AP Chemistry Question Bank

386 Unique Problems with Detailed Solutions

Section 1 - Practice Problems

Unit 1 - Composition, Atomic Structure, Periodic Trends, and Analytical Measurements

Problem 1.1 - Mass percent composition

A 4.5 g sample of what compound has the highest mass percent of oxygen?

A. Na_2O (molar mass = 62 g/mol)

B. Li_2O (molar mass = 30 g/mol)

C. MgO (molar mass = 40 g/mol)

D. SrO (molar mass = 104 g/mol)

Problem 1.2 - Empirical formula

A compound includes 14 g nitrogen and 32 g oxygen. Find the empirical formula.

A. NO

B. N_2O

C. NO_2

D. NO_3

E. N_2O_6

Problem 1.3 - Purity and percent composition

Pure glucose, $C_6H_12O_6$, is 53.3% oxygen by mass. A chemist analyzes an impure glucose sample and determines 49.7% oxygen. Which impurity could explain the lower oxygen percent?

a. n-eicosane, C_20H_{42}

b. ribose, $C_5H_10O_5$

c. fructose, $C_6H_{12}O_6$

d. sucrose, $C_{12}H_{22}O_11$

Problem 3.7 - Energy required to break all bonds

With the bond-energy table, find the energy required to break all the bonds in one mole of each substance.

a) NH_3

b) CH_3OH

Problem 3.8 - Specific heat sample: heat released by a reaction

A reaction is performed in 200 mL of solution. If the solution temperature increased from 15°C to 20°C, how much heat was released by the reaction?

Problem 3.9 - Specific heat sample: identifying a substance

Find the specific heat capacity of a substance and determine it if 43.2 kJ of heat added to 4000 g of the substance raises its temperature by 45°C.

Problem 3.10 - Specific heat sample: warming ice

What is the temperature change and final temperature if 156 J of heat is introduced to 25 g of ice at -5°C?

Problem 3.11 - Specific heat comparison

Use the specific-heat table to respond to the questions.

a) Which substance heats up most rapidly?

b) Which substance cools most rapidly?

Problem 3.12 - Heating water in an electric kettle

When 600 mL of water in an electric kettle is heated from 20°C to 85°C, what quantity of heat flows into the water?

Problem 3.53 - Bond enthalpy of F-F in NF3 formation

Thermochemical equation from source

$$N_2(g) + 3\,F_2(g) \rightarrow 2\,NF_3(g) \qquad \Delta H^\circ_{298} = -264\ \text{kJ mol}^{-1}$$

The questions relate to the synthesis reaction displayed.

a) The enthalpy change in a chemical reaction is the difference between energy absorbed in breaking bonds in the reactants and energy released by bond formation in the products. What number of bonds are formed when two molecules of NF_3 are formed?

Average bond enthalpy data from source

Bond	Average bond enthalpy (kJ/mol)
N≡N	946
N-F	272
F-F	?

b) Apply the thermochemical equation and the table of average bond enthalpies to find the average enthalpy of the F-F bond.

Problem 3.54 - Formation equations

Provide the thermochemical equation for the formation of each species.

a) $H_2O_{(l)}$

b) $C_3H_{8(g)}$

c) $CO_{2(g)}$

Problem 3.55 - Hess's law with sulfur oxides

Given:

$S_{(s)} + 3/2\ O_{2(g)} \rightarrow SO_{3(g)}$, $\Delta H = -395.2$ kJ

$2\ SO_{2(g)} + O_{2(g)} \rightarrow 2\ SO_{3(g)}$, $\Delta H = -198.2$ kJ

Calculate $\Delta H°$ for the reaction: $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$.

Problem 4.41 - Effect of changing concentration of A

A + B → products; rate = k[B]

How will a change in the concentration of A affect the rate of the reaction? No explanation necessary.

Problem 4.42 - Initial rate law for NO2 and O3

$NO_{2(g)} + O_{3(g)} \rightarrow NO_{3(g)} + O_{2(g)}$

The reaction was studied and the following data were collected.

a) Write the rate law expression for the given reaction.

b) Predict the initial rate of the reaction, with appropriate units, when $[NO_2]0 = [O_3]0 = 4.5 \times 1^{0-}5$ mol/L.

Exp.	$[NO_2]0$ (mol/L)	$[O_3]0$ (mol/L)	Initial rate (mol/(L·s))
1	$5.0 \times 1^{0-}5$	$1.0 \times 1^{0-}5$	0.022
2	$5.0 \times 1^{0-}5$	$2.0 \times 1^{0-}5$	0.044
3	$2.5 \times 1^{0-}5$	$2.0 \times 1^{0-}5$	0.022

Problem 4.43 - Low-temperature mechanism for NO2 and CO

$NO_{2(g)} + CO_{(g)} \rightarrow CO_{2(g)} + NO$

At low temperature, the reaction appears to follow this mechanism:

(1) $NO_2 + NO_2 \rightarrow NO_3 + NO$ slow

(2) $NO_3 + CO \rightarrow NO_2 + CO_2$ fast

a) State the reaction intermediate$_{(s)}$. If there are no intermediates, indicate so.

b) Account for why this is a self-catalyzed reaction.

c) Predict the rate law expression based on the mechanism.

d) Which step in the mechanism do you think will have the largest activation energy, Ea? No explanation necessary.

Problem 5.55 - SO2 and NO2 equilibrium at 733 K

At 733 K, sulfur dioxide and nitrogen dioxide establish equilibrium with nitrogen monoxide and sulfur trioxide:

$SO_{2(g)} + NO_{2(g)} \rightleftharpoons NO_{(g)} + SO_{3(g)}$; $K_{eq} = 85.00$

If SO_2 and NO_2 are initially each 0.100 mol/L, find the equilibrium concentrations.

Problem 5.56 - Ethene equilibrium with product initially present

$C_2H_{2(g)} + H_{2(g)} \rightleftharpoons C_2H_{4(g)}$; $K_{eq} = 4.00$ L/mol. A 2.00 L container includes 6.00 mol ethyne, 2.00 mol hydrogen, and 10.0 mol ethene. At equilibrium, what is the concentration of ethene?

Problem 5.57 - HF equilibrium from all gases

$H_{2(g)} + F_{2(g)} \rightleftharpoons 2\ HF_{(g)}$; $K_{eq} = 3.00$. 5.00 mol of each gas is placed in a 1.00 L flask. Find the equilibrium concentrations.

Problem 5.58 - Ozone and NO equilibrium

$O_{3(g)} + NO_{(g)} \rightleftharpoons O_{2(g)} + NO_{2(g)}$; $K_{eq} = 25.0$. If 1.25 mol of each gas is placed in a 2.00 L flask, find the equilibrium concentrations.

Problem 5.59 - HI decomposition equilibrium

At 430 °C, K = 1.84 × 102 for $2\ HI_{(g)} \rightleftharpoons H_{2(g)} + I_{2(g)}$. If 0.100 mol HI is placed in a 1.00 L container and allowed to reach equilibrium, find the equilibrium concentrations of all species.

Problem 5.60 - I2 dissociation equilibrium

At 727 °C, K = 3.80 × 105 mol/L for $I_{2(g)} \rightleftharpoons 2\ I_{(g)}$. If the initial concentration of molecular iodine is 0.200 mol/L, find the concentration of atomic iodine at equilibrium.

Problem 5.61 - NO2 dimerization equilibrium

$2\ NO_{2(g)} \rightleftharpoons N_2O_{4(g)}$. At a certain temperature, K = 0.000212 L/mol. If 0.155 mol NO_2 is introduced to a 1.00 L container, find the equilibrium concentration of N_2O_4.

Problem 7.6 - Balancing in basic solution

Complete and balance the following equations for oxidation-reduction reactions that take place in basic solution.

a) $CN^-_{(aq)} + MnO^{4-}_{(aq)} \rightarrow CNO^-_{(aq)} + MnO_{2(s)}$

b) $NO^{2-}_{(aq)} + Al_{(s)} \rightarrow NH_{3(aq)} + Al(OH)^{4-}_{(aq)}$

c) $Cr(OH)_{3(s)} + ClO^-_{(aq)} \rightarrow CrO_4^{2-}{}_{(aq)} + Cl_{2(g)}$

Problem 7.7 - Voltaic cell from dichromate and iodide

The following oxidation-reduction reaction is spontaneous:

$Cr_2O_7^{2-}{}_{(aq)} + 14\ H^+_{(aq)} + 6\ I^-_{(aq)} \rightarrow 2\ Cr^{3+}_{(aq)} + 3\ I_{2(s)} + 7\ H_2O_{(l)}$

A solution containing $K_2Cr_2O_7$ and H_2SO_4 is poured into one beaker, and a solution of KI is poured into another. A salt bridge joins the beakers. A non-reactive metallic conductor such as platinum foil is suspended in each solution, and the two conductors are connected through a voltmeter.

Indicate the reaction at the anode, the reaction at the cathode, the direction of electron and ion migration, and the signs of the electrodes.

Problem 7.8 - Anode and cathode from given half-reactions

The two half-reactions in a voltaic cell are:

$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2\ e^-$

$ClO^{3-}_{(aq)} + 6\ H^+_{(aq)} + 6\ e^- \rightarrow Cl^-_{(aq)} + 3\ H_2O_{(l)}$

a) Indicate which reaction takes place at the anode and which takes place at the cathode.

b) Which electrode is consumed in the cell reaction?

c) Which electrode is positive?

Problem 2.10 Solution

Substitutional alloys become harder when atoms of different sizes disrupt the regular layers of the metallic crystal. Copper atoms are appreciably smaller than silver atoms, so their presence makes it more difficult for layers of atoms to slide past one another. Gold and silver have more similar metallic radii, so an Ag-Au lattice is less distorted. The observation is therefore best explained by choice c.

Final answer: c. Copper is smaller than silver and distorts the lattice

Problem 2.11 Solution

Germanium has four valence electrons. A P-type semiconductor is produced by doping it with an element that has one fewer valence electron, creating electron vacancies called holes. Gallium is a Group 13 element with three valence electrons, so it acts as an acceptor impurity in Ge. Phosphorus and selenium would supply extra electrons and produce N-type behavior. Choice d is correct.

Final answer: d. gallium

Problem 2.12 Solution

With similar chain spacing, the polymer containing heavier atoms has the greater mass per unit volume. Polypropylene contains only carbon and hydrogen, whereas polyvinyl chloride contains a chlorine atom in each repeating unit. Chlorine adds substantial mass without requiring a proportionally larger volume. PVC therefore has the higher density and sinks in water, while lower-density PP floats.

Final answer: PVC sinks; PP floats

Problem 2.13 Solution

Boiling separates intact H_2O molecules from one another. The intramolecular O-H covalent bonds remain unbroken; otherwise the substance would chemically decompose. The principal intermolecular attractions that must be overcome are hydrogen bonds between water molecules. Therefore, choice b identifies the interaction disrupted during boiling.

Final answer: b. hydrogen bonds

Problem 3.67 Solution

a) $\Delta H^\circ = \Delta H^\circ_f[H_2O_{(g)}] - \Delta H^\circ_f[H_2O_{(l)}] = -241.8 - (-285.8) = +44.0$ kJ/mol.

b) Convert ΔS° to 0.1188 kJ/(K·mol).

$\Delta G^\circ = 44.0 - (298)(0.1188) = +8.60$ kJ/mol

c) The positive standard free-energy change means vaporization is not spontaneous when both phases are in their standard states at 25 °C.

d) At the same temperature, reduce the partial pressure of water vapour below its equilibrium vapour pressure by using dry air, applying a vacuum, or continuously removing vapour. This changes the nonstandard free energy and favours evaporation.

Final answer: a) +44.0 kJ/mol; b) +8.60 kJ/mol; c) not spontaneous because $\Delta G^\circ > 0$; d) lower the water-vapour partial pressure

Problem 3.68 Solution

The water and calorimeter both absorb heat during the 6.0 °C rise.

$q_{water} = (100\text{ g})(4.184\text{ J/(g·°C)})(6.0\text{ °C}) = 2510$ J

$q_{cal} = (45.7\text{ J/°C})(6.0\text{ °C}) = 274$ J

$q_{reaction} = -(2510 + 274)\text{ J} = -2.784$ kJ

n(KOH) = (3.25 g)/(56.11 g/mol) = 0.0579 mol

$\Delta H = (-2.784\text{ kJ})/(0.0579\text{ mol}) = -48.1$ kJ/mol

The measured concentrations, calorimeter conditions, and solution activities are not all standard-state values, so the result is written ΔH rather than ΔH°.

Final answer: $\Delta H \approx -48.1$ kJ/mol KOH; it is not ΔH° because the experimental conditions are nonstandard

Problem 3.69 Solution

Without a lid, some energy from the hot metal escapes to the air instead of heating the water. The measured water temperature rise is therefore smaller than it would be in an ideal insulated calorimeter. When the student assumes all the metal's lost heat entered the water, the calculated q for the metal is too small. The calculated specific heat capacity is consequently underestimated. Choice b is correct.

Final answer: b. The determined specific heat capacity is too small because heat is lost to the surroundings

Problem 7.28 Solution

A correct redox balance must conserve both matter and electrical charge. The half-reaction method supplies coefficients that satisfy these two conditions simultaneously and prevents an apparently atom-balanced equation from carrying unequal charge.

For basic solution, first obtain an atom- and charge-balanced half-reaction and then neutralize any H^+ with OH^-. Cancel the resulting water molecules so that the final equation contains the smallest necessary amounts of H_2O and OH^-.

$Cr_2O_7^{2-} + 6Cl^- + 7H_2O \rightarrow 2Cr^{3+} + 3Cl_2 + 14OH^-$

Chromium is reduced from +6 to +3, and chloride is oxidized to elemental chlorine.

Final answer: $Cr_2O_7^{2-} + 6Cl^- + 7H_2O \rightarrow 2Cr^{3+} + 3Cl_2 + 14OH^-$

Problem 7.29 Solution

The oxidation half-reaction releases the same number of electrons that the reduction half-reaction consumes. Adding the adjusted half-reactions and cancelling common species produces the lowest whole-number coefficients shown below.

For basic solution, first obtain an atom- and charge-balanced half-reaction and then neutralize any H^+ with OH^-. Cancel the resulting water molecules so that the final equation contains the smallest necessary amounts of H_2O and OH^-.

$2CrO_4^{2-} + 3SO_3^{2-} + 5H_2O \rightarrow 2Cr^{3+} + 3SO_4^{2-} + 10OH^-$

Cr(+6) gains electrons to become Cr(+3), whereas sulfur rises from +4 in sulfite to +6 in sulfate.

Final answer: $2CrO_4^{2-} + 3SO_3^{2-} + 5H_2O \rightarrow 2Cr^{3+} + 3SO_4^{2-} + 10OH^-$

Problem 7.30 Solution

Rather than treating oxygen and hydrogen first, balance the electron-transfer centres. The remaining O and H atoms are supplied by water together with H^+ or OH^- as required, and any species appearing on both sides is cancelled.

In acidic solution, H_2O is used to balance oxygen and H^+ is used to balance hydrogen. The charge is then balanced with electrons before the half-reactions are added.

$2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

The ten electrons released when five oxalate ions form carbon dioxide are consumed by two permanganate ions.

Final answer: $2MnO_4^- + 5C_2O_4^{2-} + 16H^+ \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O$

Problem 7.102 Solution

Begin with a clean, dry crucible and record its mass. Add a known mass of pure copper, then heat the uncovered or loosely covered crucible in an excess supply of oxygen. Cool the crucible in a desiccator, weigh it, and repeat the heating-cooling-weighing cycle until a constant mass is obtained.

m(O) = m(final oxide) - m(initial Cu)

n(Cu) = m(Cu)/(63.55 g/mol)

n(O) = m(O)/(16.00 g/mol)

An O:Cu ratio of approximately 1:1 identifies CuO. An O:Cu ratio of approximately 1:2, or 0.50 mol O per mole Cu, identifies Cu_2O. Heating to constant mass and cooling in a desiccator reduce error from incomplete oxidation and moisture uptake.

Final answer: Use the measured O:Cu mole ratio: 1:1 indicates CuO, while 1:2 indicates Cu_2O

Problem 7.103 Solution

Place a known mass of $CaCO_3$ in a rigid, heat-resistant sealed vessel and determine the total mass of the vessel and contents. Heat the vessel so that the reaction occurs, allow it to return to room temperature, and weigh the entire closed system again.

$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

Although the solid loses mass as CO_2 forms, the gas remains inside the sealed vessel. Therefore the combined mass of CaO, CO_2, and the container must equal the initial mass of $CaCO_3$ and the container. As an additional quantitative check, the CO_2 amount can be obtained from its pressure, volume, and temperature and compared with the one-to-one stoichiometric ratio.

Final answer: The total mass of the sealed system before and after heating should be the same

Problem 7.104 Solution

The last two measurements agree, so 25.252 g is taken as the constant mass of the crucible and oxide.

m(oxide) = 25.252 - 18.650 = 6.602 g

m(O) = 6.602 - 5.200 = 1.402 g

n(Sn) = (5.200 g)/(118.71 g/mol) = 0.04380 mol

n(O) = (1.402 g)/(16.00 g/mol) = 0.08763 mol

n(O)/n(Sn) = 0.08763/0.04380 = 2.00

The empirical ratio is one Sn atom to two O atoms, so the oxide is SnO_2 rather than SnO.

Final answer: The product is SnO_2

Problem 8.2 Solution

Before equivalence, the pH is controlled by the weak base and its conjugate acid. At equivalence, the conjugate acid formed from the base controls the pH. Only after the equivalence point is there excess added HCl, and its remaining hydronium concentration dominates the solution pH. On the standard labeled curve, this far-right post-equivalence region is point E.

Final answer: E, the post-equivalence region

Problem 8.3 Solution

First compare reactant amounts for the balanced reaction.

$2Fe(OH)_{3(s)} + 3H_2SO_{4(aq)} \rightarrow Fe_2(SO_4)_{3(aq)} + 6H_2O_{(l)}$

$n[Fe(OH)_3] = (45.0\ g)/(106.87\ g/mol) = 0.421\ mol$

$n(H_2SO_4) = (0.0450\ L)(2.00\ mol/L) = 0.0900\ mol$

The acid can react with only $(2/3)(0.0900) = 0.0600$ mol $Fe(OH)_3$, corresponding to

$m = (0.0600\ mol)(106.87\ g/mol) = 6.41\ g\ Fe(OH)_3$

Thus H_2SO_4 is limiting and 38.6 g $Fe(OH)_3$ remains. To determine a numerical $\Delta H°_{rxn}$, the temperature change and the heat capacities of all materials absorbing heat are required. Adding an unreactive object increases the total heat capacity, so the observed $|\Delta T|$ becomes smaller. If its heat capacity is ignored, the calculated magnitude of $\Delta H°_{rxn}$ is too small.

Hermione uses the same amount of $Fe(OH)_3$ but ten times as much acid. $Fe(OH)_3$ then becomes limiting, so more reaction can occur; however, the much larger solution mass also absorbs the released heat. Using the stated quantities and similar solution heat capacities, the increased thermal mass dominates and her temperature change is smaller than Ron's.

Replacing $Fe(OH)_3$ with the same mass of $Al(OH)_3$ supplies more moles because $Al(OH)_3$ has the lower molar mass. Under the stated assumption that the molar reaction enthalpy is unchanged and acid remains sufficient, more reaction occurs and the temperature change is greater.

The prompt does not provide the absolute entropy values needed for part (c). Once those values are supplied, calculate

$\Delta S°_{rxn} = \Sigma nS°(products) - \Sigma nS°(reactants)$

Final answer: a) H_2SO_4 is limiting; 6.41 g $Fe(OH)_3$ reacts, and additional temperature and heat-capacity data are needed for numerical $\Delta H°_{rxn}$. Ignoring the unreactive object makes $|\Delta H°_{rxn}|$ too small. b) Hermione has a smaller ΔT; replacing $Fe(OH)_3$ with $Al(OH)_3$ gives a greater ΔT under the stated assumptions. c) The numerical $\Delta S°_{rxn}$ cannot be determined without the missing absolute entropy values.

3.3	a exothermic; b exothermic; c exothermic; d endothermic; e endothermic
3.4	At constant pressure, $\Delta H_{system} = -q_{surroundings}$ for an isolated calorimetry setup
3.5	$\Delta H \approx -2.05 \times 10^3$ kJ/mol C_3H_8
3.6	a) -486 kJ; b) -578 kJ
3.7	a) 1167 kJ/mol; b) 2063 kJ/mol
3.8	4.18 kJ released
3.9	C = 0.240 J/(g·°C), consistent with silver
3.10	ΔT = 3.10 °C; final temperature = -1.90 °C
3.11	Iron heats and cools most rapidly
3.12	1.63×10^2 kJ
3.13	ΔT = 8.57 °C; final temperature = 28.6 °C
3.14	Aluminum
3.15	$\Delta T \approx 18$ °C
3.16	a) 14.6 MJ; b) about $77.0 per year
3.17	Thermal energy depends on both temperature and the amount of matter
3.18	$\Delta H_{solution} \approx +17$ kJ/mol KCl
3.19	$\Delta H_{solution} \approx +14$ kJ/mol urea
3.20	$\Delta H_{reaction} \approx -26$ kJ/mol $Ba(NO_3)_2$
3.21	1.54 g $C_{10}H_{22}$
3.22	$\Delta H \approx -242$ kJ for 1.00 kg
3.23	a) heat is a product; b-i) -5605.4 kJ; ii) -1401.35 kJ; iii) +2802.7 kJ; iv) +467.1 kJ
3.24	$NH_4NO_{3(s)} + 25.8$ kJ $\rightarrow NH_4^+{}_{(aq)} + NO_3^-{}_{(aq)}$; endothermic
3.25	$\Delta H_{combustion}$ = -2657.4 kJ/mol C_4H_{10}
3.26	$HCl_{(g)} \rightarrow H^+{}_{(aq)} + Cl^-{}_{(aq)}$, ΔH = -75 kJ/mol; or $HCl_{(g)} \rightarrow H^+{}_{(aq)} + Cl^-{}_{(aq)} + 75$ kJ
3.27	a) $H_2O_{(l)} \rightarrow H_2O_{(g)}$, ΔH = +40.65 kJ/mol; b) $H_2O_{(g)} \rightarrow H_2O_{(l)}$, ΔH = -40.65 kJ/mol
3.28	a) $Au_{(s)} \rightarrow Au_{(l)}$, ΔH = +12.55 kJ/mol; b) $Au_{(l)} \rightarrow Au_{(s)}$, ΔH = -12.55 kJ/mol
3.29	a) -241.8 kJ/mol H_2; b) -28.35 kJ/mol NH_3; c) +81.6 kJ/mol N_2
3.30	-114 kJ/mol H_2SO_4; -57 kJ/mol NaOH
3.31	a and b endothermic with products above reactants; c exothermic with products below reactants
3.32	$\Delta H°$ = -851.5 kJ
3.33	$\Delta H°$ = +131.3 kJ
3.34	$\Delta H°$ = -524.8 kJ
3.35	$\Delta H°$ = -271 kJ/mol; about -2.08×10^3 kJ for 200 g C_2H_2
3.36	$\Delta H°_f[C_8H_{18(l)}] \approx -250$ kJ/mol
3.37	a) it forms two moles of HCl; b) $Br_{2(g)}$ is not bromine's standard state
3.38	d. $Cu_{(s)} + 1/2O_{2(g)} \rightarrow CuO_{(s)}$
3.39	a) $6C_{(s)} + 3H_{2(g)} \rightarrow C_6H_{6(l)}$; b) $K_{(s)} + 1/2Br_{2(l)} + 3/2O_{2(g)} \rightarrow KBrO_{3(s)}$; c) $6C_{(s)} + 6H_{2(g)} + 3O_{2(g)} \rightarrow C_6H_{12}O_{6(s)}$; d) $Mg_{(s)} + H_{2(g)} + O_{2(g)} \rightarrow Mg(OH)_{2(s)}$
3.40	a) approximately -3.51×10^3 kJ/mol; b) approximately -24.8 kJ