Review sheet for ap chemistry flashcard

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Write your answer in the space provided or on a separate sheet of paper.

1) Explain the difference between a qualitative and a quantitative measurement. Provide examples to illustrate this difference. Answer: A qualitative measurement is a measurement that gives descriptive, nonnumeric results; a quantitative measurement is a measurement that gives definite, usually numeric results. " The rock is heavy" would be a qualitative measurement. " The rock weighs 110 grams" would be a quantitative measurement. Topic: Scientific Measurement

2) Explain the difference between precision and accuracy. Suppose you made three different mass measurements of a sugar sample you knew to have a mass of 1 g. How would you know whether or not the measurements were accurate? How would you know whether or not they were precise? Could the three measurements be precise, but not accurate? Explain. Answer: Precision is the reproducibility, under the same conditions, of a measurement; accuracy is the closeness of a measurement to the true value of what is being measured. The three measurements would be precise if they were very close to each other in value; they would be accurate if they were close to the actual 1-g value for the mass of the sample. If the measurements are very close to each other, they are precise, regardless of how close they are to the real value. Therefore, the measurements could be precise, but not accurate. Topic: Scientific Measurement

3) Describe the rules that are used to determine the number of significant figures in the results of addition, subtraction, multiplication, and division. Answer: The answer of an addition or subtraction can have no more digits to

the right of the decimal point than are contained in the measurement with the least number of digits to the right of the decimal point. The answer of a multiplication or division can have no more significant figures than the measurement having the least number of significant figures. For these two operations, the position of the decimal point has nothing to do with the number of significant figures. Topic: Scientific Measurement

4) Why is the metric system the preferred system of measurement for science? Answer: The primary reason for this is that the metric system is based on units that are multiples of ten, thus simplifying conversions between units. In addition, all necessary units can be derived from the seven basic units of the metric system. Topic: Scientific Measurement

5) Why is the density of a metal greater than the density of water? Answer: The student should infer from the text that this is true either because the metal’s atoms are heavier than the water molecules or because the metal atoms are more closely packed than the water molecules, or both. Topic: Scientific Measurement

6) Why do the densities of most substances decrease with temperature? Answer: The student should infer that this is because a substance’s atoms or molecules tend to move farther apart with an increase in temperature. Consequently, the volume of the substance increases. There is no change in the mass of the substance, however, and therefore the density (mass/volume) decreases. Topic: Scientific Measurement

7) Explain the difference between specific gravity and density. Answer: Density has units and is absolute, depending only on temperature. Specific

gravity has no units and is a comparison of the density of a substance to the density of a reference substance, usually at the same temperature. The density of aluminum at 25°C is 2. 76 g/cm3. The specific gravity of aluminum at 25°C, using the reference standard of water at 4°C, is simply 2. 70. Topic: Scientific Measurement

8) Explain the difference between the Celsius and Kelvin temperature scales. Answer: Both scales use the freezing point and boiling point of water as reference temperature values. The Celsius scale designates the freezing point of water as O°C and the boiling point as 100°C. The region between these two points is divided into equal intervals known as degrees. The Kelvin scale designates 0 K as the temperature at which the volume of an ideal gas would be zero. It is called the absolute zero because it is the lowest temperature that is theoretically attainable. Absolute zero corresponds to on the Celsius scale. The Kelvin scale uses degree intervals that are the same size as the intervals on the Celsius scale. The difference between the scales lies in how the zero point is chosen. On the Celsius scale, the zero point is the freezing point of water. On the Kelvin scale, it is the point at which the volume of an ideal gas would theoretically be zero. The scales are related by the following formulas: K = 0°C + 273 or 0°C = K – 273. Topic: Scientific Measurement

9) Explain the terms molecular formula and formula unit. Give an example of each. Answer: A molecular formula shows the kinds and numbers of atoms present in a molecule of a compound. CO 2 is an example of a molecular formula. A formula unit is the lowest whole-number ratio of ions in an ionic compound. BaCl 2 is an example of a formula unit. The ratio of barium to

chloride ions is 1: 2. BaCl 2 is not a molecule. Topic: Chemical Name & Formula

10) Why was it necessary for chemists to develop a system for naming chemical compounds? Answer: Common names do not describe the chemical composition of a compound. They may give a physical or chemical property, but usually do not reveal what elements are in the compound. The systematic method that was adopted not only tells what atoms are in the compound, but gives the ratio in which they have combined to form the compound. Scientists in any country can tell what components are in a compound, and their relative amounts, when the systematic method for naming compounds is used. Topic: Chemical Name & Formula

11) Compare the characteristics of ionic and molecular compounds. Answer: Both are composed of at least two elements. Ionic compounds are made of oppositely charged ions. Molecular compounds are made of molecules. Ionic compounds are composed of metallic and nonmetallic elements. Molecular compounds are composed of nonmetallic elements only. Ionic compounds are solids with high melting points. Molecular compounds can be solids, liquids, or gases at room temperature, and have low melting points. Topic: Chemical Name & Formula

12) Explain how to write an ionic formula, given an anion and a cation. As an example, use phosphate anion. Write formulas for the compounds produced by the combinations of these ions. Name the compounds for which you have written formulas. Answer: The cation is written first, the anion second. The resulting formula must indicate an electrically neutral substance. The charge

of the anion becomes the subscript of the cation, and the charge of the cation becomes the subscript of the anion. When the charges of the two ions are the same, no subscripts are written. Example: Cu^+ , copper(I) phosphate Cu_3PO_4 Cu^{2+}, copper(II) phosphate $Cu_3(PO_4)_2$ Topic: Chemical Name & Formula Many other examples are possible.

13) Name the compounds $CuBr_2$, SCl_2, and BaF_2. Explain the use or omission of the Roman numeral (II) and the prefix di-. Answer: $CuBr_2$ is copper(II) bromide. The name must include a Roman numeral because copper is a transition element. SCl_2 is sulfur dichloride. The compound is named with prefixes because sulfur is a nonmetal. BaF_2 is barium fluoride. Neither a Roman numeral nor a prefix is needed in this name because barium is a Group A metal. Topic: Chemical Name & Formula

14) What is the advantage of using the specific term, gram molecular mass, instead of the general term, gram formula mass? Answer: There may be some confusion when the terms are applied to diatomic gases such as nitrogen and oxygen. For instance, the gram molecular mass of nitrogen is 28 g. The gram formula mass could be either 14 g or 28 g, depending upon whether the subject is nitrogen atoms or nitrogen molecules. Topic: Unit Conversion

15) Why is it possible to calculate the density of a gas at STP, knowing only its gram formula mass, but it is not possible to make the same calculation for a solid or a liquid? Answer: The molar volume of any gas is 22. 4 L at STP. The density can be calculated in the following way: x = density of gas (g/L)

However, the molar volumes of different solids and liquids are not uniformly the same at any prescribed condition. Topic: Unit Conversion

16) What determines whether one metal will replace another metal from a compound in a single-replacement reaction? Answer: Whether one metal will replace another is determined by the relative reactivity of the two metals. The activity series of metals lists metals in order of decreasing reactivity. A reactive metal will replace any metal found below it in the activity series. Topic: Chemical Reactions

17) What is the importance of the coefficients in a balanced chemical reaction? Answer: The coefficients in a balanced chemical equation indicate the relative number of moles of reactants and products. From this information the amounts of reactants and products can be calculated. The number of moles may be converted to mass, volume, or number of representative particles. Topic: Stoichiometry

18) In which kind of stoichiometric calculation can the steps involving conversion to and from moles be omitted? Explain why it is possible to do so. Answer: Volume-volume conversions between gases do not require mole conversions. Molar volumes of all gases at STP are the same. The coefficients in a balanced equation indicate the relative number of moles and the relative volumes of interacting gases. Topic: Unit Conversion

19) Name the three basic assumptions of the kinetic theory. Answer: The kinetic theory is based upon the assumptions that a gas is composed of particles, that these particles are in constant random motion, and that all collisions between particles are elastic. Topic: Matter- IMF

20) Explain why the pressure exerted by a gas does not depend on the size of the gas particles. Answer: Gas particles at the same temperature have the same average kinetic energy. When the same number of particles with the same average kinetic energy are contained in the same amount of space, they should exert the same pressure, regardless of their size. Also keep in mind that the particles of a gas are very far apart with nothing but space between them. No matter how large the particles are, they are still small when compared to the volume of space occupied by the gas. Topic: Matter-IMF

21) Explain why a liquid will eventually evaporate completely if held at a constant temperature below its boiling point. Answer: The particles in a liquid have a range of energies. At any particular temperature, some particles of the liquid have enough energy to overcome the attractive forces at the surface of the liquid and escape to the vapor state. As the high energy particles leave, the temperature of the liquid would normally go down. However, when exposed to the environment, the liquid will absorb heat from the environment to maintain a constant temperature. As a result, additional particles of the liquid will gain enough energy to escape the liquid. This process of escape, absorption of heat, and more escape is repeated until the liquid has evaporated completely. Topic: Matter- IMF

22) Describe the structure of the water molecule and indicate how this structure is responsible for many of the unique properties of this vital compound. Answer: Water is a simple, triatomic molecule. Each O H covalent bond in the water molecule is highly polar. Because of its greater electronegativity, the oxygen atom attracts the electron pair of the covalent

O H bond and acquires a slightly negative charge. The hydrogen atoms, being less electronegative than the oxygen, acquire a slightly positive charge. The atoms of the water molecule are joined in a 105° angle. As a result, the slight charges on the individual atoms do not cancel each other out and the molecule itself is polar. There is a slight negative charge in the region around the oxygen and a slight positive charge in the region around the hydrogens. Because water molecules are polar, they attract each other.

The hydrogen of one molecule is attracted to the oxygen of another molecule. This attraction is termed hydrogen-bonding and it is stronger than other polar attractions because of the fact that the hydrogen nucleus is not shielded by an electron cloud in the way that other nuclei would be (hydrogen atoms have only 1 electron). It is the strong intermolecular attraction associated with hydrogen-bonding that is responsible for many of the unusual properties of water, including its high surface tension, low vapor pressure, high specific heat, high heat of vaporization, and high boiling point. Topic: Aqueous Stoichiometry

23) Why will a needle float on the surface of water but sink immediately if it breaks through the surface? Answer: The surface of the water presents greater resistance (surface tension) to the needle than does the rest of the water. This is because water molecules are packed more closely at the surface than elsewhere. This closer packing results from the one-sided intermolecular attraction that exists at the surface of the water. The water molecules at the surface are attracted to the water molecules below them. But, unlike in the rest of the liquid, there are no water molecules above them to pull them in the opposite direction. The closer-packed water molecules at

the surface form a “ skin” that is denser than the water below. Objects denser than water (e. g. the needle) can float upon this skin. All liquids will exhibit surface tension, but this phenomenon is more pronounced in water because of its hydrogen-bonding. Topic: Aqueous Stoichiometry

24) How is water’s high heat capacity related to the ability of a large body of water to moderate temperature changes in its vicinity? Answer: As a result of its high heat capacity, water can absorb and release significant amounts of energy with only relatively small changes in temperature. Consequently, large bodies of water act to moderate temperature changes in their vicinities. As the temperature of the surroundings goes up, water absorbs heat energy, but its temperature goes up only relatively slightly. This has the effect of moderating the temperature in the area around the water. In cold weather, a similar phenomenon takes place. As the temperature of the surroundings drops, water will give off energy to the surroundings, but the temperature of the water will go down only slightly. In this way, cold temperatures also are moderated in areas near bodies of water. Topic: Aqueous Stoichiometry

25) Explain why water has a relatively high heat of vaporization. Answer: Water has a high heat of vaporization as a result of its hydrogen-bonding. Because of its extensive network of hydrogen-bonds, the molecules of water are held together more tightly than are the molecules of many other liquids. The attractive force of these hydrogen-bonds must be overcome in order for water to vaporize. Topic: Aqueous Stoichiometry

26) Why is ice less dense than water? Answer: The structure of ice is a very regular, open framework in which the water molecules are farther apart from each other than they are in liquid water. When ice melts, this open framework collapses and the water molecules move closer together. As a result, the water is denser than the ice. Topic: Aqueous Stoichiometry

27) Define the terms solute, solvent, and aqueous solution. Provide an example of each. Answer: An aqueous solution is any sample of water that contains one or more dissolved substances. A solvent is the dissolving medium in a solution. A solute is the dissolved material in a solution. Salt (NaCl) in water is an example of an aqueous solution. In this solution, water is the solvent and salt is the solute. Topic: Aqueous Stoichiometry

28) Describe the process of solvation. Answer: Solvent molecules collide with solute particles and exert attractive forces on them. Whenever these forces are greater than the attractive forces within the solute, the solute particles separate from the bulk of the solute. The particles of solute then become surrounded by particles of solvent. Topic: Aqueous Stoichiometry

29) What is an electrolyte? Give examples and distinguish between a strong electrolyte and a weak electrolyte. Answer: An electrolyte is any substance that will conduct an electric current in aqueous solution or in the molten state. Strong electrolytes are substances that are completely, or almost completely, ionized in water. Examples of strong electrolytes are sodium chloride, hydrochloric acid, and sodium hydroxide. Weak electrolytes are substances that are only slightly ionized in water. Examples of weak

electrolytes are mercuric chloride and acetic acid. Topic: Aqueous Stoichiometry

30) Distinguish among efflorescent, hygroscopic, and deliquescent substances. Provide examples. Answer: Efflorescent substances are hydrates that spontaneously give off the water of hydration. A hydrate will effloresce if it has a water vapor pressure higher than that of the water vapor in the air. An example of an efflorescent compound is copper sulfate pentahydrate, $CuSO4 \bullet 5H20$. Hygroscopic substances are substances that absorb moisture from the air to form hydrates or solutions.

Examples of hygroscopic substances are phosphorus pentoxide, $P\ 2O5$; sodium hydroxide, NaOH; and calcium chloride monohydrate, $CaCl\ 2 \bullet H20$. Deliquescent substances are substances that can absorb so much moisture from the air as to actually dissolve in the absorbed water. For deliquescent substances, the water vapor pressure of the solution formed when the substance absorbs water from the air is lower than that of the water vapor in the air. An example of a deliquescent substance is sodium hydroxide. Topic: Aqueous Stoichiometry

31) Distinguish among a suspension, a colloid, and a solution. Give an example of each. Answer: Suspensions are mixtures out of which some particles will settle upon standing. Particles in a typical suspension have an average diameter greater than Suspensions are heterogeneous mixtures. An example of a suspension is a sand-water mixture. Colloids are mixtures containing particles that are intermediate in size between the particles in suspensions and the particles in true solutions.

The average diameter of a particle in colloidal suspension is between 1 nm and 100 nm. Colloids exhibit the Tyndall effect, which is the scattering of light in all directions by the particles in the suspension. (Note: Suspensions also exhibit the Tyndall effect, but solutions do not.) A solution is a homogeneous mixture. The particles in a solution have an average diameter of 1 nm. Solutions do not exhibit the Tyndall effect. Topic: Aqueous Stoichiometry

32) Explain what a saturated solution is. Give a specific example. Answer: A saturated solution contains the maximum amount of solute for a given amount of solvent at a constant temperature. For example, no more than 36. 2 g of sodium chloride will go into solution in 100 g of water. Above this concentration, there is a dynamic equilibrium between the solid and its dissolved ions. In this equilibrium, just as many ions are going out of the solution as are going in per unit time, and solid will, therefore, always be present. Topic: Solution

33) Discuss the different factors that can affect the solubility of a substance. Include specific examples in your discussion. Answer: The factors are temperature, pressure, and the nature of the substances. Specific examples include the following: Sodium chloride is more soluble in water at high temperature than at low temperature. Gases are less soluble at high temperatures than at low temperatures. The solubility of a particular gas increases as the partial pressure of that gas increases above the solution. Sodium nitrate is much more soluble in water than is barium sulfate, regardless of temperature. Polar substances tend to be soluble in water, whereas nonpolar substances tend to not be soluble in water. Topic: Solution

34) Discuss the phenomenon of supersaturation. Indicate how crystallization can be initiated in a supersaturated solution. Answer: A solution that contains more solute than it can theoretically hold at a given temperature is a supersaturated solution. A dynamic equilibrium between solid and dissolved particles does not exist because there is no solid. Crystallization can be initiated by adding a seed crystal to the solution or by exposing the solution to a rough surface. The latter can be done by scratching the inside of the container holding the solution. Topic: Solution

35) Explain on a particle basis how the addition of a solute affects the boiling point, the freezing point, and the vapor pressure of the solvent. Answer: Boiling point elevation, and freezing point and vapor pressure lowering are colligative properties. They depend on the number of particles in solution and not on the chemical nature of the solute or solvent. Boiling point elevation: Additional attractive forces exist between solute and solvent; they must be overcome for the solution to boil. Kinetic energy must be added to overcome these forces. Freezing point lowering: More kinetic energy must be withdrawn from a solution because the solute is surrounded by shells of solvent. This interferes with the formation of the orderly pattern that the particles assume as the solvent changes from liquid to solid. Vapor pressure lowering: The formation of solvent shells around the solute particles reduces the number of solvent particles that have sufficient kinetic energy to vaporize. Topic: Solution

36) Explain the effects of reactant concentration and particle size on the rate of a reaction. Answer: A small particle size increases the rate of a reaction because there is more surface area for a given mass of particles and so more

collisions are possible per second. A high concentration of reactants increases the reaction rate because more molecules are present to collide each second. Topic: Kinetics

37) What is the effect of a catalyst on the rate of a reaction? Use an example. Answer: A catalyst increases the reaction rate, by permitting the formation of a less energetic activated complex. Platinum is a catalyst for certain reactions of gases. Topic: Kinetics

38) Why are some reactions reversible? Answer: Reversible reactions have a small change in free energy, and so can easily proceed in either direction. Most reactions are reversible to some extent. Topic: Kinetics

39) What is free energy, and how is it related to spontaneity in a reaction? Answer: Free energy is a measure of the ability of a reaction to do work. Spontaneous reactions release free energy. Nonspontaneous reactions absorb free energy. Topic: Thermodynamics

40) What causes a reaction to be spontaneous? Give an example. Answer: A reaction, such as sodium with water, is spontaneous because it can result in a lower energy state or a more disordered state for the system. Sometimes the energy and the disorder both increase, and sometimes the energy and the disorder both decrease. An example of the former is the dissolution of ammonium nitrate. This reaction is spontaneous even though it is endothermic. It occurs because the favored increase in disorder that accompanies dissolution outweighs the unfavored increase in energy. Topic: Thermodynamics

41) Characterize spontaneous and nonspontaneous reactions. Answer: Spontaneous reactions are reactions that, under the conditions specified, are known to produce the written products. Nonspontaneous reactions do not give products under the specified conditions. Some spontaneous reactions go so slowly that they appear to be nonspontaneous. Topic: Thermodynamics

42) What is entropy? Give several examples. Answer: Entropy is the degree of disorder in a system. A gas has more entropy than a liquid. A chemical reaction in which there are more molecules of product than of reactant will cause an increase in entropy. A solution of sodium chloride in water has more entropy than a sodium chloride crystal. Topic: Thermodynamics

43) Compare and contrast the properties of acids and bases. Answer: Both acids and bases are electrolytes, they cause indicators to change colors, and they react with each other to form water and a salt. Acids taste sour, while bases taste bitter. Bases feel slippery. Acids react with some metals to produce hydrogen gas. Topic: Acid – Base Equilibria

44) Explain how to measure pH using a pH meter. Answer: Wash the electrodes in distilled water. Dip the electrodes in a buffer at pH 7, and calibrate the meter. Wash the electrodes again. Then measure the pH of the solution. Store the electrodes in a distilled water solution. Topic: Acid – Base Equilibria

45) What are acids and bases according to Arrhenius? Give examples. Answer: According to Arrhenius, acids give up a proton to water and bases give up a hydroxide in water. Hydrochloric acid is an Arrhenius acid because

it gives up its proton as it dissolves in water. Sodium hydroxide is an Arrhenius base because it gives up its hydroxide as it dissolves in water.

Topic: Acid – Base Equilibria

46) What are acids and bases according to the Bronsted-Lowry theory?

Answer: According to the Bronsted-Lowry theory, acids donate protons to other substances and bases accept protons from other substances. Ammonia accepts a proton from water and therefore acts as a Bronsted-Lowry base.