Exam 1

# Part 1: Multiple Choice (2 points each)

## Directions: Please circle the *best* answer for each of the following questions.

1. The equilibrium constant for the formation of calcium carbonate from the ions in solution is 2.2 × 108 according to the reaction: Ca2+ (aq) + CO32- (aq) CaCO3 (s). What is the value of the equilibrium constant for the reverse of this reaction?
	1. The same, 2.2 × 108
	2. -2.2 × 108
	3. 2.2 × 10-8
	4. 4.5 × 10-9
	5. 1
2. The equilibrium constant for the acid ionization of mercaptoethanol is 1.91 × 10-10:

HSCH2CH2OH (aq) H+ (aq) + SCH2CH2OH- (aq)

A solution of mercaptoethanol in water

* 1. is almost entirely ionized.
	2. is almost entirely unionized.
	3. is about one-half ionized.
	4. is a strong acid.
	5. is completely dissociated.
1. Which of the following titrations result in a basic solution at the equivalence point?
	1. HI titrated with NaCH3CO2
	2. HOCl titrated with NaOH
	3. HBr titrated with KOH
	4. Pb(NO3)2 titrated with NaI
	5. none of the above
2. pH indicators
	1. are weak acids.
	2. have characteristic colors in their various protonated and deprotonated forms.
	3. have characteristic pKHIn values.
	4. all of the above
	5. none of the above
3. What is the most effective method for avoiding exposure by ingestion?
	1. Taste only chemicals that your instructor gives you permission to taste.
	2. Taste only chemicals that you know are nontoxic.
	3. Never eat or drink anything while in a chemistry lab.
	4. Only eat food in the lab when you know that it cannot be contaminated with toxic chemicals.
	5. all of the above
4. Phosphoric acid is a triprotic acid, which ionizes in sequential steps:

H3PO4 (aq) + H2O (l) H2PO4- (aq) + H3O+ (aq) Ka1

H2PO4- (aq) + H2O (l) HPO42- (aq) + H3O+ (aq) Ka2

HPO42- (aq) + H2O (l) PO43- (aq) + H3O+ (aq) Ka3

Which equilibrium is most important in determining the pH of a solution of sodium phosphate?

* 1. H3PO4 (aq) + H2O (l) H2PO4- (aq) + H3O+ (aq)
	2. H2PO4- (aq) + H2O (l) HPO42- (aq) + H3O+ (aq)
	3. HPO42- (aq) + H2O (l) PO43- (aq) + H3O+ (aq)
	4. PO43- (aq) + H2O (l) HPO42- (aq) + OH- (aq)
	5. H2PO4- (aq) + H2O (l) H3PO4 (aq) + OH- (aq)
1. The chemical formula for sulfurous acid is
2. H2SO3 (aq)
3. H2S (aq)
4. H­2SO4 (aq)
5. H2S2O7 (aq)
6. SO3 (aq)
7. Which of the following species cannot act as a Lewis base?
	1. CH4
	2. O2-
	3. H2O
	4. NH3
	5. none of the above

The following plot shows two titration curves, each representing the titration of 50.00 mL of 0.100 M acid with 0.100 M NaOH.

1. Which point a-d represents the equivalence point for the titration of a weak acid?
	1. a
	2. b
	3. c
	4. d
	5. none of the above
2. At which point a-d is the pKa of the acid equal to the pH?
	1. a
	2. b
	3. c
	4. d
	5. none of the above

# Part 2: Short Answer

## Directions: Answer each of the following questions. Be sure to use complete sentences where appropriate. For full credit be sure to show all of your work.

1. The interhalogen compound ClF3 is prepared in a two-step fluorination of chlorine gas (5 points):
	* 1. Cl2 (g) + F2 (g) ClF (g)
		2. ClF (g) + F2 (g) ClF3 (g)
2. Balance each step and write the overall equation.

Cl2 (g) + F2 (g) 2 ClF (g)

+ (ClF (g) + F2 (g) ClF3 (g) ) **× 2**

Cl2 (g) + 3 F2 (g) 2 ClF3 (g)

1. Show that the overall Qc equals the products of the Qc’s for the individual steps.
2. Dintrogen trioxide decomposes into NO and NO2 in an endothermic process (∆rH° = 40.5 kJ/mol rxn). N2O3 (g) NO (g) + NO2 (g)

Predict the effect of the following changes on the position of equilibrium; that is, state which way the equilibrium will shift (left, right or no change) when each of the following changes is made (5 points):

* 1. Adding more N2O3 (g)  \_\_\_\_\_\_\_\_right
	2. Adding more NO2 (g) \_\_\_\_\_\_\_\_left
	3. Increasing the volume of the reaction flask \_\_\_\_\_\_\_\_right
	4. Lowering the temperature \_\_\_\_\_\_\_\_left
	5. Adding some He (g) \_\_\_\_\_\_\_\_no change
1. Which should be the stronger acid, cyanic acid, HOCN, or hydrocyanic acid, HCN? Explain briefly. (In HOCN, the H+ ion is attached to the O atom of the OCN- ion.) (3 points).

Cyanic acid should be a stronger acid than hydrocyanic acid because the H atom in HOCN is attached to the more electronegative O atom. The electron attachment enthalpy of OCN is thus more negative than that of C, which stabilizes the conjugate base that forms and makes the ionization of HOCN more product-favored.

1. Consider the reaction xenon and fluorine gases to produced xenon tetrafluoride gas. A reaction mixture initially contains 2.24 atm xenon and 4.27 atm fluorine gases (10 points).
	1. If the equilibrium pressure of xenon is 0.34 atm, find the equilibrium constant, Kp, for the reaction.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  | Xe (g) | + |  2 F2 (g)  |  | XeF4 (g) |
| I | 2.24 atm |  | 4.27 atm |  | 0 atm |
| C | -x |  | -2x |  | +x |
| E | 2.24 atm – x = 0.34 atmx = 1.90 atm  |  | 4.27 atm – 2 x =4.27 atm – 2(1.90 atm) =0.47 atm |  | 1.90 atm |

* 1. What is the value of the equilibrium constant, Kc, if the reaction takes place at 215 °C?
1. What is the key structural feature of all Brønsted-Lowry bases? How does this feature function in an acid-base reaction (3 points)?

All Brønsted-Lowry bases contain at least one lone pair of electrons, which binds an H+ and allows the base to act as a proton acceptor.

1. How many moles of H3O+ or OH- must you add to 5.6 L of HA solution to adjust its pH from 4.52 to 5.25? Assume a negligible volume change (6 points).

The solution increases in pH therefore OH- was added.

1. For each of the following cases, decide whether the pH is less than 7, equal to 7, or greater than 7. Explain your answer (6 points).
	1. Equal volumes of 0.10 M acetic acid, CH3CO2H, and 0.10 M potassium hydroxide.

The reaction produces acetate ion, the conjugate base of acetic acid. The solution is weakly basic. pH is greater than 7.

* 1. 25 mL of 0.015 M ammonia is mixed with 25 mL of 0.015 M hydrochloric acid.

The reaction produces NH4+, the conjugate acid of NH3. The solution is weakly acidic. pH is less than 7.

* 1. 150 mL of 0.20 M nitric acid is mixed with 75 mL of 0.40 M sodium hydroxide.

The reaction mixes equal molar amounts of strong base and strong acid. The solution will be neutral. pH will be 7.

1. A buffer consists of 0.22 M KHCO3 and 0.37 M K2CO3. Carbonic acid is a diprotic acid with Ka1 = 4.5 × 10-7 and Ka2 = 4.7 × 10-11 (12 points).
	1. Which Ka value is more important to this buffer? \_\_\_\_\_\_Ka2, because the buffer solution is between HCO32- and CO32-, so the second proton of carbonic acid, HCO3, is undergoing chemical reaction. If the buffer solution was created using carbonic acid and potassium hydrogen carbonate Ka1 would be more important. \_\_\_\_\_\_\_\_\_\_
	2. What is the buffer pH? \_\_\_\_\_\_pH = 10.55\_\_\_\_\_\_\_\_

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | HCO3- (aq) +  | H2O (g)  | CO32- (aq) | H3O+ (aq) |
| I | 0.22 M | n/a | 0.37 M | ~0 M |
| C | -x | n/a | +x | +x |
| E | 0.22 M – x =0.22 M – 2.8 × 10-11 M ≈0.22 M | n/a | 0.37 M + x = 0.37 M + 2.8 × 10-11 M ≈ 0.37 M | x = 2.8 × 10-11 M  |

Check approximation:

Check math:

Or using the Henderson-Hasselbach equation:

1. In an experiment similar to the **Determination of Ka, Kb, and Percent Ionization from pH** the experimental pH a 0.25 M NH4Cl solution is 4.89. From CRC Handbook theoretical pKa is 9.24 and (20 points).
	1. Calculate the theoretical Ka
	2. Calculate the [H+].
	3. Calculate the pOH.

* 1. Calculate the [OH-].

* 1. Write the balanced net ionic equilibrium reaction. Identify the acid, base, conjugate acid, and conjugate base. Then calculate the equilibrium concentrations of all species using an ICE table.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
|  | NH4+ (aq) + | H2O (l)  | H3O+ (aq) + | NH­3 (aq) |
|  | acid | base | conjugate acid | conjugate base |
| I | 0.25 M |  | ~0 M | 0 M |
| C | -x |  | +x = 1.3 × 10-5 M | +x  |
| E | 0.25 M - 1.3 × 10-5 M = 0.25 M  |  | 1.3 × 10-5 M | 1.3 × 10-5 M |

* 1. Calculate the experimental Ka and pKa.
	2. Calculate the percent dissociation.

* 1. What is the percent error of Ka?

|  |  |
| --- | --- |
| Volume of OH- added (mL)  | pH |
| 0.0 | 3.09 |
| 5.0 | 3.65 |
| 10.0 | 4.10 |
| 15.0 | 4.50 |
| 17.0 | 4.55 |
| 18.0 | 4.71 |
| 19.0 | 4.94 |
| 20.0 | 5.11 |
| 21.0 | 5.37 |
| 22.0 | 5.93 |
| 22.2 | 6.24 |
| 22.6 | 9.91 |
| 22.8 | 10.20 |
| 23.0 | 10.40 |
| 24.0 | 10.80 |
| 25.0 | 11.00 |
| 30.0 | 11.50 |
| 40.0 | 11.80 |

1. A 125.0 mg sample of an unknown, monoprotic acid was dissolved in 100.0 mL of distilled water and titrated with a 0.050 M solution of NaOH. The pH of the solution was monitored throughout the titration, and the following data were collected (10 points).
	1. Complete the following graph.
	2. Indicate the major species at the half equivalence point, equivalence point, and the three-half equivalence point.

Half equivalence point: HA (aq), H2O (l), Na+ (aq), A- (aq)

Equivalence point: Na+ (aq), A- (aq), H2O (l)

Three-half equivalence point: Na+­ (aq), A- (aq), OH- (aq), H2O (l)

* 1. Determine the Ka of the acid.

Equivalence point at 22.5 mL

At half-equivalence point pH = pKa. VNaOH = 11.25 mL, pH = 4.20 = pKa = 4.20

* 1. What is the molar mass of the unknown acid?