Chemistry CP

Name:

End of Year Review Sheet

Period:

Measurement & Properties of Matter

- Distinguish between the physical properties and chemical properties of matter.
- Classify changes of matter as chemical or physical.
- Explain the gas, liquid, and solid states in terms of arrangement of particles.
- Classify a sample of matter as a substance or a mixture; as homogeneous or heterogeneous.
- Explain the difference between an element and a compound.
- Distinguish the symbols of common elements, and match the names of common elements to their symbols.
- List the characteristics that distinguish between metals, nonmetals, and metalloids.
- List the SI units of measurement used in chemistry.
- Distinguish between the accuracy and precision of a measurement.
- Identify the number of significant figures in a measurement.
- Apply the rules for significant figures in calculations to round off numbers correctly.
- Calculate the density of an object from experimental data.
- Calculate the percent error of an experimentally determined measurement.

Questions

1. The table below shows the physical properties of selected metals.

Physical Properties of Selected Metals

Metal	Molecular mass (amu)	Melting point (°C)	Boiling point (°C)	Density (g/cm^3)
Bismuth	209.98	271	1560	9.80
Chromium	52.00	1857	2672	7.20
Polonium	210.05	254	962	9.40
Ruthenium	101.07	2310	3900	12.3

A cube of an unknown metal has a volume of 2.25 cm^3 and a mass of 16.2 g. Based on data in the table, what is the identity of this metal?

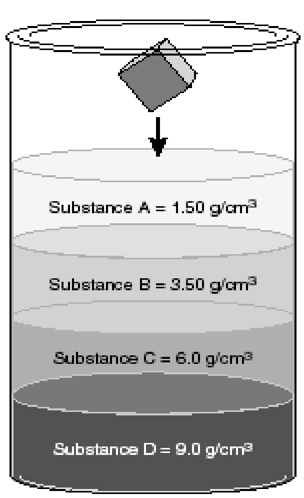
A. Bi B. Cr

C. Po D. Ru

2. A pot containing a few milliliters of water is placed on a hot burner. The water is boiled until no water is left in the pot. A frying pan is placed on a hot burner. A raw egg is taken out of its shell and placed in the frying pan until the egg white becomes solid.

a. Describe the change that takes place in the water. Be sure to indicate whether the change is chemical or physical.

b. Describe the change that takes place in the egg white. Be sure to indicate whether the change is chemical or physical.



3. A solid cube was put into a cylinder containing four liquids with different densities as shown below.

The cube fell quickly through layer A, fell slowly through layer B, and stopped upon reaching layer C. The density of the cube **most likely** lies between

A. 1.00 & 1.50 g/cm^3. B. 1.51 & 3.50 g/cm^3

C. 3.51 & 6.00 g/cm^3 D. 6.00 & 9.00 g/cm^3

4. How many significant figures are in the following numbers: 0.0045 1.23×10^5 10001

5. What is the volume of the liquid in the cylinder to the right?

6. What is the sum of 6.6412g + 12.85g + 0.046 g + 3.48 g expressed to the correct number of significant figures? A. 23 g B. 23.0 g C. 23.017 g D. 23.02 g

7. Write 0.000456g in scientific notation.

8. A student calculated the percent by mass of water in a hydrate to be 37.2%. If the accepted value is 36.0%, what is the percent error in the student's calculation?

9. Which of the following elements is a nonmetal? A. fluorine B. copper C. magnesium D. sodium

10. Which element is a noble gas? A. antimony B. krypton C. gold D. francium

11. Which element is a metalloid? A. magnesium B. silicon C. chromium D. argon

12. Which element in group 15 would most likely have luster and good electrical conductivity?
A. N B. P C. Bi D. As

Atomic Structure

- Explain the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.
- Summarize the five essential points of Dalton's atomic theory.
- Distinguish among protons, neutrons, and electrons in terms of their relative masses and charges.
- Explain the structure of an atom, including the location of the proton, neutron, and electron with respect to the nucleus.
- Explain how atomic number identifies an element.
- Summarize the observed properties of cathode rays that led to the discovery of the electron.
- Summarize Rutherford's experiment that led to the discovery of the nucleus.
- Explain how isotopes of an element differ.
- Explain, using concepts of isotopes, why the atomic masses of elements are not whole numbers.
- Infer the number of protons, electrons, and neutrons using the atomic number and mass number of an element.
- Discuss the dual wave-particle nature of light.
- Explain the relationship between frequency, wavelength, and energy of light.
- Compare and contrast the Bohr model and the quantum model of the atom.
- Relate the number of sublevels corresponding to each of an atom's main energy levels, the number of orbitals per sublevel, and the number of orbitals per main energy level.
- Apply the Aufbau principle, the Pauli Exclusion Principle, and Hund's rule to write the electron configurations of the elements.
- Explain the origin of the atomic emission spectrum of an element.

Questions:

13. Where in the atom are the protons, electrons, and neutrons? Which one(s) are represented by the atomic number? ________Which one(s) are represented by the atomic mass? ________

14. Which of the following represents a pair of isotopes?

A. ^{1}H and ^{3}H B. $^{16}O^{2-}$ and $^{19}F^{1-}$ C. ^{40}K and ^{40}Ca D. $^{16}O^{2-}$ and $^{32}S^{2-}$

15. Use arrows and boxes to write the electron configuration of Na:

16. Use noble gas notation to write the electron configuration of Fe^{3+}

Nuclear Chemistry & Periodicity

- Explain the processes of radioactivity and radioactive decay.
- Distinguish between isotopes and radioisotopes.
- Describe the characteristics of alpha, beta, and gamma radiation and list their origins.
- Define the terms nuclear stability, half-life, and transmutation.
- Write balanced nuclear equations.
- Compare fission and fusion processes.
- Calculate the amount of radioisotope remaining using the half-life method.
- Explain how radioisotopes can be used to date objects.
- Explain the roles of Mendeleev and Moseley in the development of the periodic table.
- Distinguish between a group and a period in the periodic table.
- Categorize the elements as main group element, noble gas, transition metal, or inner transition metal.
- Describe how the elements belonging to a group of the periodic table are interrelated in terms of atomic number.
- Locate and name the four blocks of the periodic table. Explain the reasons for these names.
- Discuss the relationship between group configurations and group numbers.
- Describe the locations in the periodic table of the alkali metals, the alkaline-earth metals, the halogens, and the noble gases.
- Interpret the trend shown by atomic radii within the periodic table.
- Explain the variation in ionization energies within the periodic table.
- Interpret the trend shown by atomic sizes within the periodic table.
- Interpret the trend shown by electronegativities within the periodic table.
- State how many valence electrons are present in atoms of each main-group element.
- Use factor label strategies in problem solving

Questions:

17. Is the following a fusion or fission process? Explain. $^{3}_{1}\mathrm{H} + ^{1}_{1}\mathrm{H} \rightarrow ^{4}_{2}\mathrm{He}$

18. Which statement best describes the alkaline earth elements?
    A. They have one valence electron, and they form Ions with a 1+ charge.
    B. They have one valence electron, and they form Ions with a 1- charge.
    C. They have two valence electrons, and they form Ions with a 2+ charge.
    D. They have two valence electrons, and they form Ions with a 2+ charge.

19. Write a balanced nuclear equation for each of the following nuclear reactions.
    a. Uranium-233 undergoes alpha decay.
    b. Neptunium-239 undergoes beta decay.

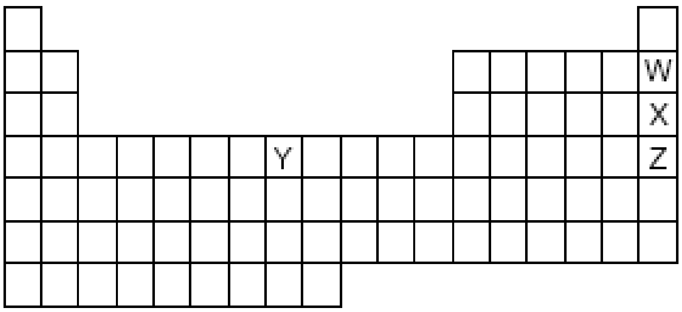
20. An isotope has a half-life of 3 days. How much of a 16 gram sample will be left after 9 days?

21. Which of the following elements has the highest electronegativity?
A. B (boron) B. C (carbon) C. O (oxygen) D. N (nitrogen)

22. Which element has the highest first ionization energy?
A. sodium B. aluminum C. calcium D. phosphorus

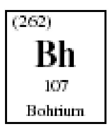
23. Which element has atoms with the largest covalent radius?
A. rubidium B. cesium d. strontium D. barium

24. The figure below represents the periodic table and the location of four different elements



A certain element has a ground state electron configuration of $1s^22s^22p^63s^23p^6$. Which letter in the diagram above represents the position of this element on the periodic table?
A. Y B. W
C. X D. Z

25. Which of the following elements can form an anion that contains 54 electrons, 74 neutrons, and 53 protons?

A.


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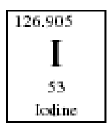
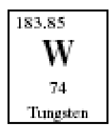
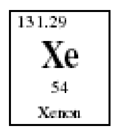
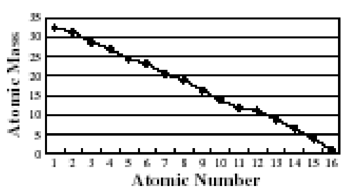


C.


D.


26. Which of the following graphs **best** shows the relationship between an element's atomic mass and its atomic number?

A.


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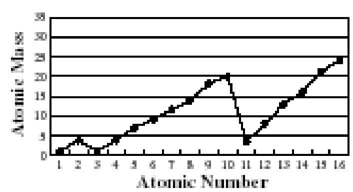

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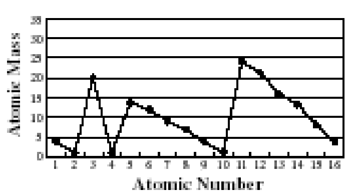
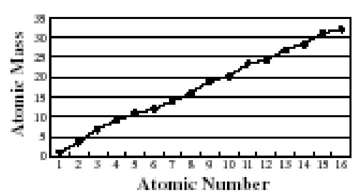


D.


27. The most reactive metal is __________ and the most reactive non-metal is _____________

28. Circle the atom/ion in each pair that has the larger radius:
Mg vs Na Cl vs Br Cl vs Cl^- Na vs Na^+

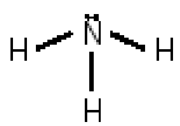
Bonding

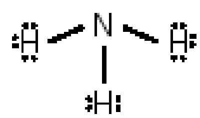
- Infer the number of valence electrons in an atom of a main-group element, and then construct its Lewis dot structure.
- Explain why most atoms form chemical bonds.
- Describe the formation of a cation from a metallic element and the formation of an anion from a non-metallic element.
- State the octet rule.
- List the characteristics of a covalent bond.
- Create Lewis structures for covalent compounds containing single, double, and triple bonds.
- Explain the modern interpretation of resonance bonding.
- List the characteristics of an ionic bond.
- Compare and contrast the properties of ionic and molecular compounds.
- Classify bonds as ionic, covalent, or polar based on electronegativity differences.
- Use electron dot formulas to determine chemical formulas for binary ionic compounds.
- Describe the shapes of simple covalently bonded molecules using VSEPR theory.
- State the diatomic elements: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2

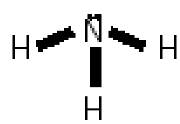
Questions:

29.How many valence electrons are present in an atom of Mg? What is its electron dot symbol?

30. The chemical formula for ammonia is NH_3. Which of the following is the correct Lewis electron dot structure for ammonia?

A.


B.


C.


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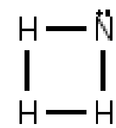

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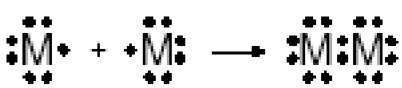
31. In potassium fluoride, the potassium atom donates an electron and the fluorine atom takes an electron. When the compound potassium fluoride is formed, which of the following are formed?
A. covalent bonds B. ionic bonds C. magnetic forces D. nuclear forces

32. An atom of which element has an ionic radius smaller than its atomic radius?
A. N B. S C. Br D. Rb

33. When elements from group 1 (1A) combine with elements from group 17 (7A), they produce compounds. Which of the following is the correct combining ratio between group 1 (1A) elements and group 17 (7A) elements? A. 1:1 B. 1:2 C. 2:1 D. 3:2

34. The figure below shows two atoms of a fictitious element (M) forming a diatomic molecule.



What type of bonding occurs between these two atoms?
A. covalent B. ionic C. nuclear D. polar

35. What is the geometric shape of a methane molecule?
A. triangular B. rectangular C. octahedral D. tetrahedral

36. Which of the following is bonding most likely to be ionic, covalent, or both:
NaOH CH_4 BaF_2

37. For the following compounds, predict the a. Lewis structure b. shape c. overall polarity
HCl CO_2 H_2O BF_3

Chemical Names and Formulas

- Infer the charge on a monatomic ion using the periodic table.
- Determine the formula of an ionic compound formed between two given ions.
- Name an ionic compound given its formula.
- Define a polyatomic ion and memorize the names and formulas of common polyatomic ions.
- Using prefixes, name a binary molecular compound from its formula.
- Write the formula of a binary molecular compound given its name.
- Classify compounds as either ionic or molecular.
- Calculate the gram formula mass of any given compound.
- Use gram formula mass to convert between mass in grams and amount in moles of a chemical compound.
- Define how Avogadro's number is related to a mole of any substance.
- Calculate the number of molecules, formula units, or ions in a given molar amount of a chemical compound.
- Calculate the percentage composition of a given chemical compound or experimental data.
- Derive the empirical formula of a compound from experimental data (either a percentage or a mass composition).
- Derive the molecular (true) formula of a compound from experimental data.

Questions:

38. Limestone is a naturally occurring form of calcium carbonate. The correct formula for limestone is: A. $Ca(CO_3)_2$ B. $CaCO_3$ C. Ca_2CO_3 D. $Ca_2(CO_3)_2$

39. Which formula is correctly paired with its name?
    A. $MgCl_2$; magnesium chlorine
    B. K_2O; phosphorus dioxide
    C. $CuCl_2$; copper(II) chloride
    D. FeO; iron(III) oxide

40. Which formula is correctly paired with its name?
    A. N_2O; nitrogen oxide B. P_4O_{10}; phosphorus oxide
    C. CBr_4; carbon bromide D. SO_3 sulfur trioxide

41. Which formula represents lead(II) phosphate?
    A. $PbPO_4$ B. $Pb4PO_4$ C. $Pb_3(PO_4)_2$ D. $Pb_2(PO_4)_3$

42. How many water molecules are contained in 30.0 g of water?

43. What is the empirical formula for C_2H_4? ___________________

44. A compound contains 50% sulfur by mass and 50% oxygen by mass. What is its empirical formula?

45. How many moles of oxygen atoms are present in 2 moles of $Mg_3(PO_4)_2$?
    A. 4 B. 8 C. 12 D. 16

Chemical Reactions

- List indirect evidence that a reaction has occurred.
- Identify the reactants and products in a chemical reaction.
- Rewrite a chemical equation from a description of a chemical reaction using appropriate symbols and formulas.
- Demonstrate the ability to write and balance chemical reactions when given the names or formulas of all reactants and products.
- Classify a reaction as synthesis, decomposition, single replacement, double displacement, or combustion.
- Predict the products of simple reactions given the reactants.
- Use the activity series of metals to predict whether a given reaction will occur and to predict the products of single replacement reactions.
- Use solubility tables to predict precipitant formation.

Questions

46. Copper in the compound $CuSO_4$ can be isolated in the following reaction with iron.

$$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$$

What type of reaction is shown above?

A. decomposition B. synthesis C. single displacement D. double displacement

47. A balanced chemical reaction is shown here: $C_5H_{12} + 8O_2 \rightarrow 5CO_2 + 6H_2O$

If 3 moles of C_5H_{12} are reacted completely, how many moles of water are formed?

A. 3 B. 6 C. 12 D. 18

48. Given the unbalanced chemical equation: ___$Pb(NO_3)_2$ + ___$K_2CrO_4 \rightarrow$___$PbCrO_4$ + ___KNO_3

Which coefficients are needed to balance this equation?

A. 2, 4, 4, 6 B. 2, 4, 4, 3 C. 1, 1, 1, 2 D. 1, 6, 6, 2

49. Calculate the molar masses of C_3H_8 and $CuSO_4 \bullet 5H_2O$

50. Potassium carbonate is an important component of fertilizer. The partially balanced equation for the reaction of 6 moles of potassium hydroxide (KOH) and 3 moles of carbon dioxide (CO_2) to produce potassium carbonate and water is given: $6\ KOH + 3\ CO_2 \rightarrow$ ___ $K_2CO_3 + 3\ H_2O$ When this equation is balanced, what is the coefficient for potassium carbonate?

A. 2 B. 3 C. 6 D. 9

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51. a) 0.20 mole O_2 = ________ grams b) 4 moles Na = _______________ atoms of Na

52. Calculate the % of carbon in K_2CO_3

53. Aluminum reacts vigorously and exothermically with copper(II) chloride. Which of the following is the balanced equation for this reaction?

A. $Al + CuCl_2 \rightarrow AlCl_3 + Cu$ B. $Al + 3CuCl_2 \rightarrow 2AlCl_3 + Cu$
C. $2Al + 3CuCl_2 \rightarrow 2AlCl_3 + 3Cu$ D. $3Al + 2CuCl_2 \rightarrow 3AlCl_3 + 2Cu$

54. Which of the following represents a double displacement reaction?

A. ABC $\rightarrow$ AB + C B. A + B $\rightarrow$ AB
C. AB + CD $\rightarrow$ AD + CB D. A + BC $\rightarrow$ AC + B

55. What are the two products of a combustion reaction of a hydrocarbon?

Stoichiometry

- Interpret balanced chemical equations in terms of interacting moles, representative particles, masses and volumes (at STP).
- Construct mole ratios from balanced chemical equations and apply these ratios in calculating mole-mole stoichiometric quantities.
- Calculate stoichiometric quantities from balanced chemical equations using units of mass.
- Calculate stoichiometric quantities from balanced chemical equations using units of moles, mass, representative particles, and volume (gases at STP).
- Identify the limiting reagent in a reaction and use it to calculate stoichiometric quantities and the amount of excess reagent(s).
- Calculate the theoretical yield, actual yield, and/or percent yield for a chemical reaction.

Questions:

56. A student burned a sample of pure carbon in an open crucible. The carbon reacted with oxygen in the air and produced carbon dioxide.

a. Write the balanced equation for the complete combustion of carbon.

b. The student observed no visible products. Why does it appear that the law of conservation of mass was violated by this reaction?

c. If three moles of carbon are burned, how many moles of oxygen gas will be consumed, and how many moles of product should be obtained? Explain how you arrived at your answers.

57. Write the balanced chemical equation for the formation of water from hydrogen and oxygen.

If you have 5 grams of hydrogen, how many grams of water should be made?

If you have 5 grams of hydrogen and 5 grams of oxygen, what is the limiting reactant?

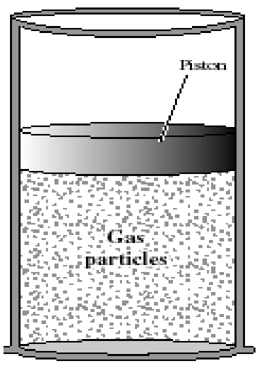
Gases

- Describe the motion of particles of a gas according to the kinetic theory.
- Relate that the temperature of a substance is a measure of the kinetic energy of the particles in that substance.
- Distinguish between real and ideal gases.
- Interpret gas pressure in terms of kinetic theory.
- Determine the pressure of a confined gas using open and closed manometers.
- Convert between different units of pressure.
- Calculate pressure or volume from the pressure-volume relationship of a contained gas at constant temperature.
- Calculate temperature or volume from the temperature-volume relationship of a contained gas at constant pressure.

- Calculate temperature or pressure from the temperature-pressure relationship of a contained gas at constant volume.
- Calculate pressure, volume, or temperature from the temperature-pressure-volume relationships of confined gases.
- Calculate the total pressure of a mixture of gases or the partial pressure of a gas in a mixture of gases.
- Calculate the amount of gas at any specified conditions of pressure, volume, and temperature.
- Explain, using kinetic theory, why molecules of small mass diffuse more rapidly than molecules of large mass.

Questions:

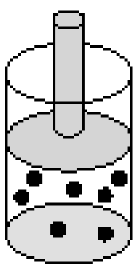
58. The picture below shows a gas at standard conditions in a container with a moveable piston.



According to Charles's law, what will happen to the piston when the gas is heated?

A. The piston will move up because the gas particles get larger.
B. There will be no change because heat will not affect the system.
C. The piston will move up because the gas particles move faster and get farther apart.
D. The piston will move down because the gas particles move slower and get closer together.

59. A cylinder of gas particles is shown below.



The cylinder is fitted with a moveable piston that can be raised and lowered. Which of the following would result in an **increase** in the pressure of the gas below the piston?

A. increasing the volume of the cylinder
B. removing some of the gas from the cylinder
C. decreasing the volume of the cylinder
D. decreasing the pressure outside the cylinder

60.The pressure on 30. mL of an ideal gas increases from 760 torr to 1520 torr at constant temperature. The new volume is:

A. 30. mL x 760 torr/1520 torr C. 30. mL x 1520 torr/760 torr
B. 1520 torr x 760 torr/30. mL D. 760 torr x 1520 torr/30. mL

61. 56 K = ______°C

62. A 1.00 L sample of a gas has a mass of 1.92 g at STP. What is the gram formula mass of the gas? (R = 0.0821 L atm/mole K)

A. 1.92 g/mol B. 19.2 g/mol C. 22.4 g/mol D. 43.0 g/mol

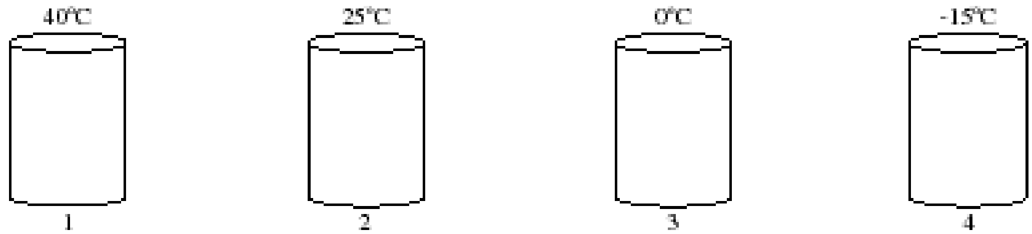
63. At constant temperature and pressure, gas volume is directly proportional to the

A. gram formula mass of the gas C. density of the gas at STP
B. number of moles of gas. D. rate of diffusion

64. A 75.0 mL sample of gas exerts 200. mm Hg pressure at 30.^{o}C. What pressure does it exert at 35.0^{o}C if the volume expands to 80.0 mL?

A. 90.0 mm Hg B. 161 mm Hg C. 190 mm Hg D. 219 mm Hg

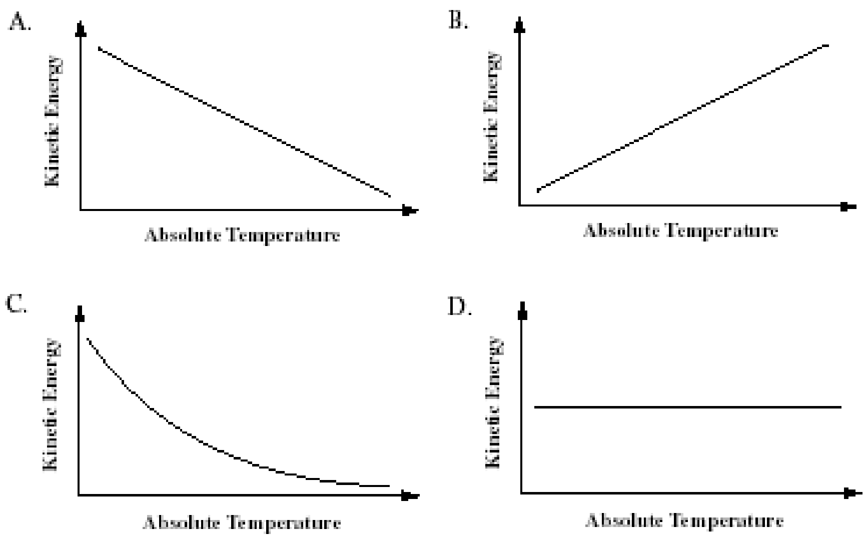
65. The illustration below shows four containers. Each container is full of helium gas at a different temperature.



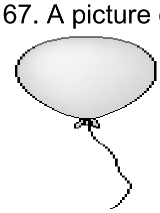
If all of the containers are closed and have a pressure of 1 atm, which container has helium particles with the **greatest** average kinetic energy?

A. 1 B. 2 C. 3 D. 4

66. Which of the following graphs **best** represents how the average kinetic energy of molecules changes with absolute temperature?



67. A picture of a balloon is shown below.



If the temperature of this balloon were to decrease suddenly, how would the balloon change?

A. Its mass would increase. B. Its mass would decrease.
C. Its volume would increase. D. Its volume would decrease.

68. What is the partial pressure of a mixture of CO_2, SO_2 and H_2O gases if each gas has a partial pressure of 250 torr? A. 250 torr B. 500 torr C. 750 torr D. 1000 torr

69. Which gas diffuses most rapidly at STP? A. Ar B. Kr C. N_2 D. O_2

70. At .500 atm and 15.0^{o}C, a sample of gas occupies 120. L. What volume does it occupy at 0.250 atm and 10.0^{o}C? A. 60 L B. 111 L C. 236 L D. 480 L

Solutions

- Distinguish between homogeneous and heterogeneous mixtures
- Compare the properties of suspensions, colloids, and solutions.
- Distinguish between electrolytes and nonelectrolytes.
- Explain the difference between saturated, unsaturated, and supersaturated solutions.
- Explain the solution process in terms of hydration and polarity.
- Use solubility graphs to discuss the effect of temperature on solubility.
- Define and work problems involving the molarity of a solution.
- Calculate percent by volume and percent by mass for solutions.

Questions

71. Polar solutes will only dissolve in what type of solvent?

72. At 25°C, 40 grams NaCl is added to 300 grams of water. Is it unsaturated or saturated? (You need to look up the solubility curve in your handouts or text)

73. When adding solute to solvent, the freezing point of the solution will (inc, dec, or remain same). What will happen to the boiling point?

74. 0.50 moles of solute are dissolved in 200 mL of solution. What is the molarity?

75. How many grams of NaCl are in 100. mL of a 0.1 M solution (NaCl is 58.5 g/mole)

76. The solubility of a substance can be described in a variety of ways. Some references may use descriptive terms for solubility, such as those in the table illustrated below.

Descriptive terms	Parts of solvent needed for 1 part solute
Very soluble	<1
Freely soluble	1–10
Soluble	10–30
Sparingly soluble	30–100
Slightly soluble	100–1,000
Very slightly soluble	1,000–10,000
Practically insoluble or insoluble	>10,000

Using the table as a reference, what descriptive term would be used for a medication that required 4,000 mg of water to dissolve 200 mg of the drug?

A. soluble
B. slightly soluble
C. sparingly soluble
D. very slightly soluble

Thermochemistry

- Distinguish between temperature and heat.
- Explain the heat capacity of objects and express it in standard units of heat.
- Describe heat changes in terms of a system and its surroundings.
- Use specific heat capacity to calculate the heat changes that occur in chemical and physical properties.
- Construct equations that show the heat changes for chemical and physical processes.

- Describe in words and in diagrams the heat changes that occur in melting, freezing, boiling, and condensing.
- Apply Hess' Law of heat summation to find heat changes for chemical and physical processes.
- Show how changes in entropy relate to a change of state, a change in temperature, and a change in the number of product particles compared with reactant particles.
- Explain how changes in energy and changes in entropy both influence the spontaneity of a reaction.

Questions:

77. Draw and label a potential energy diagram for an endothermic reaction with and without a catalyst. Label the reactants, products, and E_a of the forward reaction.

78. Which phase change results in a release of energy?
    A. $Br_2(l) \rightarrow Br_2(s)$
    B. $I_2(s) \rightarrow I_2(g)$
    C. $H_2O(s) \rightarrow H_2O(l)$
    D. $NH_3(l) \rightarrow NH_3(g)$

79. Which system tends to become LESS random as reactants form products?
    A. $C(s) + O_2(g) \rightarrow CO_2(g)$
    B. $S(s) + O_2(g) \rightarrow SO_2(g)$
    C. $I_2(s) + Cl_2(g) \rightarrow 2\ ICl(g)$
    D. $2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$

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80. Using the following equations, find the ΔH of combustion of methane.

Target equation:

$C(s) + O_2(g) \rightarrow CO_2(g)$	$\Delta H = -393.5$ kJ
$O_2(g) + 2\ H_2(g) \rightarrow 2\ H_2O(g)$	$\Delta H = -483.6$ kJ
$C(s) + 2\ H_2(g) \rightarrow CH_4(g)$	$\Delta H = -74.8$ kJ

Reaction Rates

- Interpret and express the meaning of the rate of a chemical reaction.
- Calculate the initial rate of reaction using experimental data.
- Explain how the rate of a chemical reaction is influenced by the temperature, concentration, particle size of reactants, and catalysts using collision theory.
- Interpret potential energy diagrams to find the activation energy and enthalpy change for a chemical reaction.
- Explain how activation energy affects reaction rate.

Questions:

81. A student drops a cube of sugar that has a mass of 15 g into a 2.0 L sample of water at 20°C. The student then observes the rate at which the sugar dissolves.
    a. Identify **two** changes that the student could make to the materials to increase the rate at which the sugar dissolves in the water.

b. Explain why **each** of these two changes would increase the rate at which the sugar dissolves in water.

82. Many laboratory preparations of solutions call for stirring the solvent while adding the solute. Which of the following is always an effect of this procedure?

A. It decreases the reactivity of the solute.
B. It decreases the solubility of the solute.
C. It brings the solute and solvent rapidly into contact.
D. It produces a double displacement reaction.

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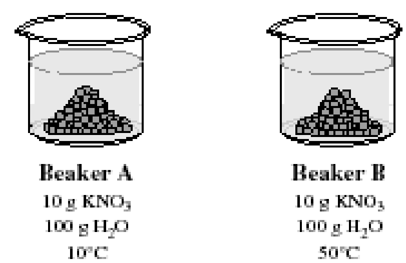
83. A student pours mineral salts into a bottle of cold water. Which of the following **best** explains why shaking the bottle will affect the dissolving rate of the salt?

A. Shaking exposes the salts to the solvent more quickly.
B. Shaking helps more water to evaporate.
C. Shaking causes more ions to precipitate out of solution.
D. Shaking equalizes the water temperature.

84. A data table and two prepared beakers are shown to the right: Solid KNO_3 was added to each beaker. Each beaker was stirred at the same rate until all of the solid dissolved. The table shows the solubilities of KNO_3 at different temperatures. How will the rates of dissolving KNO_3compare?

A. Beaker B dissolves faster because of more surface area.
B. Beaker A dissolves faster since the water molecules are farther apart.
C. Beaker B dissolves faster since the overall KE is increased.
D. Beaker A and B dissolve at the same rate since the concentrations are the same.

Temperature (°C)	Solubility of KNO_3 in 100 g H_2O
10	22 g
20	33 g
30	48 g
40	65 g
50	84 g



Equilibrium

- Define chemical equilibrium.
- Explain the nature of the equilibrium constant.
- Write chemical equilibrium expressions and carry out calculations involving them.
- Discuss the factors that disturb equilibrium.
- Discuss conditions under which reactions go to completion.
- Describe the common-ion effect.
- Explain what is meant by solubility-constant products, and calculate their values.
- Calculate solubilities using solubility-constant products.

Questions

85. What is a property of a reaction that has reached equilibrium?

A. The amount of products is greater than the amount of reactants.
B. The amount of products is equal to the amount of reactants.
C. The rate of the forward reaction is greater than the rate of the reverse reaction.
D. The rate of the forward reaction is equal to the rate of the reverse reaction.

86. Given the reaction: $HNO_2(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + NO_2^-(aq)$, the equilibrium constant, Ka, is

A. $[HNO_2]/[H_3O^+][NO_2^-]$ C. $[NO_2^-]/[H_3O^+][NO_2^-]$
B. $[H_3O^+][NO_2^-]/[HNO_2]$ D. $[H_3O^+][HNO_2]/[NO_2^-]$

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87. Which of the following actions will produce a shift of the reaction to the left?

A. removing CH_4 from the system
B. increasing the pressure at which the reaction is performed
C. increasing the temperature at which the reaction is performed
D. increasing the amount of H_2 used.

88. In the equilibrium process shown below, hydrogen gas and iodine gas react toform hydrogen iodide: $H_{2(g)} + I_{2(g)} + \text{heat} \leftrightarrow 2HI_{(g)}$

An increase in which of the following would shift the equilibrium to the left?

A. amount of heat B. amount of I2 C. concentration of HI D. concentration of H2

89. The Haber process is used to convert atmospheric nitrogen into ammonia, NH_3, a compound used in fertilizers. The reaction is shown: $N_{2(g)} + 3H_{2(g)}$ <-- --> $2NH_{3(g)}$ + energy

How can the equilibrium of this reaction be shifted to the right?

A. by decreasing the mass of N2 B. by decreasing the mass of H2
C. by increasing the pressure on the system D. by increasing the volume of the system

90. The equation for a chemical reaction is shown: $C(s) + 2H_2(g) \leftrightarrow CH_4(g)$ + heat

Which of the following actions will produce a shift of the reaction to the left?

A. removing CH_4 from the system
B. increasing the pressure at which the reaction is performed
C. increasing the temperature at which the reaction is performed
D. increasing the amount of H_2 used

Acids and Bases

- Classify a solution as neutral, acidic, or basic, given the hydrogen ion or hydroxide ion concentration.
- Calculate the pH of a solution given the hydrogen-ion or hydroxide-ion concentration.
- Calculate the hydrogen-ion or hydroxide-ion concentration given the pH of a solution.
- Define and give examples of Arrhenius acids and bases.
- Classify substances as acids or bases, and identify conjugate acid-base pairs in acid-base reactions according to Bronsted-Lowry theory.
- Distinguish between strong and weak acids and bases using the extent of ionization and the dissociation constants.
- Demonstrate knowledge of neutralization reactions by writing and balancing complete equations.
- Explain the steps of a titration.

Questions:

91. The water from hot springs near the Ebeko volcano in the Pacific Ocean has a very low pH. A low pH indicates which of the following about the water?

A. It has no detectable H^+ or OH^- ions. B. It has equal concentrations of H^+ and OH^- ions.
C. It has high concentrations of H^+ ions. D. It has equal numbers of positive and negative ions.

92. Arrhenius acids form _______________ion (name and formula) in water.

93. According to Arrhenius theory, $Ca(OH)_2$ is considered a _____________ .

94. Complete the following reactions:

Neutralization: HBr + $Mg(OH)_2$ → ________ + _________
Bronsted-Lowry (pair and label) HS^- + NH_2^- → ________ + ________

95. The table below shows the pH values of samples of substances.

Substance	pH
Rainwater	5.8
Drain cleaner	14.0
Distilled water	7.0
Soda water	3.0

According to the table, which of these substances is basic?

A. rainwater B. drain cleaner
C. distilled water D. soda water

96. A student was assigned to take water samples from a lake near his home. He measured the pH of one of the water samples to be 6.0. Which of the following **best** describes this sample of water?

A. highly acidic B. slightly acidic
C. highly basic D. slightly basic

97. The table below shows pH values of some foods.

pH Values of Some Important Foods

Vegetables	pH	Citrus	pH	Dairy/Egg	pH	Starches	pH
Asparagus	5.6	Grapefruit	3.2	Butter	6.2	Bread (white)	5.5
Beans	5.5	Lemons	2.3	Cheese	5.6	Corn	6.2
Peas	6.1	Limes	1.9	Eggs (fresh)	7.8	Crackers	7.5
Spinach	5.4	Oranges	3.5	Milk	6.5	Potatoes	5.8

A patient has chronic indigestion due to an overproduction of stomach acid. Which foods should the patient avoid until the condition is resolved?

A. vegetables B. citrus C. dairy/egg D. starches

98. What is the pH of 0.1 M NaOH? What is the pOH? Acidic or basic?

99. How many milliliters of 4.00M NaOH are required to exactly neutralize 50.0 mL of a 2.00 M solution of HNO_3? A. 25.0 mL B. 50.0 mL C. 100.0 mL D. 200.0 mL

100. Given the reaction at equilibrium: $HSO_4^- + H_2O \leftrightarrow H_3O^+ + SO_4^{2-}$
The two Bronsted bases are

A. H_2O and H_3O^+ B. H_2O and SO_4^{2-} C. H_3O^+ and HSO_4^- D. H_3O^+ and SO_4^{2-}

Redox

- Compute the oxidation number of an atom of any element in a pure substance.
- Identify the oxidizing and reducing agent in a redox reaction .
- Distinguish between redox and non-redox reactions.
- Define oxidation and reduction in terms of a change in oxidation number and identify atoms being oxidized or reduced in redox reactions.
- Apply the half-reaction method to balance redox equations.
- Relate chemical activity to oxidizing and reducing strength.
- Explain a voltaic cell using a sketch, labeling the cathode, the anode, and the direction of electron flow.
- Compute the standard cell potential of a cell using standard electrode potentials
- Distinguish between electrolytic and voltaic cells.

Questions:

101. Given the oxidation reduction reaction: $Hg^{2+} + 2I^{-} \rightarrow Hg(l) + I_2(s)$
Which equation correctly represents the half-reaction for the reduction that occurs?

A. $Hg^{2+} \rightarrow Hg(l) + 2\ e^-$
B. $Hg^{2+} + 2\ e^- \rightarrow Hg(l)$
C. $2I^{-} \rightarrow I_2(s) + 2\ e^-$
D. $2I^{-} + 2\ e^- \rightarrow I_2(s)$

102. The oxidation number of nitrogen in N_2O is

A. +1 B. +2 C. -1 D. -2

103. In any oxidation-reduction reaction, the total number of electrons gained is

A. less than the total number of electrons lost
B. greater than the total number of electrons lost
C. equal to the total number of electrons lost
D. unrelated to the total number of electrons lost

104. An electrochemical cell was set up with lithium and tin (Sn) half cells.

Sketch the electrochemical cell, including the anode, cathode, salt bridge, and direction of electron flow. Calculate the cell potential and write the balanced equation.

105. Calculate E_{cell} for the reaction $3Ni + 2Cr^{3+} \rightarrow 3Ni^{2+} + 2Cr^{2+}$. Is the reaction spontaneous?