Chemistry 20 Final Exam Review Questions

**Unit A: Chemical Bonding**

1. What are the forces that keep molecules together called?
2. What are the forces that keep atoms within a molecule together called?
3. Ionic bonds
   1. Describe ionic bonds:
   2. List properties of ionic compounds:
4. Covalent bonds
5. Describe covalent bonds:
6. List and describe the types of covalent bonds:
7. London Dispersion Forces
8. Describe London Dispersion Forces (LDF).
9. What substances exhibit London Dispersion force?
10. Explain why larger molecules have greater LDFs than smaller molecules.
11. Dipole-Dipole Forces
12. Describe intermolecular dipole-dipole forces.
13. What substances exhibit dipole Forces?
14. Hydrogen Bonding
15. Describe hydrogen bonding.
16. What elements must hydrogen be bonded to in order for hydrogen bonds to occur?
17. If substance A has a higher melting and boiling point than substance B, what can you say about the substances’ intermolecular forces?
18. Nitrogen has 7 protons and 7 electrons, sulfur has 16 protons and 16 electrons. Which of the two have a greater London dispersion forces? Explain
19. Describe metallic bonding.
20. Describe the bonding structure of network covalent structures.
21. Rank the IMFs from strongest to weakest.
22. Complete the following table on atom characteristics.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| **Atomic**  **Number** | **Atom**  **Symbol** | **Group**  **Number** | **Number of Valence Electrons** | **Number of Occupied Energy Levels** | **Lewis Diagram of atom** | **Number of Lone Electron Pairs** | **Number of Bonding Electrons** |
|  | S |  |  |  |  |  |  |
|  | Si |  |  |  |  |  |  |
|  | P |  |  |  |  |  |  |
|  | Cl |  |  |  |  |  |  |
|  | Br |  |  |  |  |  |  |
|  | Ge |  |  |  |  |  |  |
|  | H |  |  |  |  |  |  |
|  | C |  |  |  |  |  |  |
|  | N |  |  |  |  |  |  |
|  | O |  |  |  |  |  |  |

1. Complete the following table on atom characteristics.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| **Molecule or Ion Formula** | **# of Bonding Electrons** | **# of Lone Pairs** | **Lewis Structure** | **VSEPR Diagram** | **Name of VSEPR Shape** |
| **CF4** |  |  |  |  |  |
| **PH3** |  |  |  |  |  |
| **H2S** |  |  |  |  |  |
| **CO2** |  |  |  |  |  |
| **HF** |  |  |  |  |  |
| **SO2** |  |  |  |  |  |
| **SO32-** |  |  |  |  |  |
| **SO42-** |  |  |  |  |  |

**Unit B: Gases**

1. Convert the following:
   1. 235 torr to kPa
   2. 180 kPa to mm Hg
   3. 2.34 atm to kPa
   4. 24ºC to Kelvin
   5. 987 K to ºC
2. Give the conditions for STP.
3. Give the conditions for SATP.
4. An unknown gas has a pressure of 469 mm Hg and occupies 29.0 ml. What would be the new volume if the pressure was changed to 0.998 atm?
5. What would be the initial pressure (in mm Hg) of 80.9 L hydrogen gas if it was changed to 660 torr and had a final volume of 89.0 L?
6. What does absolute zero mean?
7. What would be the new volume of oxygen gas when pressure remains constant if the temperature changed from 39.0 ºC to 55.0 ºC and its initial volume was 685 ml?
8. A sample of nitrogen gas exerts 52.6 kPa at 66.0 ºC. What pressure would the gas exert at 99.0ºC if the container’s volume remains the same?
9. Helium gas in a hot air balloon experiences a temperature change from 21.0ºC to 55.0 ºC and an atmospheric pressure change from 100 kPa to 88.5 kPa. What would be the new volume of the hot air balloon if its initial volume was 125 kiloliters?
10. A 2.7-liter sample of nitrogen gas is collected at a temperature of 45 ºC and a pressure of 0.92 atm. What pressure would have to be applied to the gas to reduce its volume to 2.0 liters at a temperature of 25.0 ºC?
11. A sample of argon gas occupies a volume of 2.0 L at -35ºC at 1.2 atm. What would its Celsius temperature be at 2.0 atm if its volume decreases to 1.5 L?
12. What pressure would 2.00 Kmol of fluorine gas exert under 45ºC with a volume of 985 ml?
13. What would be the volume of 3.52 mg of chlorine gas at 21.0ºC under 99.2 kPa of pressure?

1. What volume of oxygen gas would occupy at STP that has a mass of 9.45 grams?
2. Oxygen gas and magnesium react to form 2.43 g of magnesium oxide. What volume of oxygen gas at 94.9 kPa and 25.0ºC would be consumed to produce this amount of MgO(s)?
3. Nitrogen triodide decomposes into explosive nitrogen gas and iodine. Calculate the volume of the gas produced at STP when 395 mg of NI3 (g) decomposes.
4. A 2.00 liter of sample of ethane, C2H6, is burned at 1.00 atm and 25.0 ºC with an excess of oxygen. What mass of water vapor will be produced from the burning of ethane? What would be the volume of the water vapor under the same conditions?
5. If 15 L of methane gas undergoes complete combustion, what volume of oxygen gas will be required? What volume of carbon dioxide will be produced?

**Unit C: Solutions, Acids, and Bases**

1. Define the following
   1. Solute

Solute is the substance that gets dissolved.

* 1. Solvent

Solvent is the substance that does the dissolving.

* 1. Saturated solution

A solution that contains the maximum amount of solute at a given temperature.

* 1. Miscible

me type of intermolecular forces) in all proportions.

* 1. Immiscible

A

liquid that will not form a solution with another liquid in any proportion.

* 1. Endothermic reaction

A reaction that absorbs energy.

* 1. Exothermic reaction

A reaction that releases energy.

* 1. Solubility Equilibrium

1. What is the difference between gases and solids in terms of temperature and solubility?

Generally, as the temperature increases, solids increase in solubility, because the solvent particles move more qui

ckly and can surround each solute with fewer solvent particles. However, as the temperature increases, gases become less solSolubility is measured in g/100 mL solution at a given temperature.

1. What is the molar concentration of a solution that has 4.4 grams of sucrose in 990 ml of solution?
2. What is the volume of a 0.750 mol/L solution containing 2.50 grams of salt (sodium chloride)?
3. How many liters of 1.50 mol/L solution of magnesium hydroxide would contain 40.0 g of solute?
4. Sodium phosphate solution is used to remove the scales at the bottom of a tea kettle. Calculate the mass of sodium phosphate needed to make 4.00 L of a 0.500 mol/L cleaning solution.
5. Express the following in % (v/v) or % (m/v).
6. 6.0 g of potassium chromate in 100 mL solution
7. 0.75 ml of methanol in 12.0 ml of solution
8. 4.5 g of calcium chloride in 0.0500 L of solution
9. What mass or volume of solute is contained in the following solutions?
10. 150.0 ml of 2.2% (m/v) solution containing sulfuric acid
11. 50.0 ml of 6.5% (m/v) solution of ammonium phosphate
12. 89.0 ml of 75% (v/v) solution of ethanol
13. What volume would be needed for the following?
14. 3.6 g from a 5.5% (m/v) solution of sodium acetate
15. 25.0 ml from a 10.5% (v/v) solution of hydrogen peroxide
16. 9.6 g from a 30% (m/v) solution of calcium chloride
17. What is the % (m/v) concentration of 2.50 mol/L of hydrogen peroxide?
18. What is the % (m/v) concentration of 433 ppm of sodium chloride?
19. You are diluting a solution and you want to prepare 0.500 mol/L at 200.0 ml from a stock solution of 1.00 mol/L, what volume do you need from your stock solution?
20. Commercial sulfuric acid is 17.6 mol/L. If one uses 170 ml of the concentrated acid and dilutes it to a volume of 1.50 L, what is the concentration of the diluted sulfuric acid solution?
21. Calculate the mass of solute needed to prepare a 500 ml of a 0.800 mol/L solution of potassium chromate.
22. Suppose the solution in the previous question was the stock solution and you wanted to make a secondary solution that is 250.0 ml that is 0.4500 mol/L in concentration, what will be the volume needed to make it?
23. Outline the process of making a standard solution from a mass.
24. Write the dissociation equation for the following and give the ion concentrations for each ion.
25. 2.50 mol/L of calcium nitrate
26. 3.00 mol/L of cobalt (III) acetate
27. 1.25 mol/L of zinc chloride
28. 0.750 mol/L of ammonium sulfate
29. 0.250 mol/L of acetic acid
30. Give three properties each of acids and bases.
31. Define the following:
32. Acid according to Arrhenius
33. Base according to Arrhenius
34. Acid according to modified Arrhenius theory
35. Base according to modified Arrhenius theory
36. What is the difference between strong acids and weak acids?
37. What is the difference between strong bases and weak bases?
38. Why is an equilibrium arrow given in some of the equations for acids and bases?
39. Why is a straight arrow given in some of the equations for acids and bases?
40. What is the difference between a diluted acid and a concentrated acid?
41. Write modified Arrhenius equations for the following acids:
42. boric acid
43. acetic acid
44. sulfurous acid
45. carbon dioxide
46. hydrobromic acid
47. Write modified Arrhenius equations for the following bases:
48. sodium carbonate
49. strontium oxide
50. sodium hydrogen sulfate
51. lithium hydroxide
52. calcium oxide
53. Calculate the pH for the following hydronium ion concentrations:
54. 0.02 mol/L
55. 2.5 mol/L
56. 0.0065 mol/L
57. 9.67 x 10-9 mol/L
58. 0.0874 mol/L
59. Calculate the pH for the following hydroxide ion concentrations:
60. 0.36 mol/L
61. 0.559 mol/L
62. 1.90 x 10-5 mol/L
63. 3.9 mol/L
64. 0.004 mol/L
65. Calculate the hydronium ion concentrations for the following pH or pOH values:
66. pH = 2.45
67. pH = 6.550
68. pOH = 2.990
69. pH = 8.35
70. Predict the pH of the solution made by dissolving 925 mg of nitric acid in enough water to make 850 mL.
71. Predict the pH of the solution prepared by dissolving 3.589 grams of magnesium hydroxide in 1.50 liters of water (assume all the magnesium hydroxide dissolves).
72. How much solvent will you have to add to 150 ml of perchloric acid that has a pH of 3.500 to change it to a pH of 4.000?
73. A student wants to make a solution with a pH of 7.900. What mass of sodium hydroxide will the student require to make 4.50 L of solution?

**Unit D: Quantitative Analysis**

1. What color of flame does the following show?
2. Lithium
3. Sodium
4. Lead
5. How many moles of hydrogen gas are produced if 0.500 mol of water is decomposed?
6. Sulfur reacts with barium oxide to produce barium sulfide and oxygen gas.
   1. How many moles of elemental sulfur are needed if 2.00 mol of barium oxide are used?
   2. How many moles of barium sulfide are produced from 0.100 mol of sulfur?
7. The combustion of methane gas takes place in the presence of oxygen gas to produce carbon dioxide and water vapor (the compounds produced when you burn a hydrocarbon)
8. How many moles of oxygen gas are needed to completely burn 3.00 mol of methane gas?
9. How many moles of water vapor are produced from 0.0400 mol of methane gas?
10. Iron (II) phosphate reacts with tin (IV) nitride to produce iron (II) nitride and tin (IV) phosphate.
11. How many moles of tin (IV) nitride are needed to produce 0.500 mol of iron (II) nitride?
12. How many moles of iron (II) phosphate are used when 0.045 mol of tin (IV) nitride reacts?
13. When 6.5 mol of potassium chlorate solid breaks down to the simpler compound of potassium chloride and oxygen gas, what mass of KCl (s) would be produced?
14. When an excess of silver reacts with 3.45 moles of zinc phosphate, what mass of silver phosphate would be produced?
15. In neutralization of sulfuric acid solution, 4.56 g of sodium hydroxide was used. What mass of water would be produced in this reaction?
16. When iron (II) hydroxide reacts with cobalt (II) phosphate, iron (II) phosphate and cobalt (II) hydroxide are formed. If 3.00 g of iron (II) hydroxide react, what mass of cobalt (II) phosphate is needed?
17. Calcium acetate is added to 10.0 ml of 0.0200 mol/L silver nitrate. What is the mass of the precipitate produced?
18. Predict the volume of the base (0.250 mol/L of sodium hydroxide) needed to be added to neutralize 75.0 ml of nitric acid at 0.500 mol/L.
19. A 150.0 ml sample of 0.7500 mol/L of aqueous solution of sodium sulfate is added to 250.0 ml of strontium hydroxide at 0.350 mol/L. Predict the mass of the precipitate.
20. What volume of boric acid at 0.025 mol/L will be needed to neutralize 90 ml of 0.055 mol/L magnesium hydroxide?
21. For each of the following questions indicate which reactant is the limiting reagent, and which is the reagent in excess. You may need to complete a balanced chemical equation before answering the question.
    1. 25 grams of zinc reacts with 75 g of nitric acid. What mass of zinc nitrate is produced?
    2. 15 g of sodium hydroxide reacts with 26 g of hydrochloric acid. What mass of sodium chloride is produced?
    3. 2.0 g of methane reacts with 7.0 g of oxygen. How many grams of carbon dioxide is produced?
22. When 84.8 g of iron(III) oxide react with an excess of carbon monoxide, 57.8 g of iron is produced.

Fe2O3(s) + 3CO(g) 🡪 2Fe(s) + 3CO2(g)

What is the percent yield of this reaction?

1. When 50.0 g of silicon dioxide are heated with an excess of carbon, 32.2 g of silicon carbide are produced.

SiO2(g) + 3C(s) 🡪 SiC(s) + 2CO(g)

What is the percent yield of this reaction?

1. What is the percent yield if 3.74 g of copper are produced when 1.87 g of aluminum are reacted with 14.5 g of copper (II) sulfate?
2. 50.0 ml of 0.250 mol/L potassium phosphate reacts with 25.0 ml of 1.00 mol/L lead (II) sulfate. What is the mass of the precipitate?
3. 75 ml of 1.25 mol/L of HCl (aq) is reacted with 125 ml of 1.00 mol/L KOH (aq), what is the concentration of potassium chloride produced?
4. Use the following titration data to determine the concentration and pH of the sulfuric acid.

*Titration of 20.00 ml sample of H2SO4 (aq) with 0.550 mol/L magnesium hydroxide*

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Trial** | **1** | **2** | **3** | **4** |
| **Initial Burette Reading (mL)** | 52.6 | 38.6 | 27.1 | 15.7 |
| **Final Burette Reading (mL)** | 38.6 | 27.1 | 15.7 | 4.1 |
| **Volume of Titration (mL)** |  |  |  |  |