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Chemistry 142 Exam #2 Preparation

Part I

Directions. You will work out the problems on this page then enter them on the computer. Once you enter them into the computer you can submit your work for grading. You can re-enter (and re-score) the problems as many times as you wish.

Question 1. A weak acid, HA, has a pK_a of 4.05. What is the pH of a 0.149 M solution of HA?

Question 2. A weak base, B^- , has a pK_b of 5.52. A 25.00 mL aliquot of 0.137M B^- is titrated with 0.082M HCl. Calculate the pH after 18.05 mL of acid are added.

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Question 3. A weak base, B^- , has a pK_b of 5.52. A 25.00 mL aliquot of 0.100M B^- is titrated with 0.100M HCl. Calculate the pH after 25.00 mL of acid are added.

Question 4. A weak base, B^- , has a pK_b of 5.52. A 6.63 mL aliquot of 0.148M B^- is titrated with 0.154M HCl. Calculate the pH after 10.17 mL of acid are added.

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Question 5. 25.00 mL of an acetic acid solution are titrated to the endpoint with 36.00 mL of 0.1234 M KOH solution. The K_a for acetic acid is 1.8×10^{-5} . What is the molarity of the acetic acid solution?

Question 6. A solution of the compound, X_3Z , is found to have a Z^{3-} concentration of 0.00010M. Calculate K_{sp} for this compound. Calculate your answer to 3 significant figures.

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Question 7. 4.67 g of Ag_2SO_4 are placed in 1000 mL of distilled water in a 2 L beaker. After sufficient time is allowed, how many g of solid, undissolved silver sulfate remain in the beaker? K_{sp} for silver sulfate is 1.2×10^{-5} . Assume that the addition of the silver sulfate does not change the volume of the solution.

Question 8. Write a balanced net ionic reaction for each of the following dissolution reactions, and use the appropriate K_{sp} and K_{f} values in the appendix of your text to calculate the equilibrium constant for each.

a) Zinc hydroxide in aqueous NaOH to form $\text{Zn}(\text{OH})_4^{2-}$

b) Copper(II) hydroxide in aqueous ammonia to form $\text{Cu}(\text{NH}_3)_4^{2+}$

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c) Silver bromide in aqueous NH_3 to form $\text{Ag}(\text{NH}_3)_2^+$

Question 9. What is the molar solubility of barium hydroxide in water? Would you characterize barium hydroxide as being “soluble”? Why?

Question 10. What is the molar solubility of barium hydroxide in a pH 9.0 buffer?

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Question 11. For the vaporization of benzene, $\Delta H_{\text{vap}} = 30.7 \text{ kJ/mol}$, and $\Delta S_{\text{vap}} = 87.0 \text{ J/molK}$. Calculate ΔS_{total} at $70 \text{ }^\circ\text{C}$.

Question 12. Will benzene boil at $70 \text{ }^\circ\text{C}$? Why?

Question 13. Calculate, to three significant figures, the normal boiling point for benzene in $^\circ\text{C}$.

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Question 14. For melting of NaCl, $\Delta H_{\text{fusion}} = 30.2 \text{ kJ/mol}$, and $\Delta S_{\text{fusion}} = 28.1 \text{ J/molK}$. Calculate ΔS_{total} at 1050 K. Watch your significant figures.

Question 15. Does NaCl melt at 1050 K? Why?

Question 16. Calculate the melting point of NaCl. Report your answer in $^{\circ}\text{C}$. Report your answer to three significant figures.

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Question 17. What ratio of acid to conjugate base ($[HA]/[A^-]$) is required to prepare a buffer with a pH of 3.6 using lactic acid. K_a for lactic acid, $HC_3H_5O_3$ is 1.4×10^{-4} . Give two methods by which this buffer can be prepared.

Part II

Question 1. Discuss chemical spontaneity in terms of movement towards or away from equilibrium.

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Question 2. Discuss thermodynamic stability and kinetic stability.

Question 3. How can a reaction be endothermic, yet spontaneous? Offer an example with an explanation.

Question 4. Explain the statement, "A disordered state is more probable than an ordered state." Be sure to offer examples.

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Question 5. Explain the statement, “In any spontaneous process at constant temperature and pressure, the free energy of the system always decreases.”

Question 6. Discuss the notion that free energy is the maximum useful work that can be produced by a system.

Question 7. Discuss the relationship between free energy and the equilibrium constant.

Thermodynamics

- How the tendency of a process to occur by itself is distinct from how long it takes to occur
- The distinction between a spontaneous and a non-spontaneous change.
- Why the first law of thermodynamics and the sign of ΔH° cannot predict the direction of a spontaneous reaction
- The natural tendency of a system to become disordered over time and why a disordered system is more probable than an ordered one
- How disorder is expressed quantitatively by entropy (S)
- How the second law of thermodynamics expresses the criterion of spontaneity, that a change increases S_{univ}
- How absolute entropies (S°) can be obtained because the third law of thermodynamics provides a "zero point"
- How temperature, physical state, dissolution, molar mass, and molecular complexity influence S° values
- How $\Delta S^\circ_{\text{rxn}}$ is based on the difference between the S° values of the reactants and the products
- How the surroundings add heat to or remove heat from the system and how $\Delta S^\circ_{\text{surr}}$ influences the overall $\Delta S^\circ_{\text{rxn}}$.
- The relationship between $\Delta S^\circ_{\text{surr}}$ and $\Delta H^\circ_{\text{sys}}$
- How reactions proceed spontaneously toward equilibrium ($\Delta S^\circ_{\text{univ}} > 0$) and then proceed no further at equilibrium ($\Delta S^\circ_{\text{univ}} = 0$)
- How the free energy change combines the system's entropy and enthalpy changes
- How the expression for the free energy change is derived from the second law
- The relationship between ΔG and the maximum work a system can perform and why this quantity of work is never performed in a real process
- How temperature determines spontaneity for reactions in which ΔS° and ΔH° have the same sign
- Why the temperature at which a reaction becomes spontaneous occurs when $\Delta G^\circ = 0$
- How ΔG° is the free energy change from the start of a reaction (expressed by Q) to its end (expressed by K)
- The meaning of ΔG° and its relation to K
- The relation of ΔG to ΔG° and Q
- How ΔG decreases, no matter what the starting concentrations, as the reacting system moves spontaneously toward equilibrium

Required Skills

- Predicting relative S° values of substances and mixtures
- Calculating $\Delta S^\circ_{\text{rxn}}$ for a chemical or physical change
- Determining reaction spontaneity from $\Delta S^\circ_{\text{surr}}$ and $\Delta H^\circ_{\text{sys}}$
- Calculating $\Delta G^\circ_{\text{rxn}}$ from $\Delta H^\circ_{\text{f}}$ and S° values
- Calculating $\Delta G^\circ_{\text{rxn}}$ from $\Delta G^\circ_{\text{f}}$ values
- Calculating the effect of temperature on $\Delta G^\circ_{\text{rxn}}$
- Calculating the temperature at which a reaction becomes spontaneous
- Calculating K from ΔG°
- Using ΔG° and Q to calculate ΔG at any temperature

Equations

- $\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} > 0$
- $\Delta S^\circ_{\text{rxn}} = \sum m S^\circ_{\text{products}} - \sum m S^\circ_{\text{reactants}}$
- $\Delta S_{\text{surr}} - \frac{\Delta H}{T}$
- $\Delta G_{\text{sys}} = \Delta H_{\text{sys}} - T\Delta S_{\text{sys}}$
- $\Delta G^\circ_{\text{sys}} = \Delta H^\circ_{\text{sys}} - T\Delta S^\circ_{\text{sys}}$
- $\Delta G^\circ_{\text{rxn}} = \sum m G^\circ_{\text{products}} - \sum m G^\circ_{\text{reactants}}$
- $T = \frac{\Delta H^\circ}{\Delta S^\circ}$
- $\Delta G = RT \ln \frac{Q}{K} = RT \ln Q - RT \ln K$
- $\Delta G = \Delta G^\circ + RT \ln Q$