Name ____________________ Section ________ Date ________

Chapter 15: Basic Review Worksheet

1. Define a *solution*.

2. Describe how an ionic solute such as NaCl dissolves in water to form a solution. How does a molecular solid such as sugar dissolve in water?

3. List three ways to increase the rate of dissolution of a solute.

4. Define a *saturated* solution, an *unsaturated* solution, and a *supersaturated* solution.

5. Define *percent by mass*, *molarity*, and *normality*.

6. What is one *equivalent* of an acid? What does an equivalent of a base represent?

7. What is a one *normal* solution of an acid or a base?

8. A 12.5 g sample of glucose ($C_6H_{12}O_6$) is dissolved in 225 g of water. Calculate the percent by mass of glucose in the solution.

9. A chemist prepares some standard solutions for use in the lab using 500.0-mL volumetric flasks to contain the solutions. If the following masses of solutes are used, calculate the resulting molarity of each solution.
   a. 4.865 g NaCl　　b. 78.91 g $AgNO_3$

10. Suppose each of the following solutions is diluted by adding the indicated amount of water. Calculate the new concentrations of the solutions.
    a. 255 mL of 3.02 *M* HCl; 375 mL water added
    b. 75.1 g of 1.51% $AgNO_3$; 125 g water added

11. Calculate the volume (in mL) of each of the following acid solutions that would be required to neutralize 36.2 mL of 0.259 *M* NaOH solution.
    a. 0.271 *M* HCl　　b. 0.119 *M* H_2SO_4　　c. 0.171 *M* H_3PO_4

12. What volume of 0.242 *M* H_2SO_4 can furnish the same number of moles of H^+ ions as each of the following?
    a. 41.5 mL of 0.118 *M* HCl
    b. 27.1 mL of 0.121 *M* H_3PO_4

13. Calculate the normality of the following solutions.
    a. 0.204 *M* HCl
    b. 0.328 *M* H_2SO_4

14. What volume of 0.10 *M* $Ba(NO_3)_2$ solution is required to react completely with 100.0 mL of 0.50 *M* NaCl solution to form $BaCl_2(s)$?

15. What is meant by the term *colligative property*?

Name ______________________ Section ________ Date _________

Chapter 15: Review Worksheet

1. How are the strong bonding forces in a crystal of ionic solute overcome to allow the solid to dissolve in water?

2. What forces between water molecules and the molecules of a molecular solid tend to help the solute dissolve?

3. Does the term *saturated* for a solution mean the same thing as *concentrated*? Explain.

4. When a solution is diluted by adding additional solvent, the *concentration* of solute changes but the *amount* of solute present does not change. Explain.

5. Suppose 250.0 mL of water is added to 125.0 mL of 0.551 *M* NaCl solution. Explain *how* you would calculate the concentration of the solution after dilution.

6. How is the equivalent weight of an acid or a base related to the substance's molar mass?

7. Give an example of an acid and a base that have equivalent weights *equal* to their molar masses. Give an example of an acid and a base that have equivalent weights that are *not equal* to their molar masses.

8. How is the *normality* of an acid or a base solution related to its *molarity*?

9. Suppose that 4.25 g of NaCl and 7.50 g of KCl are both dissolved in 125 g of water. Calculate the percent by mass of each component of the solution.

10. Calculate the molarity of a solution in which enough water is added to 250.0 g of calcium chloride to make a solution with a volume of 2.25 L.

11. To prepare 125 mL of 0.100 *M* sulfuric acid (H_2SO_4) in the lab requires 0.69 mL of concentrated acid. What is the concentration of concentrated sulfuric acid?

12. Silver chromate (Ag_2CrO_4) is a blood red ionic solid which has very low solubility in water. What mass of silver chromate can be produced by mixing 100.0 mL of 0.100 *M* silver nitrate and 75.0 mL of 0.100 *M* sodium chromate?

13. Calculate the volume (in mL) of 0.104 *M* HCl that would be required to neutralize 25.2 mL of 0.00491 *M* barium hydroxide solution.

14. What volume (in mL) of 0.50 *N* NaOH is required to neutralize 15.0 mL of 0.35 *N* H_2SO_4?

15. Explain in your own words why adding a solute to water raises its boiling point.

Name ______________________________ Section ________ Date __________

Chapter 15: Challenge Review Worksheet

1. Why do some substances *not* dissolve in water to any appreciable extent?

2. Why does a solute dissolve only to a particular extent in water?

3. How does formation of a saturated solution represent an equilibrium?

4. The concentration of a solution may be expressed in various ways. Two means of expressing concentration were introduced in Chapter 15 – mass percent and molarity. How are these two concentration expressions the same, and how do they differ? Suppose 5.0 g of NaCl is dissolved in 15.0 g of water to give a solution volume of 16.1 mL. Explain how you would calculate the mass percent of NaCl and the molarity of NaCl in this solution. Which number did you *not* use for the mass percent calculation? Which number did you *not* use for the molarity calculation? Suppose instead of being given the volume of the solution after mixing, you had been given the density of the solution. Explain how you would calculate the molarity of this solution with this information.

5. Give an example of a solution whose normality is equal to its molarity, and an example of a solution whose normality is *not* the same as its molarity.

6. You make two solutions, one containing silver nitrate and one containing sodium chloride. You make these separate solutions by adding equal masses of the solid to enough water to make 1.0-L of each solution. You then mix these solutions together and note the formation of a white solid. You collect and dry the solid, and you find the mass to be 143.35 g. Determine the concentration of all of the ions remaining in the solution.

7. What mass of glucose ($C_6H_{12}O_6$) must be dissolved in 1.0 L of water so that the solution has the same freezing point as a solution containing 10.0 g of NaCl in 1.0 L of water?

15. The simple model we use to explain many properties of metallic elements is called the electron sea model. In this model we picture a regular lattice array of metal cations in a "sea" of mobile valence electrons. The electrons can move easily to conduct heat or electricity through the metal, and the lattice of cations can be deformed fairly easily, allowing the metal to be hammered into a sheet or stretched to make a wire.

16. Substitutional alloys consist of a host metal in which some of the atoms in the metal's crystalline structure are replaced by atoms of other metallic elements of comparable size to the atoms of the host metal. For example, sterling silver consists of an alloy in which approximately 7% of the silver atoms have been replaced by copper atoms. Brass and pewter are also substitutional alloys. An interstitial alloy is formed when other smaller atoms enter the interstices (holes) between atoms in the host metal's crystal structure. Steel is an interstitial alloy in which carbon atoms occupy the interstices of a crystal of iron atoms. The presence of the interstitial carbon atoms markedly changes the properties of the iron, making it much harder and tougher. Depending on the amount of carbon introduced into the iron crystal, the properties of the resulting steel can be carefully controlled.

Chapter 14: Challenge Review Worksheet

1. The fact that the boiling point of water is so much higher than that of the other covalent hydrogen compounds of the Group 6 elements is evidence for the special strength of hydrogen bonding (it takes more energy to vaporize water because of the extra strong forces between the molecules in the liquid state).

2. The high heat of vaporization of water is essential to life on earth, since much of the excess energy striking the earth from the sun is dissipated in vaporizing water.

3. A simple experiment to determine vapor pressure is shown in Figure 14.11. Samples of a liquid are injected into a sealed tube containing mercury. Since mercury is so dense, the liquids float to the top of the mercury where they evaporate. As the vapor pressures of the liquids develop to the saturation point, the level of mercury in the tube changes as an index of the magnitude of the vapor pressures.

4. The stronger the interparticle forces in a solid, the higher the melting and boiling points.

5. N_2 is a more ideal gas than CO. The carbon monoxide molecule is polar, whereas N_2 is a nonpolar molecule. Thus, the intermolecular forces for CO are greater than those for N_2. One of the premises of the kinetic molecular theory is that the gas particles exert no forces on each other. This is more true for N_2 than for CO thus, N_2 is a more ideal gas than CO.

Chapter 15: Basic Review Worksheet

1. A solution is a homogeneous mixture, a mixture in which the components are uniformly intermingled.

2. In a crystal of sodium chloride there is a negative chloride ion at one of the corners of the crystal. When this crystal is placed in water, water molecules surround the chloride ion, and orient themselves with the positive end of their dipoles aimed at the negative chloride ion. When

enough water molecules have so arranged themselves, the resultant attraction of the several water molecules for the chloride ion becomes stronger than the attractive forces from the positive sodium ions in the crystal, and the chloride ion separates from the crystal and enters the solution (still surrounded by the group of water molecules). Similarly, a positive sodium ion in a similar position would be attracted by a group of water molecules arranged with the negative ends of their dipoles oriented toward the positive ion. When enough water molecules are present to surpass the attractive forces from negative ions in the crystal, the sodium ion enters the solution. Once the chloride ion and the sodium ion are in solution, they remain surrounded by water molecules (called a hydration sphere), which diminishes the effective charge each ion would feel from the other and which prevents them from easily recombining. For a molecular solid (such as sugar) to dissolve in a solvent, there must be some portion or portions of the molecule that can be attracted by molecules of solvent. For example, common table sugar (sucrose) contains many hydroxyl groups, -OH. These hydroxyl groups on a sugar molecule are attractive to water molecules (naturally, the –OH groups can also hydrogen bond with water). When the attractive forces between water molecules and solute molecules are strong enough, the solid dissolves.

3. To increase the rate of dissolution we can: heat the solution, stir the solution, increase the surface area of the solute (in all cases we expose more solute to fresh solvent).

4. A saturated solution is one that contains as much solute as can dissolve at a particular temperature. An unsaturated solution is one that contains less solute than can dissolve at a particular temperature. A supersaturated solution is one in which more solute is dissolved than is normally possible at a particular temperature.

5. Percent by mass (or mass percent) is defined as (mass of solute)/(mass of solution) x 100%. Molarity is defined as the number of moles of solute per liter of solution. Normality is defined as the number of equivalents per liter of solution.

6. One equivalent of an acid is the amount of acid that can furnish one mole of H^+ ions. One equivalent of a base is the amount of a base that can furnish one mole of OH^- ions.

7. The normality of a solution is defined as the number of equivalents of solute contained in one liter of the solution: a 1 N solution of an acid contains 1 mole of H^+ per liter; a 1 N solution of a base contains 1 mole of OH^- per liter.

8. 12.5 g glucose/237.5 g solution x 100% = 5.26% glucose.

9. a. 500.0 mL = 0.5000 L

   molar mass NaCl = 58.44 g

   $$4.865\text{ g NaCl} \times \frac{1\text{ mol}}{58.44\text{ g NaCl}} = 0.08325\text{ mol NaCl}$$

   $$M = \frac{0.08325\text{ mol NaCl}}{0.5000\text{ L}} = 0.1665\text{ M}$$

b. molar mass $AgNO_3$ = 169.9 g

$$78.91\text{ g }AgNO_3 \times \frac{1\text{ mol}}{169.9\text{ g }AgNO_3} = 0.4644\text{ mol }AgNO_3$$

$$M = \frac{0.4644\text{ mol }AgNO_3}{0.5000\text{ L}} = 0.9288\text{ M}$$

10. a. $\frac{(3.02\text{ M})(255\text{ mL})}{(255+375)\text{ mL}} = 1.22\text{ M}$

b. $\frac{((1.51\%)(75.1\text{ g}))}{(75.1+125)\text{ g}} = 0.567\%$

11. It would be convenient to first calculate the number of moles of NaOH present in the sample, since this information will be needed for each part of the answer.

36.2 mL = 0.0362 L

$$\text{mol NaOH} = 0.0362\text{ L} \times \frac{0.259\text{ mol NaOH}}{1\text{ L}} = 9.38 \times 10^{-3}\text{ mol NaOH}$$

a. $HCl + NaOH \rightarrow NaCl + H_2O$

$$9.38 \times 10^{-3}\text{ mol NaOH} \times \frac{1\text{ mol HCl}}{1\text{ mol NaOH}} = 9.38 \times 10^{-3}\text{ mol HCl}$$

$$9.38 \times 10^{-3}\text{ mol HCl} \times \frac{1\text{ L}}{0.271\text{ mol HCl}} = 0.0346\text{ L} = 34.6\text{ mL}$$

b. $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

$$9.38 \times 10^{-3}\text{ mol NaOH} \times \frac{1\text{ mol }H_2SO_4}{2\text{ mol NaOH}} = 4.69 \times 10^{-3}\text{ mol }H_2SO_4$$

$$4.69 \times 10^{-3}\text{ mol }H_2SO_4 \times \frac{1\text{ L}}{0.119\text{ mol }H_2SO_4} = 0.0394\text{ L} = 39.4\text{ mL}$$

c. $H_3PO_4 + 3NaOH \rightarrow Na_3PO_4 + 3H_2O$

$$9.38 \times 10^{-3}\text{ mol }H_3PO_4 \times \frac{1\text{ mol }H_3PO_4}{3\text{ mol NaOH}} = 3.13 \times 10^{-3}\text{ mol }H_3PO_4$$

$$3.13 \times 10^{-3}\text{ mol }H_3PO_4 \times \frac{1\text{ L}}{0.171\text{ mol }H_3PO_4} = 0.0183\text{ L} = 18.3\text{ mL}$$

12. a. $\frac{(41.5\text{ mL})(0.118\text{ M})(1)}{(0.242\text{ M})(2)} = 10.1\text{ mL } H_2SO_4$

b. $\frac{(27.1\text{ mL})(0.121\text{ M})(3)}{(0.242\text{ M})(2)} = 20.3\text{ mL } H_2SO_4$

13. a. 0.204 M HCl = 0.204 N HCl
    b. 0.328 M H_2SO_4 = 0.656 N H_2SO_4

14. The balanced equation is $Ba(NO_3)_2 + 2NaCl \rightarrow BaCl_2 + 2NaNO_3$

$$\frac{0.50\text{ mol NaCl}}{1000\text{ mL}} \times \frac{1\text{ mol Ba(NO}_3)_2}{2\text{ mol NaCl}} \times \frac{1000\text{ mL}}{0.10\text{ mol Ba(NO}_3)_2} = 250.0\text{ mL } 0.10\text{ M Ba(NO}_3)_2$$

15. A colligative property is a property of a solution that depends on the number of particles in solution, not the nature (or identity) of the particles.

Chapter 15: Review Worksheet

1. When an ionic substance is dissolved in water to form a solution, the water plays an essential role in overcoming the strong interparticle forces in the ionic crystal. Water is a highly polar substance: one end of the water molecule dipole is partially negative, and the other is partially positive.

2. In order for a substance to dissolve, the molecules of the substance must be capable of being dispersed among water molecules. (See the discussion of question #2, Chapter 15: Basic Review Worksheet)

3. To say that a solution is saturated does not necessarily mean that the solute is present at a high concentration. For example, magnesium hydroxide only dissolves to a very small extent before the solution is saturated, whereas it takes a great deal of sugar to form a saturated solution (the saturated sugar solution is extremely concentrated).

4. Adding additional solvent to a solution so as to dilute the solution does not change the number of moles of solute present but only changes the volume in which the solute is dispersed. If we are using the molarity of the solution to describe its concentration, the number of liters is changed when we add solvent, and the number of moles per liter (the molarity) changes, but the actual number of moles of solute does not change.

5. For example, 125 mL of 0.551 M NaCl contains 68.9 millimol of NaCl. The solution will still contain 68.9 millimol of NaCl after the 250 mL of water is added to it, only now the 68.9 millimol of NaCl will be dispersed in a total volume of 375 mL. This gives the new molarity as 68.9 mmol/375 mL = 0.184 M. The volume and the concentration have changed, but the number of moles of solute in the solution has not changed.

6. The equivalent weight of an acid or base is related to the molar mass of the substance, by taking into account how many H^+ or OH^- ions the substance furnishes per molecule.

7. HCl and NaOH have equivalent weights equal to their molar masses, since each of these substances furnishes one H^+ or OH^- ion per unit, respectively

$$HCl \rightarrow H^+ + Cl^- \qquad NaOH \rightarrow Na^+ + OH^-$$

However, sulfuric acid (H_2SO_4) has an equivalent weight that is half the molar mass, since each H_2SO_4 molecule can produce two H^+ ions. Therefore, only half a mole of H_2SO_4 is needed to provide one mole of H^+ ion. Similarly, the equivalent weight of phosphoric acid (H_3PO_4) is one third of its molar mass, since each H_3PO_4 molecule can provide three H^+ ions (and so only one third mole of H_3PO_4 is needed to provide one mole of H^+ ions).

$$H_2SO_4 \rightarrow 2H^+ + SO_4^{2-} \qquad H_3PO_4 \rightarrow 3H^+ + PO_4^{3-}$$

Similarly, bases like $Ca(OH)_2$ and $Mg(OH)_2$ have equivalent weights that are half their molar masses, since each of these substances produces two moles of OH^- ion per mole of base (and so only half a mole of base is needed to provide one mole of OH^-).

8. Since the equivalent weight and the molar mass of a substance are related by small whole numbers (representing the number of H^+ or OH^- a unit of the substance furnishes), the normality and molarity of a solution are also simply related by these same numbers. In fact, $N = n \times M$ for a solution where n represents the number of H^+ or OH^- ions furnished per molecule of solute.

9. mass of solution = 4.25 g + 7.50 g + 125 g = 136.75 g (137 g)

$$\frac{4.25\text{ g NaCl}}{136.75\text{ g}} \times 100 = 3.11\ \%\text{NaCl} \qquad \frac{7.50\text{ g KCl}}{136.75\text{ g}} \times 100 = 5.48\ \%\text{ KCl}$$

10. $250.0\text{g } CaCl_2 \times \dfrac{1\text{ mol } CaCl_2}{110.98\text{g } CaCl_2} = 2.253\text{ mol } CaCl_2$

$$\frac{2.253\text{ mol}}{2.25\text{ L}} = 1.00\text{ M}$$

11. $125\text{ mL} \times \dfrac{0.100\text{ mol}}{1000\text{ mL}} = 0.0125\text{ mol } H_2SO_4$

$$\frac{0.0125\text{ mol } H_2SO_4}{0.00069\text{ L}} = 18\text{ M } H_2SO_4$$

12. The balanced equation is

$$2AgNO_3(aq) + Na_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + 2\,NaNO_3(aq)$$

First, we need to find which reactant is limiting:

$$100.0 \text{ mL} \times \frac{0.100 \text{ mol AgNO}_3}{1000 \text{ mL}} \times \frac{1 \text{ mol Ag}_2\text{CrO}_4}{2 \text{ mol AgNO}_3} \times \frac{331.8\text{g Ag}_2\text{CrO}_4}{1 \text{ mol Ag}_2\text{CrO}_4} = 1.66\text{g Ag}_2\text{CrO}_4$$

$$75.0 \text{ mL} \times \frac{0.100 \text{ mol Na}_2\text{CrO}_4}{1000 \text{ mL}} \times \frac{1 \text{ mol Ag}_2\text{CrO}_4}{1 \text{ mol Na}_2\text{CrO}_4} \times \frac{331.8\text{g Ag}_2\text{CrO}_4}{1 \text{ mol Ag}_2\text{CrO}_4} = 2.49\text{g Ag}_2\text{CrO}_4$$

Thus, 1.66g Ag_2CrO_4 can be produced (the amount of $AgNO_3$ is limiting).

13. The balanced equation is $2\ HCl + Ba(OH)_2 \rightarrow BaCl_2 + 2\ H_2O$

$$25.2 \text{ mL} \times \frac{0.00491 \text{ mol Ba(OH)}_2}{1000 \text{ mL}} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Ba(OH)}_2} \times \frac{1000 \text{ mL}}{0.104 \text{ mol HCl}} = 2.38 \text{ mL HCl}$$

14. The balanced equation is $H_2SO_4 + 2\ NaOH \rightarrow Na_2SO_4 + 2\ H_2O$

$$15.0 \text{ mL} \times \frac{0.35\text{eq. H}_2\text{SO}_4}{1000 \text{ mL}} \times \frac{1 \text{ eq. NaOH}}{1 \text{ eq. H}_2\text{SO}_4} \times \frac{1000 \text{ mL}}{0.50 \text{ eq NaOH}} = 10.5 \text{ mL } 0.50 \text{ N NaOH}$$

15. This is explained in the text, but the students should phrase the answer in their own words. Basically adding a solute decreases the vapor pressure of the water (students should include their own phrasing of the "bubbles" that form in the water). Recall as well that the boiling point is the point at which the vapor pressure of the solution is equal to the atmospheric pressure. Thus, at 100°C, the vapor pressure of an aqueous solution is less than 1 atm, so the solution must be heated to a higher temperature in order for the solution to boil.

Chapter 15: Challenge Review Worksheet

1. If the water-solute interactions are not comparable to the water-water interactions, the substance will not dissolve. (See the discussion of question #2, Chapter 15: Basic Review Worksheet)

2. A saturated solution is one which is in equilibrium with undissolved solute: as molecules of solute dissolve from the solid in one place in the solution, dissolved molecules rejoin the solid phase in another place in the solution.

3. As with the development of vapor pressure above a liquid, formation of a solution reaches a state of dynamic equilibrium: once the rates of dissolving and "undissolving" become equal, there will be no further net change in the concentration of the solution and the solution will be saturated.

4. The mass percent and the molarity are similar in that both methods of expressing the concentration of a solution represent ratios: that is, they both express the amount of solute per unit of solution. The mass percent for a solution represents the number of grams of solute that would be present in 100. g of the solution. Since the mass percent is based only on mass, it is invariant for a given solution under all conditions. The molarity of a solution represents the

number of moles of solute that would be present in one liter of the solution. Since the volume of a liquid varies slightly with temperature, the molarity of a solution varies also with temperature.

If 5.0 g of NaCl were dissolved in 15.0 g of water, the mass percent composition of the solution could be calculated as

$$\frac{5.00\text{ g NaCl}}{(15.0\text{ g }H_2O + 5.0\text{ g NaCl})} \times 100\% = 25\%\text{ NaCl}$$

Since the mass percent is based only on the masses of the components of the solution, the volume of the solution is not needed for this calculation.

If 5.0 g of NaCl (molar mass 58.4 g) were dissolved in enough water to give a total solution volume of 16.1 mL, the molarity of the solution could be calculated as

$$5.0\text{ g NaCl} \times \frac{1\text{ mol}}{58.4\text{ g}} = 0.0856\text{ mol NaCl}$$

$$16.1\text{ mL} = 0.0161\text{ L}$$

$$M = \frac{0.0856\text{ mol NaCl}}{0.0161\text{ L solution}} = 5.3\text{ M}$$

Since the molarity is based on the amount of solute per liter of solution, the mass of solvent present is not needed for the calculation. If the density (1.24 g.mL) of the solution had been given, rather than the explicit volume, the volume could be calculated as shown

$$V = m/d = \frac{(5.0\text{ g} + 15.0\text{ g})}{1.24\text{ g/mL}} = \frac{20.0\text{ g}}{1.24\text{ g/mL}} = 16.1\text{ mL}$$

5. For example, a 0.521 M HCl solution is also 0.521 N, since each HCl furnishes one H^+ ion. However, a 0.475 M H_2SO_4 solution would have a normality equal to

   $$N = n \times M = 2 \times 0.475\text{ M} = 0.950\text{ N}$$

   since each H_2SO_4 molecule furnishes two H^+ ions.

6. The relevant balanced equation is

   $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

   $$143.35\text{g AgCl} \times \frac{1\text{ mol AgCl}}{143.35\text{g AgCl}} = 1\text{ mol AgCl}$$

   Thus, 1 mol AgCl is formed. Therefore, we must have at least 1 mole $AgNO_3$ (169.91g) and 1 mol NaCl (58.44g). Since we are told we have equal masses of each reactant, we must have 169.91g of each.

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

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,
$[NO_3^-] = 0.500$ M
$[Na^+] = 1.46$ M
$[Cl^-] = 0.955$ 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