AP Chemistry – Units 4–6 Practice Set

AP Chemistry (College Board) · Total marks: 34 · Answers on the last page

This practice set covers Unit 4 (Chemical Reactions), Unit 5 (Kinetics) and Unit 6 (Thermodynamics). A calculator is permitted, as on the free-response section of the AP Chemistry exam.

Section A – Reactions and Stoichiometry (Unit 4)

1. Write the net ionic equation for the reaction between aqueous silver nitrate and aqueous sodium chloride. [2]

2. A 25.0 mL sample of 0.150 M HCl is titrated with 0.100 M NaOH. Calculate the volume of NaOH required to reach the equivalence point. [3]

3. Classify the reaction $\mathrm{Zn}(s) + 2\mathrm{HCl}(aq) \rightarrow \mathrm{ZnCl_2}(aq) + \mathrm{H_2}(g)$ and identify the oxidizing agent. [2]

Section B – Kinetics (Unit 5)

4. The rate law for a reaction is rate $= k[A]^2[B]$. If the concentration of A is tripled and the concentration of B is doubled, by what factor does the rate increase? [2]

5. For a first-order reaction, the rate constant is $k = 0.0231\ \mathrm{s^{-1}}$. Calculate the half-life of the reaction. [3]

6. Explain, in terms of collision theory, why increasing the temperature of a reaction increases its rate. [3]

7. A reaction has an activation energy of 75 kJ/mol. Using the Arrhenius equation qualitatively, explain the effect of adding a catalyst on the rate constant k. [2]

Section C – Thermodynamics (Unit 6) (FRQ-style, calculator required)

8. Calculate ΔH°_{rxn} for $\mathrm{CH_4}(g) + 2\mathrm{O_2}(g) \rightarrow \mathrm{CO_2}(g) + 2\mathrm{H_2O}(l)$, given ΔH°_f: $\mathrm{CH_4}(g) = -74.8$ kJ/mol, $\mathrm{CO_2}(g) = -393.5$ kJ/mol, $\mathrm{H_2O}(l) = -285.8$ kJ/mol, $\mathrm{O_2}(g) = 0$. [4]

9. A reaction has $\Delta H = -120$ kJ/mol and $\Delta S = -250$ J/(mol K). Calculate ΔG at 298 K and determine whether the reaction is spontaneous under these conditions. [4]

10. For the reaction in the previous question, determine the temperature above which the reaction becomes non-spontaneous. [3]

11. In a coffee-cup calorimeter, 50.0 g of water at 25.0°C is mixed with 50.0 g of water at 85.0°C. Assuming no heat loss to the surroundings, calculate the final temperature of the mixture. [3]

12. Justify, using entropy considerations, why the dissolution of $\mathrm{NaCl}(s)$ in water results in an increase in the entropy of the system, despite the ordering effect of hydration shells forming around the ions. [3]

Answers

Check your working against these. If an answer does not match, try the question again before looking for the method.

1. $\mathrm{Ag^+}(aq) + \mathrm{Cl^-}(aq) \rightarrow \mathrm{AgCl}(s)$

2. $n(\mathrm{HCl}) = 0.0250\,\mathrm{L} \times 0.150\,\mathrm{M} = 3.75 \times 10^{-3}\,\mathrm{mol}$. Since the mole ratio is 1:1, $n(\mathrm{NaOH}) = 3.75 \times 10^{-3}\,\mathrm{mol}$, so $V = \dfrac{3.75 \times 10^{-3}}{0.100} = 0.0375\,\mathrm{L} = 37.5\,\mathrm{mL}$.

3. This is a single-replacement (oxidation-reduction) reaction. HCl is the oxidizing agent, since $\mathrm{H^+}$ is reduced to $\mathrm{H_2}$ (Zn is oxidized and is the reducing agent).

4. New rate $= k(3[A])^2(2[B]) = k \times 9[A]^2 \times 2[B] = 18 \times k[A]^2[B]$. The rate increases by a factor of 18.

5. For a first-order reaction, $t_{1/2} = \dfrac{\ln 2}{k} = \dfrac{0.693}{0.0231} = 30.0\,\mathrm{s}$.

6. At higher temperature, particles have greater average kinetic energy and move faster, so they collide more frequently. More importantly, a greater proportion of collisions have energy equal to or greater than the activation energy, E_a, so a greater fraction of collisions are successful, increasing the rate.

7. A catalyst provides an alternative reaction pathway with a lower activation energy, E_a. Since $k = Ae^{-E_a/RT}$, a smaller E_a gives a larger value of k (the exponent is less negative), so the rate constant, and hence the rate, increases without the catalyst itself being consumed.

8. $\Delta H^\circ_{rxn} = \sum \Delta H^\circ_f(\text{products}) - \sum \Delta H^\circ_f(\text{reactants})$
$= [-393.5 + 2(-285.8)] - [-74.8 + 2(0)]$
$= -965.1 - (-74.8) = -890.3\,\mathrm{kJ/mol}$.

9. $\Delta G = \Delta H - T\Delta S = -120\,000\,\mathrm{J/mol} - (298\,\mathrm{K})(-250\,\mathrm{J/(mol\,K)}) = -120\,000 + 74\,500 = -45\,500\,\mathrm{J/mol} = -45.5\,\mathrm{kJ/mol}$. Since $\Delta G < 0$, the reaction is spontaneous at 298 K.

10. $\Delta G = 0$ when $T = \dfrac{\Delta H}{\Delta S} = \dfrac{-120\,000}{-250} = 480\,\mathrm{K}$. Since $\Delta S < 0$, ΔG becomes positive (non-spontaneous) above $T = 480\,\mathrm{K}$.

11. Since the two samples have equal mass and specific heat capacity, heat lost by the hot water equals heat gained by the cold water: $mc(85.0 - T) = mc(T - 25.0) \Rightarrow 85.0 - T = T - 25.0 \Rightarrow T = \dfrac{85.0 + 25.0}{2} = 55.0^\circ\mathrm{C}$.

12. Dissolving breaks up the highly ordered ionic lattice into freely moving, dispersed $\mathrm{Na^+}$ and $\mathrm{Cl^-}$ ions in solution, greatly increasing the number of accessible microstates (positional disorder). Although water molecules become somewhat more ordered as hydration shells form around the ions (a local decrease in entropy), this decrease is smaller than the large increase from disrupting the crystal lattice, so the overall entropy of the system increases.

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