***FINAL REVIEW PRACTICE SOLUTIONS***

**INTRODUCTION TO CHEMISTRY AND MEASUREMENT**

1. Is the following set of data accurate, precise, both, or neither if the actual value is 23.4 mL?

20.4 mL 20.3 mL 20.1 mL 20.5 mL

The measurements are precise but not accurate.

2. Indicate the number of significant figures in the following:

a. 1040 3 b. 600 1 c. 12.40 4 d. 0.00230 3 e. 0.82 2

3. Solve each expression using correct number of significant figures and units.

a. 24.4g + 110g 130 g b. 145.30 mL – 12.8 mL 132.5 mL

c. 9.8 g / .123 L 80. g/L d. 1.4500 cm x 2.55 cm x 1.5 cm 5.5 cm3

4. Solve the following using dimensional analysis (factor label method).

a. Convert 23.5 yd3 to ft3. 635 ft3

b. A pound of coffee beans yields 50 cups (this number is obtained by counting) of coffee. How many milliliters of coffee can be obtained from 1.00 g of coffee beans? (1 cup = 240mL) 26.4 mL

c. Convert 23.50 m/s2 to km/min2. 84.60 km/min2

5. Convert the following to the indicated temperatures

a. -34 ̊C to ̊F -29 F b. 213 K to ̊C -60. C

6. Explain how a Bunsen burner works. Gas and air are controlled by knobs on the burner. The gas and air mix resulting in a smokeless flame that produces very high temperatures.

7. What is the most practical and accurate tool used to measure volume in a laboratory?

Graduated Cylinder

8. Explain how you would set up a laboratory apparatus for

a. heating a crucible of magnesium – ring stand with an iron ring supporting a clay triangle

b. heating a beaker of copper II oxide and water – ring stand with an iron ring supporting a wire gauze. The beaker will be placed on the gauze and supported by a second iron ring.

9. A student takes an object with an accepted mass of 200.00 grams and masses it on his own balance.  He records the mass of the object as 196.5 g.   What is his percent error? 1.8%

10. What value does each of the following metric unit represent?

a. centi .01 or 1 x 10-2 b. nano .000000001 or 1 x 10-9 c. kilo 1000 or 1 x 103 d. milli .001 or 1 x 10-3 e. micro .000001 or 1 x 10-6

**MATTER AND ENERGY**

1. Solve the following density problems using significant figures.

a. What is the density (g/mL) of a block of metal 6cm x 10cm x 100cm that weighs 6820 dg? 0.11366667 = 0.1 g/mL

What would be the volume (in L) of 2.8 kg of the metal? 24.63343 = 20 L

b. A block of metal has a mass of 42.3 g. It is placed in a graduated cylinder containing 56.0 cm3 of water. If the final volume is 75.1 mL, what is the density of the metal block? 2.21 g/mL

2. How would you separate…..

a. iron filings and pepper magnet b. salt and sand dissolve the salt and filter the sand then evaporate the water c. methanol and water distillation

3. What is the specific heat of water? 4.184 J/g C

4. Which requires more heat to warm from 22.0 ̊C to 85.0̊C if they have exactly the same mass, a sample of aluminum (c = .902 J/ g ̊C) or a sample of water. Explain. It takes more energy to heat one gram of water than 1 gram of aluminum because water has a higher specific heat.

5. A .0234 kg sample of copper cools from 53 ̊C to - 9.4 ̊C. What is the change in its energy if the specific heat of copper is .387 J/g ̊C? - 570 J

6. Distinguish between a physical and chemical change. A physical change is a change that results in a change in appearance, but not in chemical composition. A chemical change is a change in the chemical composition and thus a new substance is formed.

7. Classify the following as either a physical or chemical change:

a. spoiling of milk C b. softening glass to bend it into a new shape P

c. burning a piece of paper C d. rusting of a nail C

8. Identify the following as a pure substance, heterogeneous mixture, or homogeneous mixture.

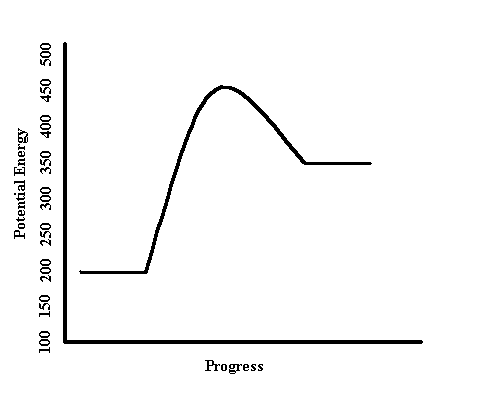
a. copper Pure substance - element

b. sweetened tea Homogeneous mixture - solution

c. calcium carbonate Pure substance - compound

d. sandy water – Heterogeneous mixture - suspension

9. What is the difference between an exothermic reaction and an endothermic reaction? Relate each reaction type to enthalpy. Endothermic reactions absorb energy and the products have more potential energy than the reactants. The enthalpy change for an endothermic reaction is positive. Exothermic reactions release energy and the products will have less potential energy than the reactants. Exothermic reactions have a negative enthalpy change.



10. Using the potential energy diagram above, answer the following questions.

a. How much energy do the reactants have? 200 kJ

b. How much energy is needed to get the reaction started? 250 kJ

c. What type of reaction is this, endothermic or exotherimic? endothermic

**CHEMICAL FOUNDATIONS**

1. What is the atomic mass of silver? What is silver’s mass number? 107.87 amu ; 108

2. Determine the symbol for the atom, ion, and or isotope that fits the description below.

a. 16 p, 18 e, 16 n 32S-2

b. 26 p, 26 e, 30 n 56Fe

c. 13 p, 10 e, 11 n 24Al+3

d. 35 p, 35 e, 45 n 80Br

3. Determine the identity of an element with a mass number of 75, that contains 42 neutrons. As

4. Determine the number of protons, neutrons, and electrons in the following:

a. Ar – 40 18p; 22n; 18 e b. F-1 9p; 10n; 10e c. 29 Al 13p; 16n; 13e d. Co 27p; 32n; 27e e. 56 Fe 26p; 30n; 26e

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5. Distinguish between an alpha and beta particle. Alpha decay is the emission of a helium nucleus (2 protons and 2 neutrons = alpha particle). Beta decay is the splitting of a neutron into a proton and an electron = beta particle. The proton is added to the nucleus of the atom, changing the identity of the atom.

6. If 36 grams of water are produced and 4 grams of hydrogen are used in its production, how many grams of oxygen is needed? What law is represented here? 32.0 grams of O2 = Law of the Conversation mass

7. What is the difference between the mass number for Carbon 14 and carbon’s average atomic mass of 12.011 amu? Mass number gives the total number of protons and neutrons for one isotope of carbon. The average atomic mass is the calculated sum of all of the isotopes of carbon and their % abundance in nature.

8. Calculate the atomic mass of lithium if one isotope has a mass of 6.0151 amu and a percent abundance of 7.59% and a second isotope has a mass of 7.0600 amu and a percent abundance of 92.41%. 6.9588 amu

9. Describe the debate between Democritus and Aristotle in terms of the atom.

Democritus – small pieces cannot be broken into smaller pieces forever – sooner or later, a single particle will exist that is so small that cannot be broken down any further. He calls this end particle “atomos”. He states that all matter consists of a collection of atoms and if there was space between the atoms, it contains nothing.

Aristotle – he believes that it makes more sense to believe that everything can be broken into smaller and smaller pieces forever. He questions that if matter is composed of particles and empty space, what holds the particles together? Due to Aristotle’s fame and popularity, he wins the debate.

10. List the four postulates to Dalton’s theory. Explain two reasons why it is no longer accepted.

1. All matter consists of tiny particles.
2. When elements react, their atoms combine in simple, whole number ratios.
   1. Atoms combine in simple whole number ratios to form compounds.
3. Atoms of an element are indivisible and indestructible.
4. All atoms of the same element have the same mass, atoms of different elements have different masses.

11. Compare and contrast Rutherford and Thomson’s experiments and atomic models.

Thomson – experiments using cathode ray tubes and discovers the electron. His model is referred to as the Plum Pudding Model where atoms are sphere shaped and composed of a space of positive charge. The negative electrons are found within this area of positive charge.

Rutherford – experiments using sheets of gold foil and alpha particles and discovers that the alpha particles are deflected back differently than what the plum pudding model would predict. His model is referred to as the nuclear model where the electrons are found in the empty space and the positive charge or protons are found in the center of the atom in what he refers to as the nucleus.

12. What significance does each subatomic particle have in regards to an element’s properties?

Protons – used to identify an individual atom

Neutrons – effect the average atomic mass of an atom

Electrons – determine an atom’s chemical behavior

13. Radioactive strontium (Sr - 90) has a half life of 29 years. If you had a 10.0 gram sample, how many years will it take you to get 2.5 grams of sample?

10.0 grams 🡪 5.0 grams = 29 years; 5.0 grams 🡪 2.5 grams = 29 years TOTAL = 58 years

14. What element undergoes alpha decay to form lead-208? Explain.

When a radioactive element undergoes alpha decay, it emits an alpha particle that is 2 protons and 2 neutrons (Helium nucleus). The element is Po – 212.

**MODERN ATOMIC THEORY/QUANTUM MECHANICS**

1. What is the wavelength (in m) of a photon of light containing 3.19 x 10-13 J of energy?

6.23 x 10 -13 m

1. How much energy in Joules will a photon of light with a wavelength of 345.0 nm possess?

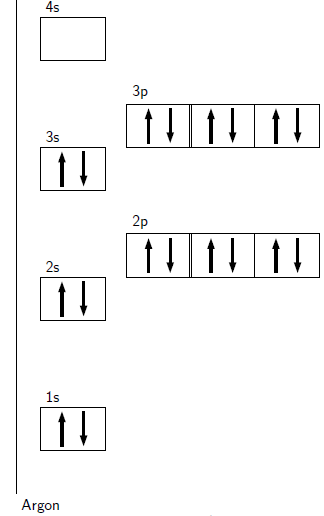
5.758 x 10-19 J

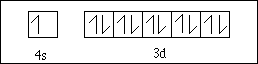
3. What is the maximum number of electrons that can occupy the following?

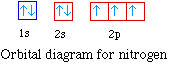
a. 3d 10 b. 3rd energy level 18 c. 4s 2 d. one orbital 2

e. 4th energy level 32

4. Draw the orbital diagrams for the following elements:

a. argon 

b. copper 

c. nitrogen 

5. Write the electron configuration for the following elements/ions:

a. calcium 1s22s22p63s23p64s2  b. arsenic 1s22s22p63s23p64s23d104p3

c. molybdenum 1s22s22p63s23p64s23d104p6 5s14d5

d. magnesium ion 1s22s22p6 e. sulfur ion 1s22s22p63s23p6

6. Write the noble gas configuration for the following elements.

a. Tin [Kr] 5s24d105p2 b. Gold [Xe] 6s1 4f14 5d10 c. Chromium [Ar] 4s13d5

7. Answer the following questions about the electron configuration of ***aluminum***:

a. What does the 3 in 3s2 mean? Third energy level

b. What does the 2 in 3s2 mean? Two electrons in sublevel

c. How many valence electrons does this atom have? 3

d. How many electrons will an atom of aluminum (lose or gain) to form an ion? Lose 3

e. What orbitals will aluminum lose its electrons from? 3s and 3p

8. What is needed for an electron to move from one energy level to a higher energy level?

A quanta of energy must be absorbed

9. What can be determined from an atom’s atomic emission spectra? The wavelength/frequency/energy of light emitted by an atom when electrons return to ground state.

10. What happens when an electron falls back to its ground state? It emits light of a specific wavelength/frequency/energy.

**THE PERIODIC TABLE**

1. Write the symbol of the element that fits each description:

a. alkali metal in period 3 Sodium

b. alkaline earth metal with 5 energy levels Strontium

c. the only metalloid chalcogen (oxygen group) Tellurium (Polonium is acceptable too)

d. halogen with four energy levels Bromine

2. How many valence electrons does each of the following possess? What ion does each form?

a. calcium 2; Ca+2 b. silver 1; Ag+1 c. phosphorus 5; P-3 d. lead 4; Pb+2/Pb+4  e. carbon 4; C-4

3. Which is smaller?

a. Ca or Ca+2 b. P or O c. N or F d. Sulfur ion or Aluminum ion

4. Does each of the following trends increase or decrease down a family or across (left to right) a period?

|  |  |  |
| --- | --- | --- |
|  | Down a Family | Across a Period |
| Electronegativity | Decreases | Increases |
| Ionization Energy | Decreases | increases |
| Atomic Radius | Increases | decreases |

5. Define electron affinity. Which elements have a negative value for electron affinity? Why?

Electron affinity is the energy change that occurs when a neutral atom acquires an electron. The more negative the value, the greater the amount of energy that is released. Elements that readily gain electrons like the halogens will have the most negative values because little energy is required to force an atom to acquire the electron.

Metals such as the alkaline earth metals will have positive values. Metals wish to lose electrons, not gain them. In order to add electrons to a metal, energy must be added forcing the addition of the electron.

6. Draw the Lewis dot diagrams for the following elements.

a.  b.  c. 

7. Why are the alkali metals one of the most reactive families of elements? They have low ionization energy.

8. Which groups make up the representative or main group elements? 1, 2, 13, 14, 15, 16, 17, and 18

9. Give three physical properties that differ between metals and nonmetals. Metals will lose electrons; are good conductors of electricity and heat; have luster; exist as solids; high boiling points and melting points. Nonmetals will gain electrons; are poor conductors of electricity and heat, lack luster, are normally found as gases or solid; are brittle; low boiling points and melting points.

10. What is a metalloid? List each of them. Elements that behave as both a metal and a nonmetal. Boron, Silicon, Germanium, Arsenic, Antimony, Tellurium, Polonium, and Astatine

**NOMENCLATURE**

1. Name each of the following:

1. Ba(CN)2
2. CoBr2
3. P2O5 (STOCK)
4. NaH
5. MgI2
6. Ca3(AsO4)
7. HCl
8. Sn(SO4)2
9. SO2 (CLASSIC)
10. BaSO3
11. Zn(ClO4)2
12. Be(BrO)2
13. HNO3 (aq)
14. Al(OH)3
15. HBrO4 (aq)
16. KMnO4
17. Hg3(PO3)2 (CLASSIC)
18. Na2O2
19. LiC2H3O2
20. FeCl3 ∙ 6 H2O (CLASSIC)
21. CCl4
22. AgNO3
23. N2O3 (STOCK)
24. Pb3N2 (STOCK)
    1. Barium Cyanide
    2. Cobalt (II) Bromide
    3. Phosphorus (V) Oxide
    4. Sodium Hydride
    5. Magnesium Iodide
    6. Calcium Arsenate
    7. Hydrogen Chloride
    8. Tin (IV) Sulfate
    9. Sulfur Dioxide
    10. Barium Sulfite
    11. Zinc Perchlorate
    12. Beryllium Hypobromite
    13. Nitric Acid
    14. Aluminum Hydroxide
    15. Perbromic Acid
    16. Potassium Permanganate
    17. Mercuric Phosphite
    18. Sodium Peroxide
    19. Lithium Acetate
    20. Ferric Chloride Hexahydrate
    21. Carbon Tetrachloride
    22. Silver Nitrate
    23. Nitrogen (III) Oxide
    24. Lead (II) Nitride

Write the formula for each of the following:

1. Diboron Hexahydride
2. Ammonium Chloride
3. Iron (III) Nitrate
4. Phosphoric Acid
5. Titanium (IV) Bromide
6. Cuprous Phosphide
7. Selenium (VI) Fluoride
8. Sodium Cyanide
9. Hydroselenic Acid
10. Phosphorus Triiodide
11. Barium Chloride Dihydrate
12. Potassium Hydroxide
13. Mercuric Sulfate
14. Auric Bromide
15. Cadmium Oxide
16. Calcium Hypochlorite
17. Lead (II) Phosphate
18. Magnesium Sulfate Heptahydrate
19. Silicon Dioxide
20. Sulfuric Acid
21. Aluminum Chlorate
22. Magnesium Sulfide
23. Plumbous Acetate
24. Manganese (III) Sulfite
25. B2H6
26. NH4Cl
27. Fe(NO3)3
28. H3PO4(aq)
29. TiBr4
30. Cu3P
31. SeF6
32. NaCN
33. H2Se(aq)
34. PI3
35. BaCl2 · 2H2O
36. KOH
37. HgSO4
38. AuBr3
39. CdO
40. Ca(ClO)2
41. Pb3(PO4)2
42. MgSO4 · 7H2O
43. SiO2
44. H2SO4(aq)
45. Al(ClO3)3
46. MgS
47. Pb(C2H3O2)2
48. Mn2(SO3)3