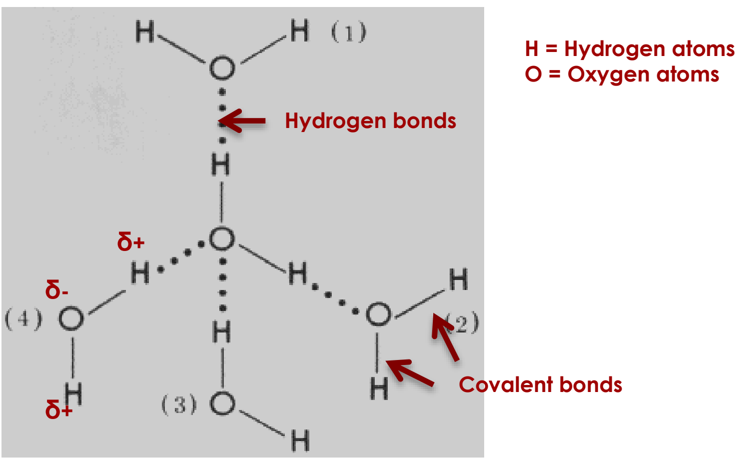
CHEMISTRY CP Name:_____ **KEY** ______________________ Period:_____
TEST DATE: _________________

Unit 8 Review Sheet KEY:
Properties of Water, Solutions, Concentration, Acids and Bases

PROPERTIES OF WATER

1. Define the following terms: *polarity, surface tension, vapor pressure, specific heat, and capillary action.*
   POLARITY: **Polarity is the separation of charges, positive and negative that can describe a bond or an entire molecule and is caused by differences in electronegativity.**
   SURFACE TENSION: **The tendency for molecules at the surface of a liquid to be pulled inward resulting in a smooth surface.**
   VAPOR PRESSURE: **The vapor pressure of a liquid is the equilibrium pressure of a vapor above its liquid (or solid)**
   SPECIFIC HEAT: **The amount of energy needed to raise the temperature of 1g of substance by 1°C.**
   CAPILLARY ACTION: **The rise of liquids up a narrow tube**
2. Draw four water molecules. Label the types of bonds, oxygen atoms, hydrogen atoms, and respective charges on the atoms.



3. Is water polar or nonpolar? Explain. **Water is a polar molecule because the oxygen is more electronegative that the hydrogen so there is unequal sharing of the bonding electrons. Oxygen attracts the bonding electrons more than each hydrogen, and this causes the oxygen to be slightly negative and each hydrogen becomes slightly positive.**
4. What type of bond forms between individual molecules of water? **Hydrogen bonds**
5. What type of bond forms between each hydrogen and the oxygen within a water molecule? **Covalent**
6. Why is water considered the universal solvent? **Water is considered the universal solvent because it has the ability to dissolve many substances.**
7. What are the special properties of water and why do they occur? **High surface tension, high specific heat, low vapor pressure, capillarity, less dense in the solid state; these properties are due to the hydrogen bonds**
8. Explain why solid ice is less than liquid water with regard to particle arrangement. **Ice actually has a very different structure than liquid water, in that the molecules align themselves in a regular lattice rather than more randomly as in the liquid form. It happens that the lattice arrangement allows water molecules to be more spread out than in a liquid, and, thus, ice is less dense than water.**
9. Why does sugar dissolve in water, but oil does not? **Water is a polar molecule and sugar is polar (like dissolves like). Oil is nonpolar so it will not dissolve.**

SOLUTIONS

10. Define the following terms: *solution, solvent, solute, dilute, concentrated, dissociate, solubility, saturated, supersaturated and unsaturated.*
    SOLUTION: **a homogeneous mixture where one substance is dissolved inside of another.**
    SOLVENT: **the substance that does the dissolving**
    SOLUTE: **the substance that is dissolved**
    DILUTE: **a solution that has excess solvent; the solution has a lower concentration of solute per solvent**
    CONCENTRATED: **a solution that contains more a higher concentration of solute per solvent**
    DISSOCIATE: **when ionic compounds break into their respective ions completely**
    SOLUBILITY: **the measure of the amount of solute that can be dissolved in a given amount of solvent**

SATURATED: **a solution where the maximum of solute is added to solvent**
SUPERSATURATED: **a solution where there are more solute particles than are needed to form a saturated solution**
UNSATURATED: **a solution where there are less solute particles than are needed to form a saturated solution**

11. Give an example of solid, liquid, and gas solution. Identify the solute and solvent.
Solid: Steel. Solute-carbon, Solvent-iron
Liquid: Soda. Solute-sugar, CO_2, etc. Solvent-water
Gas: Air. Solute-O_2, CO_2, etc. Solvent-N_2
12. What is a solution? Give an example of a solution, and an example of a mixture that is not a solution.
Solution is a homogeneous mixture. Examples of solutions include steel, Kool-Aid, and air.
A mixture that is not a solution is cereal and milk.
13. Describe how temperature and pressure affects the solubility of solid and gas solutes in water.
As temperature increases, the solubility of solid solutes increase and the solubility of gas solutes decrease. As pressure increases, the solubility of gas solutes increase. Pressure does not affect the solubility of solid solutes.
14. A glass of water has 10g of sugar dissolved in it. If more sugar can be added to dissolve in the water, is the solution unsaturated, saturated, or supersaturated? **unsaturated**
15. How do intermolecular forces affect solvation? **Like dissolves like: ionic and polar solutes dissolve in polar solvents, nonpolar solutes dissolve in nonpolar solvents. The energy released by forming intermolecular bonds between the solutes and solvent needs to be greater than the energy it takes to break apart the bonds.**
16. Why do vinegar and oil not mix? **Differences in polarity; vinegar is polar and oil is nonpolar**

17. *On the line at the left, write the letter of the definition that best matches each term.*

_**f**__ solution	a. measure of how much solute will dissolve in a solvent
_**c**_ solute	b. solution with water as the solvent
_**g**_ solvent	c. substance that is dissolved in a solution
_**i**__ soluble	d. substance that dissolves in water to form a solution that conducts electricity
_**e**_ alloy	e. solid solution containing two or more metals
_**b**_ aqueous	f. homogeneous mixture of two or more substances in a single physical state
_**d**_ electrolyte	g. substance that does the dissolving in a solution
_**a**_ solubility	h. Liquids that are insoluble in each other are considered this
_**h**_ immiscible	i. capable of being dissolved

CALCULATING CONCENTRATION

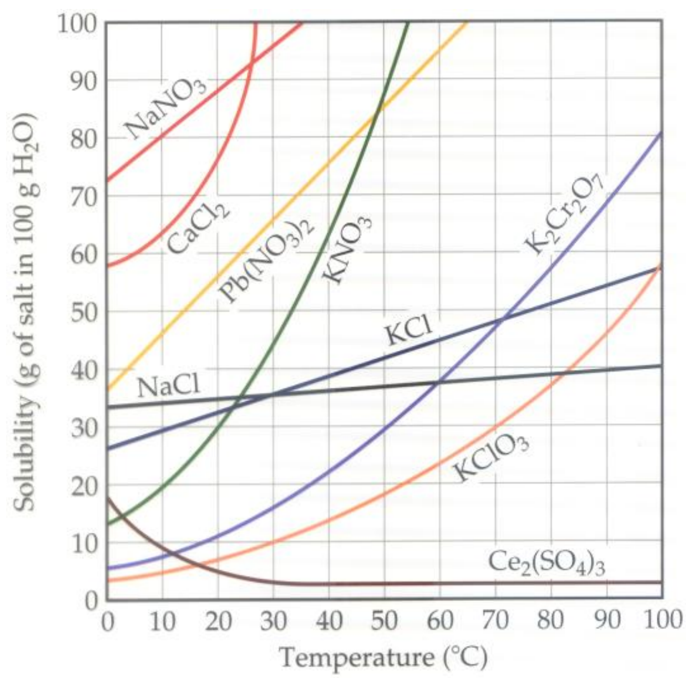
COMMON CONVERSIONS: 1 g = 1 mL for water 1 kg = 1000 g 1 L = 1000 mL

EQUATIONS: (1) Molarity= Moles solute/liters solution
(2) Molality= Moles solute/kg solvent
(3) ppm = (grams of solute/g solution) x 1,000,000
(4) % mass = (grams of solute/ grams of solution) x 100
(5) % volume = (volume of solute/ volume of solution) x 100
(6) grams per liter = grams of solute/liter of solution

18. If 8.7 g of Na_2CO_3 is dissolved in 800 mL of water, what is the **molarity** of the solution?
8.7 g / 106 g = 0.0821 moles Molarity = 0.0821moles / 0.800 L = 0.103*M*
19. How many **grams** of $MgCl_2$ would be needed to make 1.5 L of a **0.40 M** solution?
Molarity = moles/Liter
0.40*M* = x/1.5 L x = 0.60 moles
0.60 moles x 95.2 g/M = 57.12 grams
20. What is the **molarity** of a bleach solution containing 9.5g of NaOCl per liter of bleach?
9.5 g / 74.4 g = 0.128 moles Molarity = 0.128 moles / 1 L = 0.128*M*
21. What is the **percent by volume** of ethanol in a solution that contains 35 mL of ethanol dissolved in 115 mL of water?
% volume = (35 mL / (35 mL + 115 mL)) x 100 = 23.3%
22. Calculate the **molarity** of 1.60L of a solution containing 1.55 g of dissolved KBr.
1.55 g / 119 g = 0.0130 moles Molarity = 0.0130 moles / 1.60 L = 0.00814*M*

23. If .5mL of blood are added to 10.0 L of water what is the concentration in **PPM?**
    .5 mL = . 5 g   10.0 L = 10.0 kg or 10000 g
    (0.5 g/10000 g) x 1,000,000 = 50 ppm
24. A solution is made up of 123 g NaOH and 289 g water. The total volume is 300.0 mL. Determine the following:
    a. Moles of NaOH **123 g/ 40 g = 3.075 moles**
    b. Moles of H_2O **289 g/ 18 g = 16.1 moles**
    c. Mass Percent **(123 g / 412 g) x 100 = 29.9 %**
    d. Mole fraction **X_{NaOH} = 3.075 moles / 19.175 moles = 0.160**
       X_{H2O} = 1.00 - 0.160 = 0.84
    e. Molarity **3.075 moles / .300 L = 10.25 *M***
    f. PPM **(123 g / 412 g) x 1000000 = 2.99×10^5 ppm**
25. What is the **molality** of a solution that contains 63.0 g HNO_3 in 0.500 kg H_2O?
    63.0 g / 63.0 g = 1 mole   molality = 1 mole / 0.500 kg = 2*m*
26. What **mass of water** is required to dissolve 100. g NaCl to prepare a **1.50*m*** solution?
    100 g / 58.5 g = 1.71 moles
    molality = moles / kg
    1.50*m* = 1.71 moles / x kg   x = 2.57 kg

SOLUBILITY **(NOTE CHANGES TO #28 and #30)**



27. What is the solubility of potassium nitrate at 30^0 C?
    Potassium nitrate, KNO_3
    44 g per 100 g H_2O
28. How many grams of ammonia can I dissolve in 200 grams of water at a temperature of 45^0 C? **SKIP, NH_3 IS NOT ON THIS GRAPH****
29. At what temperature is the solubility of sodium chloride the same as the solubility of potassium chloride? **About 30 degrees C**
30. How many grams of **potassium** chloride would I need to make 300 grams of a saturated solution at 70^0 C? **48 g x 3 = 144 g**
31. What do all of the compounds that decreased in solubility over the temperature range in the graph have in common? **They are gases**
32. What compound is least soluble at 40^0 C? **$Ce_2(SO_4)_3$**
33. What ionic compound is least soluble at 40^0 C? **$Ce_2(SO_4)_3$**
34. Using the solubility graph, determine if the following solutions are saturated, unsaturated or supersaturated. If they are anything but saturated, list two things you can do to make them saturated (include numbers).

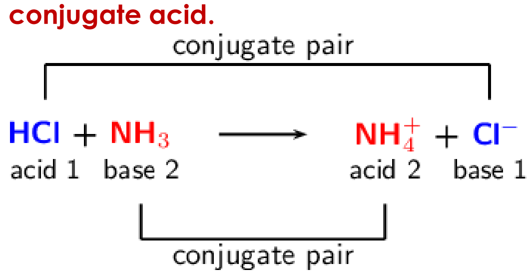
Solution (in 100g H_2O)	Sat, Unsat, Supersat	+/- how many °C to make saturated?	+/- how many g to make saturated?
10 g of $KClO_3$ at 30^oC	**SATURATED**	**N/A**	**N/A**
30 g NaCl at 40^oC	**UNSATURATED**	**VALUE IS OFF THE CHART**	**36 g – 30 g = ~6 g**
60 g KNO_3 at 30^oC	**SUPERSATURATED**	**37°C – 30°C = ~ 7°C**	**60 g – 45 g = ~15 g**
40 g $K_2Cr_2O_7$ at 80^oC	**UNSATURATED**	**80°C – 63°C = ~ 17°C**	**57 g – 40 g = ~17 g**

ACIDS AND BASES
What are the properties of acids and bases?
Acids: taste sour, pH < 7, turns blue litmus red.
Bases: taste bitter, slippery, pH > 7, turns red litmus blue.

35. What is an Arrhenius acid and base? How is it different from a Bronsted Lowry acid and base? **The two theories are just different ways to describe acids and bases. Bronsted-Lowry definition is generally more inclusive of all acid-bass. Arrhenius acid increases H+ concentration in solution while Arrhenius base increases OH- concentration. Bronsted Lowry acid is defined as a proton donor, while the Bronsted Lowry base is defined as a proton acceptor.**

36. Write a chemical equation that shows how HCl (hydrochloric acid), creates hydronium ions in an aqueous solution.
$HCl + H_2O \rightarrow Cl^- + H_3O^+$

37. What is a conjugate acid-base pair? Give an example. **Conjugate acid-base pair consists of two substances related to each other by the donating and accepting of a single hydrogen ion. When an acid donates a hydrogen ion, it forms the conjugate base. When a base accepts a hydrogen ion, it forms a conjugate acid.**



38. Identify the following as acids or bases:

a. HF **ACID**
b. NH_4OH **BASE**
c. HCN **ACID**
d. NH_4^+ **ACID**

40. Identify the conjugate acid for the following:

a. CN^- **HCN**
b. NH_3 **NH_4^+**
c. CO_3^{2-} **HCO_3^-**
d. Br^- **HBr**
e. H_2O **H_3O^+**
f. CH_3NH^- **CH_3NH_2**

41. Identify the conjugate base for the following:

a. H_2S **HS^-**
b. H_2SO_3 **HSO_3^-**
c. HCO_3^- **CO_3^{2-}**
d. HI **I^-**
e. HNO_3 **NO_3^-**
f. $H_2PO_4^-$ **HPO_4^{-2}**

42 .Fill in the following table:

Equation	Acid	Base	Conjugate Base	Conjugate Acid
$H_2O + CH_3NH_2 \rightarrow OH- + CH_3NH_3$	H_2O	CH_3NH_2	**OH-**	**CH_3NH_3**
$HF + NH_3 \rightarrow F^- + NH_4^+$	**HF**	**NH_3**	**F^-**	**NH_4^+**
$H_2O + H_2O \rightarrow OH^- + H_3O^+$	**H_2O**	**H_2O**	**OH-**	**H_3O+**
$HCl + H_2O \rightarrow Cl^- + H_3O^+$	**HCl**	**H_2O**	Cl^-	H_3O^+

43. Why do scientists tend to express the acidity of a solution in terms of pH rather than in terms of molarity of hydrogen ion present? **Since pH is the –log [H+], it makes it easier to express dilute concentrations of [H+] and compare the relative acidity of various solutions by understanding that each step in the pH scale is a difference of 10x. For example, instead of looking at 0.00000001 M [H+] and 0.000001 M [H+], it makes it easier to compare a pH = 8 and pH = 6. Also, since we have learned the relationship between pH and pOH, using the pH scale allows us to express the relative [H+] and [OH-] using one scale instead of measuring both concentrations.**

44. What is the mathematical definition of pH? **pH = -log [H+]**

45. As the hydrogen ion concentration of a solution increases, does the pH of the solution increase or decease? **pH decreases! pH < 7 = acidic. The lower the pH # gets, the more acidic a solution is.**

46. What is the relationship between pH and pOH? **pH + pOH = 14**

47. Indicate which of the following solutions is more acidic:

a. **$[H^+] = 5.69 \times 10^{-8}$ M** or $[OH^-] = 4.49 \times 10^{-6}$ M
pH = 7.24 pOH = 5.35, pH = 14 – 5.35 = 8.65

b. **$[H^+] = 2.6 \times 10^{-3}$ M** or $[H^+] = 4.5 \times 10^{-4}$ M

48. Complete the following table:

$[H^+]$	$[OH^-]$	pH	pOH	Acidic/Basic?
3.4×10^{-3} M	**2.95×10^{-12} M**	**2.47**	**11.53**	**Acidic**
2.09×10^{-6} M	4.8×10^{-9} M	**5.68**	**8.32**	**Acidic**
7.94×10^{-6} M	**1.26×10^{-9}**	5.1	**8.9**	**Acidic**
1.58×10^{-11} M	**6.31×10^{-4} M**	**10.8**	3.2	**Basic**

49. Calculate the pH and determine whether the solution is acidic or basic:

a. 8.6 M solution of HCl
$HCl + H_2O \rightarrow H_3O^+ + Cl$ (releases a hydronium ion, so acid)
pH = -log [8.6] = 0.93
Acidic

b. 0.000701 M solution of NaOH
$NaOH \rightarrow Na^+ + OH^-$ (releases a hydroxide ion, so base)
pOH = -log [0.000701] = 3.15
pH = 14-3.15 = 10.85
Basic

c. $[OH^-] = 0.000084$ M
pOH = -log [0.000084] = 4.08
pH = 14-4.08 = 9.92
Basic

d. $[H^+] = 6.9 \times 10^{-9}$ M
pH = -log [6.9×10^{-9}] = 8.16
Basic

50. What are the products of a neutralization reaction? **Water and a salt**