**ANSWERS**

**Ch.10 - Energy**

1. How is the concept of energy defined?

The ability to do work

1. What does temperature measure?

The average kinetic energy of particles in a sample of matter; the lowest temperature possible is 0K.

1. Explain what is meant by the terms exothermic and endothermic.

Exothermic is when a substance gives off heat, and in a reaction, the product has a lower potential energy than reactants; endothermic is when a substance absorbs heat, and in a reaction, the products have a higher potential energy than reactants

1. What is meant by the specific heat capacity of a material?

Specific heat is the amount of energy required to raise the temperature of one gram of that substance 1°C

5. Solve the following

a) What energy is required to heat 55.5 g of carbon from -10°C to 47°C (Carbon = 0.71 J/g °C)

55.5 g X 0.71 J/g°C X 57°C = 2,246 J

b) Calculate the mass (in grams) of iron that could be warmed from 25°C to 125°C by application of exactly 1.0 kJ of energy. (Ciron = 0.45 J/g °C)

1000J / (0.45 J/g°C\* 100°C ) = 22.2 g Fe

**Ch. 11 – Modern Atomic Theory**

1. How many s,p,d, and f orbitals can occupy any given energy level?

s-2, p-3, d-5, f-7

1. How many electrons can occupy an orbital?

Two electrons per orbital

1. What is the shape of a p-orbital? s-orbital?

The shape of the p-orbital is a dumbbell; the shape of the s-orbital is spherical

1. Which is the lowest energy level that can have a p orbital? d orbital? f orbital?

P orbital – Level 2, d orbital – Level 3, f orbital – Level 4

1. Is it possible for two electrons in the same atom to have exactly the same quantum numbers? What must be different about electrons that are in the same orbital?

No, no two electrons have the same quantum numbers

1. When is an atom’s electron configuration considered to be stable?

When the subshells (orbital’s) are full; like the Noble Gases

12. Distinguish between an atom in its ground state and an excited atom.

An atom is in the ground state if the electrons in the atom have the lowest possible energies

1. What happens to cause a photon of light to be emitted from an atom?

A photon of light emits when an electron goes from its excited state back to ground state; an eletron drops from a higher to a lower energy level

14. For the following elements list the electron configuration.

oxygen, cesium, krypton, titanium, scandium, nitrogen, chlorine

oxygen [He]1s2 2p4

cesium [Xe]6s1

krypton [Ar]4s23d104p6

titanium [Ar]4s23d2

scandium [Ar]4s23d1

nitrogen [He]2s22p3

chlorine [Ne]3s23p5

1. On the periodic table, in which directions does ionization energy increase?

Ionization energy increases as you move from bottom to top and from left to right. Fluorine has a very high ionization energy, Fr has a very low ionization energy. On the other hand, the atomic radaii increase from top to bottom on a column and decreases ionization energy.

**Ch. 12 – Chemical Bonding**

1. Define electronegativity

Electronegativity is the ability of an atom to attract electrons.

1. Does electronegativity increase or decrease as the atomic number of an element increases within the same period of the periodic table?

Electronegativity increases as atomic number increase. It increases from left to right across the periodic table.

1. How is the strength of a bond between two elements in a molecule related to their electronegativity?

If the difference in electronegativities of two different atoms is very big, they will have a stron bond. If the difference in electronegativitieis is very small, they will have a weaker bond.

1. What is the difference between an ionic and a covalent bond?

An ionic bond is between a metal and a non-metal and has a great electronegativity difference Two two atoms bond with electrostatic attraction between + and –

A covalent bond is between two non-metals with a very small electronegativity difference. They share electrons.

20. Referring to the table of electronegativity; classify each of the following bonds as either ionic (I) or covalent (C):

\_\_ Ia. Al-O \_C\_ b. Al-S \_C\_ c. Bi-Cl \_C\_ d. Bi-O \_C\_ e. C-Cl \_C\_ f. N-O

\_I\_ g. Na-S \_C\_ h. P-O \_C\_ i. S-O \_I\_ j. Ti-Br \_I\_ k. Ca-F \_I\_ l. Ba-S

1. What atoms (elements) form diatomic molecules?

Br2I2N2Cl2H2O2F2

1. Give an example of a polar covalent and a non-polar covalent molecule.

Polar covalent means two different atoms; example: Carbon – Hydrogen

Non-polar covalent means the same atoms; example – Nitrogen=Nitrogen

23. What does VSEPR stand for?

Valence Shell Electron Pair Repulsion model

1. What does the VSEPR theory predict?

This is a model for predicting the shape of molecules.

1. Draw the Lewis dot structure for SiO2

**O = S = O**

**Ch.13 - Gases**

1. Convert 1.20 atm to units of mm Hg, torr, and pascals.
   1. tm \* 760mmHg/1atm = 912 mmHg

1.20 atm \* 760 torr/1atm = 912 torr

1.20 atm \* 101,325 Pa/ 1atm= 121590 Pa

1. What does “STP” stand for? What conditions correspond to STP?

Standard Temperature and Pressure: Temperature is 0°C

Pressure: 1 atm = 760 mm Hg = 760 torr Hg = 14.69 PSI = 101,325 Pa

1. As temperature increases at constant pressure, what happens to volume?

A temperature increases, volume increases at constant pressure.

1. As volume increases at constant temperature, what happens to pressure?

As volume increasures, pressure decreases at constant temperature.

1. A sample of gas in a 10.0 L container exerts a pressure of 565 mm Hg. Calculate the pressure exerted by the gas if the volume is changed to 15.0 L at constant temperature.

. P1V1 = P2V2 10.0L \* 565 mmHg = 15.0L \* P2, P2 = 377 mmHg

1. A sample of gas in a 5.00 L container at 35.0°C is heated at constant pressure to a temperature of 70.0°C at constant pressure. Determine the volume of the heated gas.

V1/T1 = V2/T2 5.0 L / 308K = V2 / 343K V2 = 5.56L

1. A sample of gas at 24°C occupies a volume of 3.45 L and exerts a pressure of 2.10 atm. The gas is cooled to –12°C and the pressure is increased to 5.20 atm. Determine the new volume occupied by the gas.

P1V1/T1 = P2V2/T2 2.10atm\*3.45L/297K = 5.20atm\*V2/261K V2 = 1.22L

**Ch.14 – Liquids and Solids**

1. Define molar heat of fusion and molar heat of vaporization.

Molar heat of fusion is the energy required to melt (or freeze) one mole of a solid substance. Molar heat of vaporization is the energy required to vaporize (or condense) one mole of a liquid substance.

1. What is a dipole-dipole attraction? What is hydrogen bonding?

A dipole-dipole attraction is when molecules with dipoles line up and can attract to each other so that the positive and negative ends line up. Hydrogen bonding is a type of dipole-dipole attract that is very strong. It involved H of one atom being attracted to O, N or F or another atom.

1. Define London dispersion forces.

London dispersion forces is an instantaneous temporary dipole attraction that occurs when the electrons move around the nucleus. All molecules have a London Dispersion Forces.

1. What is vaporization? What is condensation?

Vaporization is the process of atoms or molecules escaping the surface of a liquid and becoming a gas.

Condensation is the process by which vapor molecules form a liquid.

37. Energy \_\_INCREASES\_\_\_\_\_\_\_\_\_\_\_ as a substance goes from liquid to gas.

38. Energy \_\_\_\_DECREASES\_\_\_\_\_\_\_\_\_ as a substance goes from liquid to solid.

39. How are kinetic energy and temperature related?

Temperature is the measure of the average kinetic energy of a substance.

40. Draw a heating curve for water as in increases in temperature from -20C to 120C.

Label all states of matter and each phase change.

**Ch.15 - Solutions**

41. Define homogeneous and heterogeneous mixtures.

A homogeneous mixture is a mixture of two or more substances in the same state.

A heterogeneous mixture is a mixture of two or more substances that are not in the same state (ex. oil and water, sand and water, etc.)

42. What is a saturated, unsaturated, and supersaturated solution?

Unsaturated is a solution that has not reached the limit of a solute that will dissolve in it.

Saturated is a solution that contains as much solute that will dissolve.

Supersaturated is when a solid is dissolved to the saturation limit at an elevated temperature and then allowed to cool, all of the solute may remain dissolved.

43. What is molarity?

Molarity is the concentration of solute in a solvent measured in mol/L

44. A chemist prepares some standard solutions for use in the lab using 500.0 mL volumetric flasks to contain the solutions. If the following masses of solutes are used, calculate the resulting molarity of each solution.

a. 4.865 g NaCl

* 1. g NaCl \*1mol/58.44g NaCl = .083 mol NaCl/.5L = .16M

b. 78.91 g AgNO3

78.91 g AgNO3 \* 1 mol/169.91g AgNO3 = .464 mol AgNO3/.5L = .92M

**Ch.16 – Acids and Bases**

45. What are the properties of acids and bases?

Acids are sour and can eat through metal (react.) Bases are bitter and slippery.

46. What is the pH range of an acidic solution?

The pH range of an acid is 0-6.

47. What is the pH range of a basic solution?

The pH range of a base is 8-14.

48. What an acid and base are mixed, what are the products?

Water and Salt

49. What are the Arrhenius and Bronsted-Lowry definitions of acids and bases?

An Arrhenius Acid produces hydrogen ions (H+) in aqueous solution.

An Arrhenius Base produces hydroxide (OH-) in aqueous solution.

A Bronsted-Lowry Acid is a Proton Donor.

A Bronsted-Lowery Base is a Proton Acceptor.

50. Label the reactants and products in the following reaction as an acid, base, conjugate acid, conjugate base

HCl + H2O →H3O+ + Cl-

Answer: HCl (acid) + H2O (base) →H3O+(conjugate acid) + Cl- (conjugate base)

52. Calculate the pH of a 0.00515 M HCl solution.

pH = -log [0.00515] = 2.29

53. Calculate the hydrogen ion concentration of a solution with a pH of 9.4

[H+] = 10-9.4 = 3.98 x 10-10

54. Calculate the pOH of a solution that is 4.96 x 10-11 H+

pH = -log[4.96x10-11] = 10.3. pH + pOH = 14 pOH = 14 - 10.3 = 3.7

**Ch.17 - Equilibrium**

55. What do we mean by an equilibrium position?

An equilibrium position is where the equilibrium lies. If there are more products than reactants, equilibrium will lie to the right. If there are more reactants, equilibrium will lie to the left.

56. What is a catalyst?

A catalyst affects a chemical reaction by providing an alternate pathway with a lower activation energy. A catalyst lowers the activation energy in a reaction.

57. What are the four factors affecting reaction rate?

Surface area, Heat, Stirring (movement of molecules) and Nature of the Reactants

58. Write the equilibrium constant expressions for each of the following reactions.

a. 2NO(g) + O2(g) ↔ 2NO2(g) [NO2]2/[NO]2[O2]

b. N2H4(l) + O2(g) ↔ N2(g) + 2H2O(g) [H2O]/N2H4][O2]

c. CO(g) + NO2(g) ↔ CO2(g) + NO(g) [NO][CO2]/[CO][NO2]

59. For the reaction: 2SO2(g) + O2(g) ↔ 2SO3(g) at a particular temperature the equilibrium system contains

[SO3(g)] = 0.42 M, [SO2(g)] = 1.4 x 10-3 M, and [O2(g)] = 4.5 x 10-4 M.

Calculate K for the process.

[SO3]2/[SO2]2[O2] = [.42]2/[1.4x10-3]2[4.5x10-4] = 2.0 x 108

60. Explain the collision model for chemical reactions. How does the collision model account for the observation that higher concentrations and higher temperatures tend to make reactions occur faster?

The collision model simply states that molecules or atoms must collide with enough force in order to react. If there are higher temperatures, molecules or atoms move faster therefore colliding with more force and reacting more. If there is a higher concentration, molecules or atoms collide more often increasing the chance of a collision in which a reaction occurs.

61. What is Le Chatelier’s Principle?

LeChatelier’s Principle states when a change is imposed on a system at equilibrium, the position of the equilibrium shifts in a direction that tends to reduce the effect of that change

62. Suppose the reaction: 2SO2(g) + O2(g) ↔2SO3(g) has already reached equilibrium. Predict the effect of each of the following changes on the position of the equilibrium:

a. Additional SO2(g) is added to the system. Shifts right

b. The SO3(g) is liquefied and removed from the system. Shifts right

c. A very efficient catalyst is used. No change in equilibrium

d. The volume of the container is drastically reduced. Shifts right

**Ch. 18 – Redox**

**63.** Which of the following are oxidation-reduction reactions?

a. PCl3 + Cl2 → PCl5

This IS a Redox.

b. Cu + 2AgNO3 → Cu(NO3)2 + 2Ag

This IS a Redox.

c. CO2 + 2LiOH → Li2CO3 + H2O

This IS NOT a Redox.

d. FeCl2 + 2NaOH → Fe(OH)2 + 2 NaCl

This IS NOT a Redox.

64. What is the oxidation state of nitrogen in each of the following?

a. HNO3 b. NH4Cl c. N2O d. NO2 e. NaNO2

a. + 5

b. - 3

c. + 4

d. + 3

**Ch.20 – Organic Chemistry**

65. What are alkanes, alkenes, and alkynes and what suffixes do they have?

They are all saturated hydrocarbons.

Alkanes – have single bonds; suffix is “-ane”

Alkenes – have one or more double bonds; suffix is “-ene”

Alkynes – have one or more triple bonds; suffic is “-yne”

66. What is a polymer and monomer?

A polymer is repeated subunits of monomers (i.e. proteins, carbohydrates, nucleic acids)

A monomer is a single subunit (i.e. amino acid, monosaccharide)

Polymers are chain-like molecules built from small molecules called monomers.

67. Proteins are polymers. What are the monomers that proteins are made from?

Amino Acids

68. Why is carbon a good building block for organic molecules?

Carbon can form long chains, forming single and double bonds readily due to it having four valence electrons.

69. How would you identify each of the following?

Alcohols

Ketones

Aldehydes

Ethers

Carboxylic acids

Aromatics

70. Name the following.

4-butyle-3-isopropyl-6-methylheptane