}{1 \times 10^{-3}} = 1 \times 10^{-11}$ M

A d) $[OH^-] = 6.3 \times 10^{-10}$ $\frac{K_w}{[OH^-]} = [H^+] = \frac{1.0 \times 10^{-14}}{6.3 \times 10^{-10}} = 1.6 \times 10^{-5}$ M

2. Calculate $[OH^-]$ of a solution when its $[H^+]$ has the following values in molarity:

a) $[H^+] = 1 \times 10^{-3}$ A $\frac{K_w}{[H^+]} = [OH^-] = \frac{1.0 \times 10^{-14}}{1 \times 10^{-3}} = 1 \times 10^{-11}$ M

b) $[H^+] = 3.6 \times 10^{-5}$ A $\frac{K_w}{[H^+]} = [OH^-] = \frac{1.0 \times 10^{-14}}{3.6 \times 10^{-5}} = 2.8 \times 10^{-10}$ M

c) $[H^+] = 1 \times 10^{-2}$ A $\frac{K_w}{[H^+]} = [OH^-] = \frac{1.0 \times 10^{-14}}{1 \times 10^{-2}} = 1 \times 10^{-12}$ M

d) $[H^+] = 7.8 \times 10^{-8}$ B $\frac{K_w}{[H^+]} = [OH^-] = \frac{1.0 \times 10^{-14}}{7.8 \times 10^{-8}} = 1.3 \times 10^{-7}$ M

3. Go through each solution described in questions 1 and 2, and state if it is acidic, basic, or neutral.

see pg 22

The pH scale and pOH scale

K_w is the basis for the pH scale. The pH of a solution is calculated as the negative base 10 logarithm of the hydronium ion concentration $[H_3O^+]$ or $[H^+]$.

$$pH = -\log [H^+] \quad \text{and} \quad [H^+] = 10^{-pH}$$

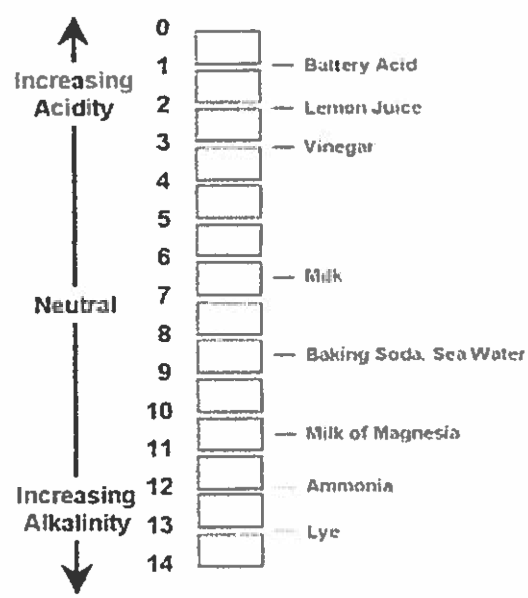
Since $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$ and pure H_2O has $[H^+] = 1.0 \times 10^{-7}$ M. and $[OH^-] = 1.0 \times 10^{-7}$ M. This is why the concentration of a neutral solution is 7 and the range of the pH scale is 0-14. **If $[H^+] > [OH^-]$ the pH is less than 7 and the solution is mostly H^+ and therefore acidic. If $[OH^-] > [H^+]$ the pH is greater than 7 and the solution is mostly OH^- and therefore basic** (alkaline)

The pOH of a solution is calculated as the negative base 10 logarithm of the hydroxide ion concentration $[OH^-]$

$$pOH = -\log [OH^-] \quad \text{and} \quad [OH^-] = 10^{-pOH}$$

Another helpful relationship is: **pH + pOH = 14.**

> Note that the pH and pOH scales go up by powers of 10 so an increase of 1 unit means the concentration of H^+ increases tenfold and the concentration of OH- decreases tenfold. This means something with pH of 3 is 10 times more acidic than something that is pH 4, and something with a pH of 7 is 10 times less basic than something with a pH of 8.
>
> *****When solving pH and pOH problems if the acid or base is a strong acid or base assume it dissociates 100% and plug the concentration [mol / liter] values into the pH and pOH formula. *****



Sig fig rules: when going from [ ] to pH or pOH, the # of decimal places in the pH/pOH value is equal to the # of sig figs in the [ ].
ex. $[H^+] = \underline{1.00} \times 10^{-2}$ M so
$pH = 2.\underline{000}$

Review:

1. Calculate the pH and pOH for each of the following solutions, a) – f). All given values have the unit M.

   a) $[H^+] = 1 \times 10^{-3}$ $pH = -\log[H^+] = -\log(1 \times 10^{-3}) = 3.0$; $pOH = 14 - 3.0 = 11.0$

   b) $[H^+] = 3.6 \times 10^{-5}$ $pH = -\log[H^+] = -\log(3.6 \times 10^{-5}) = 4.44$; $pOH = 14 - 4.44 = 9.56$

   c) $[OH^-] = 2.7 \times 10^{-4}$ $pOH = -\log[OH^-] = -\log(2.7 \times 10^{-4}) = 3.57$;
   $pH = 14 - 3.57 = 10.43$

d) $[OH^-] = 6.3 \times 10^{-10}$ $pOH = -\log[OH^-] = -\log(6.3 \times 10^{-10}) = 9.20$

$pH = 14 - 9.20 = 4.80$

2. For each of the following pH or pOH values, calculate the corresponding [H+] and [OH-].

a) pH = 3.092 $[H^+] = 10^{-3.092} = 8.09 \times 10^{-4}\ M$; $[OH^-] = \frac{Kw}{[H^+]} = \frac{1.0 \times 10^{-14}}{8.09 \times 10^{-4}} = 1.2 \times 10^{-11} M$

b) pH = 8.319 $[H^+] = 10^{-8.319} = 4.80 \times 10^{-9}\ M$; $[OH^-] = \frac{Kw}{[H^+]} = \frac{1.0 \times 10^{-14}}{4.80 \times 10^{-9}} = 2.1 \times 10^{-6} M$

c) pOH = 6.942 $[OH^-] = 10^{-6.942} = 1.14 \times 10^{-7}\ M$; $[OH^-] = \frac{Kw}{[H^+]} = \frac{1.0 \times 10^{-14}}{1.14 \times 10^{-7}} = 8.77 \times 10^{-8} M$

d) pOH = 2.574 $[OH^-] = 10^{-2.574} = 2.67 \times 10^{-3} M$; $[OH^-] = \frac{Kw}{[H^+]} = \frac{1.0 \times 10^{-14}}{3.74 \times 10^{-12}} = 2.7 \times 10^{-3}\ M$

3. Be careful! If you place one drop of hydrochloric acid with a concentration of $[H^+] = 1.0 \times 10^{-2}$ M into a full barrel of plain water, what will be the resulting pH? pH will be $\underline{7}$.

Always write an ionization equation (for acids) or dissociation equation (for bases) if you are not directly given the hydronium or hydroxide ion concentration. Simple stoichiometry will help you figure out the ion concentration you need.

4. What is the pH of a solution that contains 25 grams of hydrochloric acid (HCl) dissolved in 1.5 liters of water?

① Calculate molarity

$$\text{molarity (M)} = \frac{\left(\frac{25g\ HCl \mid 1\ mol\ HCl}{\mid 36.461 g}\right)}{1.5\ L} = .46 M$$

② Write ionization equation for HCl

$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$

According to the mole: mole ratios .46 M HCl yields .46 M of $[H^+]$

③ Calc. pH: $pH = -\log[H^+] = -\log(.46 M) = .337$

5. What is the pH and pOH of a solution that contains 1.32 grams of sodium hydroxide dissolved in 750 mL of water?

① $$\text{molarity (M)} = \frac{\left(\frac{1.32g\ NaOH \mid 1\ mole}{\mid 39.997 g}\right)}{.750 L} = .0330 M$$

② $NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$

According to mole: mole ratios, .0330M NaOH yields .0330M OH^-

③ $pOH = -\log(OH^-) = -\log(.0330M) = 1.481$; $pH = 14 - 1.481 = 12.519$

6. What is the molarity of a calcium hydroxide solution that has a pH of 10.07?

If pH = 10.07, then pOH = 14 − 10.07 = 3.93

$[OH^-] = 10^{-pOH} = 10^{-3.93} = 1.2 \times 10^{-4} M$

$Ca(OH)_2 \longrightarrow Ca^{2+} + 2OH^-$

If $[OH^-] = 1.2 \times 10^{-4}$ M then $[Ca(OH)_2] = \frac{1.2 \times 10^{-4}}{2} = 6.0 \times 10^{-5} M$

7. What is the pH of a solution that contains 10.0 g of nitric acid (HNO_3) and 15.0 g ~~moles~~ of hydrochloric acid (HCl) dissolved in 1000 liters of water? You need to know the molarity of the resulting solution

$$\# \text{moles } HNO_3 = \frac{10.0\text{g } HNO_3 \mid 1 \text{ mole}}{\mid 32.982\text{ g}} = .303 \text{ moles } HNO_3 = .303 \text{ moles } H^+$$

$$\# \text{moles } HCl = \frac{15.0\text{g } HCl \mid 1 \text{ mole}}{\mid 36.461\text{g}} = .411 \text{ moles } HCl = .411 \text{ moles } H^+$$

$$\text{total moles } H^+ = .303 + .411 = .714 \text{ moles}$$

$$\text{molarity} = \frac{.714 \text{ moles}}{1000 \text{ L}} = 7.14 \times 10^{-4} \text{M}, \text{ so } pH = -\log(7.14 \times 10^{-4})$$

$$pH = 3.146$$

Lesson 6.10: Neutralization and Solution Stoichiometry

What happens when an acid is mixed with a base? NEUTRALIZATION!

Products of Neutralization: The products of acid-base neutralization are always a metallic salt and H_2O.

The definition of a metallic salt is a class of compounds formed when the hydrogen ion of an acid is partly or wholly replaced by a metal. In order for neutralization to occur an equal number of moles of acid and base must combine, so that every 1 mole of H^+ combines with exactly 1 mole OH^- to form 1 mole of H_2O. Likewise, 1 mole of negative anion from the acid combines with 1 mole of positive cation from the base to form 1 mole of metallic salt.

It is a misconception that acid-base neutralization *always* yields a pH of 7. The resulting pH is 7 **only if** the number of moles of acid and base are exactly enough to react completely with each other. If there is any excess unreacted acid (or base) left over, it dictates the resulting pH of the solution. Excess unreacted hydronium ion, for example, will make the pH after neutralization be under 7. It is more correct to say that pH after neutralization will be *closer* to 7.

Examples of neutralizations:

$HCl + NaOH \rightarrow NaCl + H_2O$ $\quad$ $H_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4 + 2\,H_2O$ $\quad$ $2\,HNO_3 + Mg(OH)_2 \rightarrow Mg(NO_3)_2 + 2\,H_2O$

Review: Write the balanced neutralization equation for each reaction. Assume that ionization is complete for all acids.

1. Hydrobromic acid and sodium hydroxide

2. Hydroiodic acid and calcium hydroxide

3. Sulfuric acid and rubidium hydroxide

In solution stoichiometry (including acid-base stoichiometry), if you know the molarity and volume of a reactant or product in aqueous solution, you can calculate how many moles of solute are contained in it.

 # moles = molarity (in M) x volume (in L)

From there you can calculate the # of moles of any other reactant or product in the chemical reaction. If precipitation occurs, you can calculate and predict the mass of precipitate that should be recovered.

Lesson 6.9 hwk

2. Write an equation for the dissociation of each of the following bases in water.

a) RbOH

c) $Sr(OH)_2$

b) CsOH

3. Using your knowledge of the Brønsted-Lowry theory of acids and bases, write equations for the following acid-base reactions and indicate each conjugate acid-base pair.

a) $CH_3CH_2NH_2 + H_3O^+ \leftrightarrow$

b) $NH_3 + HSO_4^- \leftrightarrow$

c) $HC_2H_3O_2 + H_2O \leftrightarrow$

d) $OH^- + HNO_3 \leftrightarrow$

Lesson 6.9

1) Complete the following table by calculating the desired quantities. All concentrations are in *M*.

	$[H^+]$	$[OH^-]$	pH	pOH	acid, basic, neutral
a	3.2×10^{-3}	3.1×10^{-12}	2.49	11.51	acid
b	0.56	1.8×10^{-14}	0.26	13.74	acid
c	5.9×10^{-11}	1.7×10^{-4}	10.23	3.77	basic
d	6.0×10^{-8}	1.7×10^{-7}	7.22	6.78	basic
e	0.0050	2.0×10^{-12}	2.30	11.70	acidic
f	9.1×10^{-10}	0.000011	9.04	4.96	basic
g	1.1×10^{-11}	9.1×10^{-4}	10.97	3.03	basic
h	1.0×10^{-7}	9.8×10^{-8}	6.99	7.01	acidic (barely)

2. What would be the pH and pOH of each of the following solutions? pH in black / pOH in green

a) 0.0010 M HCl 3.00 11.00

b) 0.0010 M HNO_3 3.00 11.00

c) 0.010 M $HClO_4$ 3.00 11.00

g) 1.0 M $Ca(OH)_2$ $[OH^-] = 2.0$ M 14.30 −.30

h) 0.075 M KOH 12.88 1.12

i) 0.000034 M $Ba(OH)_2$ $[OH^-] = .000068$ M 9.83 4.17