Name ______________________ Section ________ Date _________

Chapter 16: Basic Review Worksheet

1. Compare the Arrhenius and Bronsted-Lowry definitions of an acid and a base.

2. Describe the relationship between a conjugate acid-base pair in the Bronsted-Lowry model.

3. Write balanced chemical equations showing the following molecules behaving as Bronsted-Lowry acids in water: HCl, H_2SO_4.

4. How is the *strength* of an acid related to the *position* of its ionization equilibrium?

5. Write the equations for the dissociation (ionization) of HCl, HNO_3, and $HClO_4$ in water.

6. Explain how water is an *amphoteric* substance.

7. What values does K_w have at 25°C? What are $[H^+]$ and $[OH^-]$ in pure water at 25°C?

8. How does $[H^+]$ compare to $[OH^-]$ in an acidic solution? How does $[H^+]$ compare to $[OH^-]$ in a basic solution?

9. How is the pH scale defined?

10. What range of pH values corresponds to acidic solutions? What range corresponds to basic solutions?

11. When the pH of a solution changes by one unit, by what factor does the hydrogen ion concentration change in the solution?

12. How is pOH defined?

13. Describe a *buffered* solution. Why is buffering so important in biological systems?

14. Write the conjugate acid of each of the following bases.
    a. SO_3^{2-}
    b. HS^-
    c. F^-
    d. CH_3COO^-

15. Write the conjugate base of each of the following acids.
    a. H_2SO_4
    b. H_2S
    c. H_2CO_3
    d. $HC_2H_3O_2$

16. For each of the given items, calculate the indicated quantity.
    a. $[H^+] = 4.01 \times 10^{-3}\ M$, pH = ?
    b. $[OH^-] = 7.41 \times 10^{-8}\ M$, pOH = ?
    c. $[H^+] = 9.61 \times 10^{-6}\ M$, pOH = ?

17. Calculate the pH and pOH values for each of the following solutions.
    a. 0.00141 M HNO_3
    b. $2.13 \times 10^{-3}\ M$ NaOH

18. A 25.0-mL sample of 0.50 M HCl is titrated to the endpoint with 12.4 mL of NaOH. Calculate the concentration of the NaOH solution.

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Chapter 16: Review Worksheet

1. Write balanced chemical equations showing the following molecules/ions behaving as Bronsted-Lowry acids in water: H_3PO_4, NH_4^+.

2. Acetic acid is a weak acid in water. What does this indicate about the affinity of the acetate ion for protons compared to the affinity of water molecules for protons? If a solution of sodium acetate is dissolved in water, the solution is basic. Explain. Write equilibrium reaction equations for the ionization of acetic acid in water and for the reaction of the acetate ion with water in a solution of sodium acetate.

3. Are aqueous solutions of NaCl, $NaNO_3$, or $NaClO_4$ acidic, basic, or neutral? Explain.

4. Write the chemical equation for the autoionization of water. Write the expression for the equilibrium constant, K_w, for this reaction.

5. Why is pH = 7.00 considered *neutral*?

6. How are pH and pOH for a given solution related? Explain.

7. Give three examples of buffered solutions. For each of your examples write equations that explain how the components of the buffered solution consume added strong acids or bases.

8. Write the conjugate acid of each of the following bases.
   a. HSO_4^-
   b. HSO_3^-
   c. H_2O
   d. S^{2-}

9. Write the conjugate base of each of the following acids.
   a. H_3PO_4
   b. HS^-
   c. HCO_3^-
   d. NH_3

10. For each of the given items, calculate the indicated quantity.
    a. $[OH^-] = 6.62 \times 10^{-3}\ M$, pH = ?
    b. pH = 6.325, $[OH^-]$ = ?
    c. pH = 9.413, $[H^+]$ = ?

11. Calculate the pH and pOH values for each of the following solutions.
    a. 0.00515 M HCl
    b. $5.65 \times 10^{-5}\ M$ $Ca(OH)_2$

12. You titrate a 50.0-mL sample of HCl with 0.10 M NaOH. The titration requires 23.8 mL of the base. Calculate the concentration of the HCl solution.

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Chapter 16: Challenge Review Worksheet

1. For each of the following species, write the equation for the reaction of each species with water if it acts like an acid, and if it acts like a base.
   a. HS^-
   b. HCO_3^-
   c. HSO_4^-
   d. $H_2PO_4^-$
   e. HSO_3^-

2. At 35°C, the value of K_w is 2.09×10^{-14}. Determine the pH of water at this temperature. What does the term neutral mean at 35°C?

3. Determine the pH of a solution made by dissolving 1.0×10^{-12} mol HCl is dissolved in enough water to make 1.0 L of solution.

4. You are titrating a 50.0 mL sample of HNO_3 with 0.20 *M* NaOH. After adding 45.3 mL of the base, you remember that you have not added phenolphthalein indicator. Upon adding the indicator the solution turns bright pink. You decide to "back-titrate" by adding 0.10 *M* HCl to the solution. You reach the stoichiometric point by adding 11.6 mL of the 0.10 *M* HCl. Determine the concentration of the original HNO_3 solution.

$$169.91\text{g NaCl} \times \frac{1\text{ mol NaCl}}{58.44\text{g NaCl}} = 2.91\text{ mol NaCl}$$

Using the net ionic equation,

	Ag^+	+	Cl^-	→	AgCl
Initial	1.00 mol		2.91 mol		
Change	- 1.00 mol		-1.00 mol		
End	0		1.91 mol		

When the reaction is complete, we have
0 mol Ag^+
1.00 mol NO_3^-
2.91 mol Na^+
1.91 mol Cl^-

Since our total volume is 2.00 L,
$[NO_3^-] = 0.500\text{ M}$
$[Na^+] = 1.46\text{ M}$
$[Cl^-] = 0.955\text{ M}$

7. Freezing point depression depends on the number of particles.

$$10.0\text{g NaCl} \times \frac{1\text{ mol NaCl}}{58.44\text{g NaCl}} \times \frac{2\text{ mol ions}}{1\text{ mol NaCl}} = 0.342\text{ mol ions (particles)}$$

$$0.342\text{ mol } C_6H_{12}O_6 \times \frac{180.156\text{g}}{1\text{ mol } C_6H_{12}O_6} = 61.6\text{g glucose}$$

Chapter 16: Basic Review Worksheet

1. Arrhenius defined acids and bases in terms of the ions that show acidic and basic properties in aqueous solution: an acid is a substance that produces H^+ ions in aqueous solution, and a base is a substance that produces OH^- ions in aqueous solution. This definition is too restrictive, since it only considers aqueous systems. For example, Arrhenius recognized HCl as an acid when it is dissolved in water, but what about gaseous hydrogen chloride? Or how about HCl dissolved in some other solvent? The Arrhenius theory also only allows for only one kind of base, the hydroxide ion.

   The Bronsted-Lowry model for acids and bases extends some of the concepts of the Arrhenius theory and adapts them to more general situation. Acids and bases are defined in a more fundamental manner. In the Bronsted-Lowry model, an acid still represents a source of H^+ ions (specifically, a Bronsted-Lowry acid is a proton donor), but there is no restriction as to the solvent being only water. So in the Bronsted-Lowry model, if HCl transfers H^+ to another species, then HCl is an acid regardless of any other consideration. Where the Bronsted-Lowry theory really extends our idea of acid-base chemistry is in defining what represents a base: a

Bronsted-Lowry base is any species that accepts a proton from an acid. According to the Bronsted-Lowry theory, the hydroxide ion in aqueous solution is a base, because if an acid is added to a hydroxide ion solution, the hydroxide ions would accept protons from the acid

$$HCl + OH^- \rightarrow H_2O + Cl^-$$

The Bronsted-Lowry theory allows for other bases, however, and does not require water to be the solvent. For example, ammonia and hydrogen chloride react with each other in the gas phase.

$$HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$$

This is an acid-base reaction because in the Bronsted-Lowry model a proton has been transferred from the acid (HCl) to the base (NH_3) to form NH_4^+.

2. A conjugate acid-base pair consists of two species related to each other by the donating or accepting of a single proton, H^+. An acid has one more H^+ than its conjugate base; a base has one less H^+ than its conjugate acid.

3. $HCl(aq) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)$

   $H_2SO_4(aq) + H_2O(l) \rightarrow HSO_4^-(aq) + H_3O^+(aq)$

4. The strength of an acid is a direct result of the position of the acid's ionization equilibrium. We call an acid for which the ionization equilibrium position lies far to the right a strong acid, and we call an acid whose equilibrium position lies only slightly to the right a weak acid. For example, HCl, HNO_3, and $HClO_4$ are all strong acids, which means that they are completely ionized in aqueous solution (the position of the ionization equilibrium is very far to the right).

5. $HCl(aq) + H_2O(l) \rightarrow Cl^-(aq) + H_3O^+(aq)$

   $HNO_3(aq) + H_2O(l) \rightarrow NO_3^-(aq) + H_3O^+(aq)$

   $HClO_4(aq) + H_2O(l) \rightarrow ClO_4^-(aq) + H_3O^+(aq)$

6. When we say that water is an amphoteric substance, we are just recognizing that water will behave as a Bronsted-Lowry base if a strong acid is added to it, but will behave as a Bronsted-Lowry acid if a strong base is added to it. For example, water behaves as a base when HCl is dissolved in it

   $$HCl + H_2O \rightarrow Cl^- + H_3O^+$$

   However, water would behave as an acid if the strong base $NaNH_2$ were added to it

   $$NH_2^- + H_2O \rightarrow NH_3 + OH^-$$

7. The equilibrium constant K_w has the value 1.0×10^{-14} at 25°C. Because of the fact that hydronium ions and hydroxide ions are produced in equal numbers when water molecules undergo autoionization, and with the value for the equilibrium constant given, we know that in pure water $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ M

8. If a solution has a higher concentration of H_3O^+ ion than OH^- ion, we say the solution is acidic. If a solution has a lower concentration of H_3O^+ ion than OH^- ion we say the solution is basic.

9. The pH of a solution is defined as the negative of the base 10 logarithm of the hydrogen ion concentration in the solution; that is $pH = -\log[H^+]$.

10. Solutions in which the hydrogen ion concentration is greater than 1.0×10^{-7} M (pH < 7.00) are acidic; solutions in which the hydrogen ion concentration is less than 1.0×10^{-7} M (pH > 7.00) are basic.

11. Since the pH scale is logarithmic, when the pH changes by one unit, this corresponds to a change in the hydrogen ion concentration by a factor of ten.

12. In some instances, it may be more convenient to speak directly about the hydroxide ion concentration present in a solution, and so an analogous logarithmic expression is defined for the hydroxide ion concentration: $pOH = -\log[OH^-]$

13. A buffered solution is one that resists a change in its pH even when a strong acid or base is added to it. Buffered solutions consist of approximately equal amounts of two components: a weak acid (or base) and its conjugate base (or acid). The weak acid component of the buffered solution is capable of reacting with added strong base. The conjugate base component of the buffered solution is able to react with added strong acid. By reacting with (and effectively neutralizing) the added strong acid or strong base, the buffer is able to maintain its pH at a relatively constant level.

    Buffered solutions are crucial to the reactions in biological systems because many of these reactions are extremely pH dependent: a change in pH of only one or two units can make some reactions impossible or extremely slow. Many biological molecules have 3-dimensional structures that are extremely dependent on the pH of their surroundings. For example, protein molecules can lose part of their necessary structure and shape if the pH of their environment changes (if a protein's structure is changed it may not work correctly). For example, if a small amount of vinegar is added to whole mile, the milk instantly curdles: the solid which forms is the protein portion of the milk, which becomes less soluble at lower pH values.

14. a. HSO_3^-
    b. H_2S
    c. HF
    d. CH_3COOH

15. a. HSO_4^-
    b. HS^-
    c. HCO_3^-
    d. $C_2H_3O_2^-$

16. a. $[H^+] = 4.01 \times 10^{-3}$ M, $pH = -\log[4.01 \times 10^{-3}] = 2.397$
    b. $[OH^-] = 7.41 \times 10^{-8}$ M, $pOH = -\log[7.41 \times 10^{-8}] = 7.130$
    c. $[H^+] = 9.61 \times 10^{-6}$ M, $pH = -\log[9.61 \times 10^{-6}] = 5.017$, $pOH = 14.00 - 5.017 = 8.98$

17. a. $[HNO_3] = [H^+] = 0.00141$ M, pH = -log(0.00141) = 2.851, pOH = 14.00 – 2.851 = 11.15
    b. $[NaOH] = [OH^-] = 2.13 \times 10^{-3}$ M, pOH = -log[2.13×10^{-3}] = 2.672, pH = 14.00 – 2.672 = 11.33

18. The equation is $HCl + NaOH \rightarrow H_2O + NaCl$

$$25.0\ \text{mL} \times \frac{0.50\ \text{mol HCl}}{1000\ \text{mL}} \times \frac{1\ \text{mol NaOH}}{1\ \text{mol HCl}} = 0.0125\ \text{mol NaOH}$$

$$\frac{0.0125\ \text{mol NaOH}}{0.0124\ \text{L}} = 1.01\ \text{M NaOH}$$

Chapter 16: Review Worksheet

1. $H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_2PO_4^-(aq) + H_3O^+(aq)$

   $NH_4^+(aq) + H_2O(l) \rightleftharpoons NH_3(aq) + H_3O^+(aq)$

2. When we say that acetic acid is a weak acid, we can take either of two points of view. Usually we say that acetic acid is a weak acid because it doesn't ionize very much when dissolved in water: we say that not very many acetic acid molecules dissociate. However, we can describe this situation from another point of view. We could say that the reason acetic acid doesn't dissociate much when we dissolve it in water is because the acetate ion (the conjugate base of acetic acid) is extremely effective at holding on to protons, and specifically is better at holding on to protons than water is in attracting them.

   $$HC_2H_3O_2 + H_2O \rightleftharpoons H_3O^+ + C_2H_3O_2^-$$

   Now what would happen if we had a source of free acetate ions (for example, sodium acetate) and placed them into water? Since acetate ion is better at attracting protons than is water, the acetate ions would pull protons out of water molecules, leaving hydroxide ions. That is,

   $$C_2H_3O_2^- + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-$$

   Since an increase in hydroxide ion concentration would take place in the solution, the solution would be basic. Because acetic acid is a weak acid, the acetate ion is a base in aqueous solution.

3. Since HCl, HNO_3, and $HClO_4$ are all strong acids, we know that their anions (Cl^-, NO_3^-, ClO_4^-) must be very weak bases, and that solutions of the sodium salts of these anions would not be basic. Since these acids have a strong tendency to lose protons, there is very little tendency for the anions (bases) to gain protons.

4. $H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$

   Because this reaction is so important to our understanding the relative acidity and basicity of aqueous solution, the equilibrium constant for the autoionization of water is given a special symbol, $K_w = [H_3O^+][OH^-]$.

5. Since in pure water, the amount of $H^+(aq)$ ion present is equal to the amount of $OH^-(aq)$ ion, we say that pure water is neutral. Since $[H^+] = 1.0 \times 10^{-7}$ M in pure water, this means that the pH of pure water is $-\log[1.0 \times 10^{-7} \text{ M}] = 7.00$.

6. The concentrations of hydrogen ion and hydroxide ion in water (and in aqueous solutions) are not independent of one another, but rather are related by the dissociation equilibrium constant for water,

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25°C.$$

From this constant we can see that pH + pOH = 14.00 for water (or an aqueous solution) at 25°C.

7. Answers will vary, but consider the following three buffered solutions:

$0.10 \text{ M } HC_2H_3O_2/0.10 \text{ M } NaC_2H_3O_2$
$0.50 \text{ M } HF/0.50 \text{ M } KF$
$0.25 \text{ M } NH_4Cl/0.25 \text{ M } NH_3$

The acidic component of each of these buffered solutions can neutralize added OH^- ion as shown below:

$$HC_2H_3O_2 + OH^- \rightarrow C_2H_3O_2^- + H_2O$$

$$HF + OH^- \rightarrow F^- + H_2O$$

$$NH_4^+ + OH^- \rightarrow NH_3 + H_2O$$

Notice that in each case, the added hydroxide ion has been neutralized and converted to a water molecule. The basic component of the buffered solutions can neutralize added H^+ ions as shown below:

$$C_2H_3O_2^- + H^+ \rightarrow HC_2H_3O_2$$

$$F^- + H^+ \rightarrow HF$$

$$NH_3 + H^+ \rightarrow NH_4^+$$

Notice that in each case the added hydrogen ion is converted to some other species.

8. a. H_2SO_4
   b. H_2SO_3
   c. H_3O^+
   d. HS^-

9. a. $H_2PO_4^-$
   b. S^{2-}
   c. CO_3^{2-}
   d. NH_2^-

10. a. $[OH^-] = 6.62 \times 10^{-3}$ M, pOH = $-\log[6.62 \times 10^{-3}] = 2.179$, pH = 14.00 – 2.179 = 11.82
    b. pH = 6.325, pOH = 14.00 – 6.325 = 7.68, $[OH^-]$ = {inv}{log}(-7.68) = 2.1×10^{-8}
    c. pH = 9.413, $[H^+]$ = {inv}{log}(-9.413) = 3.86×10^{-10}

11. a. $[HCl] = [H^+] = 0.00515$ M, pH = -log(0.00515) = 2.288, pOH = 14 – 2.288 = 11.71
    b. $[OH^-] = 2 \times 5.65 \times 10^{-5}$ M $= 1.13 \times 10^{-4}$ M,
    pOH = -log$[1.13 \times 10^{-4}]$ = 3.947, pH = 14.00 – 3.947 = 10.05

12. The equation is $HCl + NaOH \rightarrow H_2O + NaCl$

$$23.8\text{ mL} \times \frac{0.10\text{ mol NaOH}}{1000\text{ mL}} \times \frac{1\text{ mol HCl}}{1\text{ mol NaOH}} = 0.00238\text{ mol HCl}$$

$$\frac{0.00238\text{ mol HCl}}{0.0500\text{ L}} = 0.0476\text{ M HCl}$$

Chapter 16: Challenge Review Worksheet

1. a. acid: $HS^- + H_2O \rightleftharpoons S^{2-} + H_3O^+$
   base: $HS^- + H_2O \rightleftharpoons H_2S + OH^-$
   b. acid: $HCO_3^- + H_2O \rightleftharpoons H_3O^+ + CO_3^{2-}$
   base: $HCO_3^- + H_2O \rightleftharpoons H_2CO_3 + OH^-$
   c. acid: $HSO_4^- + H_2O \rightleftharpoons SO_4^{2-} + H_3O^+$
   base: $HSO_4^- + H_2O \rightleftharpoons H_2SO_4 + OH^-$
   d. acid: $H_2PO_4^- + H_2O \rightleftharpoons H_3O^+ + HPO_4^{2-}$
   base: $H_2PO_4^- + H_2O \rightleftharpoons H_2PO_4^- + H_3O^+$
   e. acid: $HSO_3^- + H_2O \rightleftharpoons SO_3^{2-} + H_3O^+$
   base: $HSO_3^- + H_2O \rightleftharpoons H_2SO_3 + OH^-$

2. $[H^+][OH^-] = 2.09 \times 10^{-14}$

   $[H^+] = [OH^-] = 1.45 \times 10^{-7}$

   pH = 6.84

   The solution is neutral because $[H^+] = [OH^-]$. The pH varies with temperature because the value of K_w varies with temperature.

3. The answer is 7.00. Many students will claim that the answer is 12.00 because $[HCl] = [H^+] = 1.0 \times 10^{-12}$ M, and -log(1.0×10^{-12}) = 12.00. However, we must realize that water is also in the solution and contributes 1.0×10^{-7} M H^+. While it is true that the additional H^+ from the HCl affects the autoionization of water equilibrium, the amount of H^+ from the HCl is too small to have a detectable effect. Students should also see that it doesn't make sense that the pH becomes basic when a tiny amount of a strong acid is added to water.

4. The equation is $HNO_3 + NaOH \rightarrow NaNO_3 + H_2O$

   Thus, moles acid = moles base (1:1 ratio)

   $$\text{Total moles base} = 45.3\text{ mL} \times \frac{0.20\text{ mol NaOH}}{1000\text{ mL}} = 0.00906\text{ mol base}$$

$$\text{mol HCl} = 11.6\text{ mL} \times \frac{0.10\text{ mol HCl}}{1000\text{ mL}} = 0.00116\text{ mol HCl}$$

$$0.00906\text{ mol (base)} = \text{mol } HNO_3 + 0.00116\text{ mol (HCl)}$$

$$\text{mol } HNO_3 = 0.0079\text{ mol } HNO_3$$

$$\frac{0.0079\text{ mol } HNO_3}{0.0500\text{ L}} = 0.158\text{ M } HNO_3$$

Chapter 17: Basic Review Worksheet

1. Chemists envision that a reaction can only take place between molecules if the molecules physically collide with each other.

2. The activation energy for a reaction represents the minimum energy the reactant molecules must possess for a reaction to occur when the molecules collide.

3. A catalyst is a substance that speeds up a reaction without being consumed. Biological catalysts are called enzymes.

4. Chemists define equilibrium as the balancing of two exactly opposing processes. When a chemical reaction is started by combining pure reactants, the only process possible initially is

   reactants → products

   However as the concentration of product molecules increases, it becomes more and more likely that product molecules will collide and react with each other

   products → reactants

   giving back molecules of the original reactants. At some point in the process the rates of the forward and reverse reactions become equal, and the system attains chemical equilibrium.

5. Consider this example: suppose we have a reaction for which K = 4, and we begin this reaction with 100 reactant molecules. At the point of equilibrium, there should be 80 molecules of product and 20 molecules of reactant remaining (80/20 = 4). Suppose we perform another experiment involving the same reaction, only this time we begin the experiment with 500 molecules of reactant. This time, at the point of equilibrium, there will be 400 molecules of product present and 100 molecules of reactant remaining (400/100 = 4). Since we began the two experiments with different numbers of reactant molecules, it's not troubling that there are different absolute numbers of product and reactant molecules present at equilibrium: however, the ratio, K, is the same for both experiments. We say that these two experiments represent two different positions of equilibrium: an equilibrium position corresponds to a particular set of equilibrium concentrations that fulfill the value of the equilibrium constant. Any experiment that is performed with a different amount of starting material will come to its one unique equilibrium position, but the equilibrium constant ratio, K, will be the same for a given reaction regardless of the starting amounts taken.