CollegeBoard
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Program

AP® Chemistry
2004 Scoring Guidelines

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AP® CHEMISTRY
2004 SCORING GUIDELINES

Question 1

Answer the following questions relating to the solubilities of two silver compounds, Ag_2CrO_4 and Ag_3PO_4.

Silver chromate dissociates in water according to the equation shown below.

$$Ag_2CrO_4(s) \rightleftarrows 2\,Ag^+(aq) + CrO_4^{2-}(aq) \qquad K_{sp} = 2.6 \times 10^{-12} \text{ at } 25°C$$

(a) Write the equilibrium-constant expression for the dissolving of $Ag_2CrO_4(s)$.

$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$	1 point for correct expression

(b) Calculate the concentration, in mol L^{-1}, of $Ag^+(aq)$ in a saturated solution of Ag_2CrO_4 at 25°C.

$Ag_2CrO_4(s) \rightleftarrows 2\,Ag^+(aq) + CrO_4^{2-}(aq)$ I – 0 0 C – $+2x$ $+x$ F – $0 + 2x$ $0 + x$	1 point for correct stoichiometry of $Ag^+(aq)$ and $CrO_4^{2-}(aq)$
$2.6 \times 10^{-12} = [Ag^+]^2[CrO_4^{2-}]$ $2.6 \times 10^{-12} = [2x]^2[x] = 4x^3$ $8.7 \times 10^{-5} = x = [CrO_4^{2-}]$ $[Ag^+] = 2x = 2 \times (8.7 \times 10^{-5}\,M) = 1.7 \times 10^{-4}\,M$	1 point for substituting and calculating $[Ag^+]$

(c) Calculate the maximum mass, in grams, of Ag_2CrO_4 that can dissolve in 100. mL of water at 25°C.

$\dfrac{8.7 \times 10^{-5} \text{ mole } Ag_2CrO_4}{1\text{ L}} \times \dfrac{331.7\text{ g}}{1 \text{ mole } Ag_2CrO_4} = \dfrac{0.029\text{ g } Ag_2CrO_4}{1\text{ L}}$	1 point for molar mass of Ag_2CrO_4
$\dfrac{0.029\text{ g } Ag_2CrO_4}{1\text{ L}} \times 0.100\text{ L} = 0.0029\text{ g } Ag_2CrO_4$	1 point for mass of Ag_2CrO_4 in 100 mL

Question 1 (cont'd.)

(d) A 0.100 mol sample of solid $AgNO_3$ is added to a 1.00 L saturated solution of Ag_2CrO_4. Assuming no volume change, does $[CrO_4^{2-}]$ increase, decrease, or remain the same? Justify your answer.

The $[CrO_4^{2-}]$ will decrease. Adding $[Ag^+]$ will make Q (nonequilibrium reaction quotient) greater than K. To re-establish equilibrium, the reaction goes from right to left, decreasing the quotient to return to equilibrium.	1 point for correct prediction and explanation in terms of Q or LeChâtelier's principle

In a saturated solution of Ag_3PO_4 at 25°C, the concentration of $Ag^+(aq)$ is 5.3×10^{-5} M. The equilibrium-constant expression for the dissolving of $Ag_3PO_4(s)$ in water is shown below.

$$K_{sp} = [Ag^+]^3[PO_4^{3-}]$$

(e) Write the balanced equation for the dissolving of Ag_3PO_4 in water.

$Ag_3PO_4(s) \rightarrow 3\ Ag^+(aq) + PO_4^{3-}(aq)$	1 point for correct, balanced chemical equation

(f) Calculate the value of K_{sp} for Ag_3PO_4 at 25°C.

$[Ag^+] = 5.3 \times 10^{-5}\ M$ $[PO_4^{3-}] = 5.3 \times 10^{-5}\ M\ Ag^+ \times \frac{1\ \text{mol}\ PO_4^{3-}}{3\ \text{mol}\ Ag^+} = 1.8 \times 10^{-5}\ M$	1 point for correct $[PO_4^{3-}]$
$K_{sp} = [Ag^+]^3[PO_4^{3-}]$ $= (5.3 \times 10^{-5})^3(1.8 \times 10^{-5}) = 2.6 \times 10^{-18}$	1 point for K_{sp}

(g) A 1.00 L sample of saturated Ag_3PO_4 solution is allowed to evaporate at 25°C to a final volume of 500. mL. What is $[Ag^+]$ in the solution? Justify your answer.

$[Ag^+] = 5.3 \times 10^{-5}\ M$ The $[Ag^+]$ in a saturated solution of Ag_3PO_4 is independent of the volume of the solution.	1 point for correct $[Ag^+]$ and explanation

Question 2

$$2\ \mathrm{Fe}(s) + \frac{3}{2}\ \mathrm{O_2}(g) \rightarrow \mathrm{Fe_2O_3}(s) \qquad \Delta H_f^\circ = -824\ \mathrm{kJ\ mol^{-1}}$$

Iron reacts with oxygen to produce iron(III) oxide, as represented by the equation above. A 75.0 g sample of Fe(*s*) is mixed with 11.5 L of $O_2(g)$ at 2.66 atm and 298 K.

(a) Calculate the number of moles of each of the following before the reaction begins.

(i) Fe(*s*)

$n_{\mathrm{Fe}} = 75.0\ \mathrm{g\ Fe} \times \left(\frac{1\ \mathrm{mol\ Fe}}{55.85\ \mathrm{g\ Fe}}\right) = 1.34\ \mathrm{mol\ Fe}$	1 point for number of moles of Fe(*s*)

(ii) $O_2(g)$

$PV = nRT$ $n_{\mathrm{O_2}} = \frac{PV}{RT} = \frac{2.66\ \mathrm{atm} \times 11.5\ \mathrm{L}}{0.0821 \frac{\mathrm{L\ atm}}{\mathrm{mol\ K}} \times 298\ \mathrm{K}}$ $n_{\mathrm{O_2}} = 1.25\ \mathrm{mol\ O_2}$	1 point for number of moles of $O_2(g)$

(b) Identify the limiting reactant when the mixture is heated to produce $Fe_2O_3(s)$. Support your answer with calculations.

$n_{\mathrm{O_2}}$ reacting $= 1.34\ \mathrm{mol\ Fe} \times \left(\frac{1.5\ \mathrm{mol\ O_2}}{2\ \mathrm{mol\ Fe}}\right)$ $= 1.01\ \mathrm{mol\ O_2}$ There is 1.25 mol O_2 initially, so there is an excess of O_2 , and Fe is the limiting reactant. *OR* n_{Fe} reacting $= 1.25\ \mathrm{mol\ O_2} \times \left(\frac{2\ \mathrm{mol\ Fe}}{1.5\ \mathrm{mol\ O_2}}\right)$ $= 1.67\ \mathrm{mol\ Fe}$ There is 1.34 mol Fe initially, so there is not enough Fe, and Fe is the limiting reactant.	1 point for identifying limiting reactant 1 point for supporting calculation

Question 2 (cont'd.)

(c) Calculate the number of moles of $Fe_2O_3(s)$ produced when the reaction proceeds to completion.

$n_{Fe_2O_3} = 1.34 \text{ mol Fe} \times \left(\frac{1 \text{ mol } Fe_2O_3}{2 \text{ mol Fe}}\right) = 0.670 \text{ mol } Fe_2O_3$	1 point for number of moles of Fe_2O_3 produced

(d) The standard free energy of formation, ΔG_f°, of $Fe_2O_3(s)$ is $-740.$ kJ mol^{-1} at 298 K.

(i) Calculate the standard entropy of formation, ΔS_f°, of $Fe_2O_3(s)$ at 298 K. Include units with your answer.

$\Delta G_f^\circ = \Delta H_f^\circ - T\,\Delta S_f^\circ$ $-740. \text{ kJ mol}^{-1} = -824 \text{ kJ mol}^{-1} - (298 \text{ K})\ \Delta S_f^\circ$ $+84 \text{ kJ mol}^{-1} = -(298 \text{ K})\ \Delta S_f^\circ$ $\Delta S_f^\circ = \frac{+84 \text{ kJ mol}^{-1}}{-298 \text{ K}} = -0.28 \text{ kJ mol}^{-1} \text{ K}^{-1}$	1 point for calculation of ΔS_f° 1 point for correct units

(ii) Which is more responsible for the spontaneity of the formation reaction at 298 K, the standard enthalpy of formation, ΔH_f°, or the standard entropy of formation, ΔS_f°? Justify your answer.

ΔH_f° is the more important factor. The reaction is exothermic, which favors spontaneity. ΔS_f° is negative, which means the system becomes more ordered as the reaction proceeds. Greater order will not increase the spontaneity of the reaction.	1 point for indicating that ΔH_f° is responsible and for an explanation that addresses the signs of ΔH° and ΔS°

Question 2 (cont'd.)

The reaction represented below also produces iron(III) oxide. The value of ΔH° for the reaction is $-280.$ kJ per mole of $Fe_2O_3(s)$ formed.

$$2\,FeO(s) + \frac{1}{2}\,O_2(g) \rightarrow Fe_2O_3(s)$$

(e) Calculate the standard enthalpy of formation, ΔH_f°, of $FeO(s)$.

$\Delta H^\circ_{rxn} = \Sigma\Delta H_f^\circ \text{ (products)} - \Sigma\Delta H_f^\circ \text{ (reactants)}$ $\Delta H^\circ_{rxn} = \Delta H_f^\circ\ Fe_2O_3(s) - [2\,\Delta H_f^\circ \text{ of } FeO(s) + \frac{1}{2}\Delta H_f^\circ\ O_2(g)]$ $-280.\text{ kJ mol}^{-1} = -824\text{ kJ mol}^{-1} - [2\,\Delta H_f^\circ\ FeO(s) + \frac{1}{2}(0)]$	1 point for correct stoichiometry
$+544\text{ kJ mol}^{-1} = -2\,\Delta H_f^\circ\ FeO(s)$ $-272\text{ kJ mol}^{-1} = \Delta H_f^\circ\ FeO(s)$	1 point for correct calculation

Question 3

The first-order decomposition of a colored chemical species, X, into colorless products is monitored with a spectrophotometer by measuring changes in absorbance over time. Species X has a molar absorptivity constant of 5.00×10^3 cm^{-1} M^{-1} and the path length of the cuvette containing the reaction mixture is 1.00 cm. The data from the experiment are given in the table below.

[X] (M)	Absorbance	Time (min)
?	0.600	0.0
4.00×10^{-5}	0.200	35.0
3.00×10^{-5}	0.150	44.2
1.50×10^{-5}	0.075	?

(a) Calculate the initial concentration of the colored species.

$A = abc$ $c = \dfrac{A}{ab} = \dfrac{0.600}{(5.00\times10^3\ \text{cm}^{-1}M^{-1})(1.00\ \text{cm})}$ $= 1.20 \times 10^{-4}\ M$ *OR* $A_0 = abc_0 \qquad A_1 = abc_1$ $\dfrac{A_0}{c_0} = \dfrac{A_1}{c_1} \qquad \dfrac{0.600}{c_0} = \dfrac{0.200}{4.00\times10^{-5}}$ $c_0 = 1.20 \times 10^{-4}\ M$	1 point for concentration of X

(b) Calculate the rate constant for the first-order reaction using the values given for concentration and time. Include units with your answer.

Using the first two readings, $\ln[\text{X}]_t - \ln[\text{X}]_0 = -kt$ *OR* $\ln\dfrac{[\text{X}]_t}{[\text{X}]_0} = -kt$ $\ln\dfrac{4.00\times10^{-5}}{1.20\times10^{-4}} = -k(35.0\ \text{min})$ $\ln(0.333) = -k(35.0\ \text{min})$ $-1.10 = -k(35.0\ \text{min})$ $k = 3.14 \times 10^{-2}\ \text{min}^{-1}$	1 point for magnitude and correct sign of rate constant 1 point for correct units

Question 3 (cont'd.)

(c) Calculate the number of minutes it takes for the absorbance to drop from 0.600 to 0.075.

$\ln \frac{[X]_t}{[X]_0} = -kt$ $\ln \frac{1.50\times10^{-5}}{1.20\times10^{-4}} = (-3.14 \times 10^{-2}\ \text{min}^{-1})\ t$ $\ln (0.125) = (-3.14 \times 10^{-2}\ \text{min}^{-1})\ t$ $-2.08 = (-3.14 \times 10^{-2}\ \text{min}^{-1})\ t$ $t = 66.2$ min	1 point for correct substitution 1 point for correct answer <u>Note</u>: students may use half-lives to answer this question.

(d) Calculate the half-life of the reaction. Include units with your answer.

$\ln \frac{[X]_t}{[X]_0} = -kt$ $\ln \frac{0.5\,[X]_0}{[X]_0} = (-3.14 \times 10^{-2}\ \text{min}^{-1})\ t_{1/2}$ $\ln (0.5) = (-3.14 \times 10^{-2}\ \text{min}^{-1})\ t_{1/2}$ $-0.693 = (-3.14 \times 10^{-2}\ \text{min}^{-1})\ t_{1/2}$ $22.1\ \text{min} = t_{1/2}$ *<u>OR</u>* $t_{1/2} = \frac{0.693}{k}$ $t_{1/2} = \frac{0.693}{3.14\times10^{-2}\ \text{min}^{-1}} = 22.1$ min	1 point for correct magnitude 1 point for the correct units (1 point for the half-life equation if no k is given)

(e) Experiments were performed to determine the value of the rate constant for this reaction at various temperatures. Data from these experiments were used to produce the graph below, where T is temperature. This graph can be used to determine the activation energy, E_a, of the reaction.

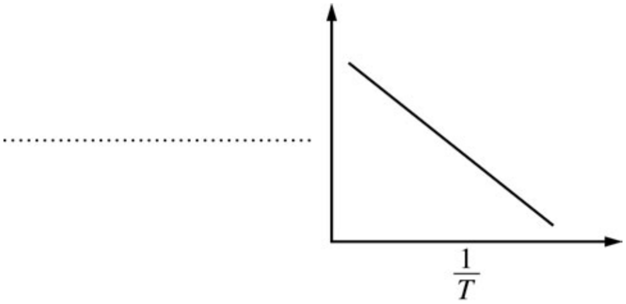
(i) Label the vertical axis of the graph.

The vertical axis should be labeled $\ln k$.	1 point

Question 3 (cont'd.)

(ii) Explain how to calculate the activation energy from this graph.

The slope of the line is related to the activation energy: $$\text{slope} = -\frac{E_a}{R}$$	1 point for recognizing that the slope must be measured
To determine the activation energy for the reaction, multiply the slope by -8.314 J mol^{-1} K^{-1}.	1 point for the correct explanation of how to obtain the activation energy



Question 4

Write the formulas to show the reactants and the products for any FIVE of the laboratory situations described below. Answers to more than five choices will not be graded. In all cases, a reaction occurs. Assume that solutions are aqueous unless otherwise indicated. Represent substances in solution as ions if the substances are extensively ionized. Omit formulas for any ions or molecules that are unchanged by the reaction. You need not balance the equations.

Example: A strip of magnesium is added to a solution of silver nitrate.

Ex.	$Mg + Ag^+ \rightarrow Mg^{2+} + Ag$

(a) A solution of copper(II) sulfate is spilled onto a sheet of freshly polished aluminum metal.

$Cu^{2+} + Al \rightarrow Cu + Al^{3+}$	1 point for reactant(s), 2 points for product(s)

(b) Dimethyl ether is burned in air.

$C_2H_6O + O_2 \rightarrow CO_2 + H_2O$ *OR* $C_2H_6O + O_2 \rightarrow CO_2 + H_2O + CO$	1 point for reactant(s), 2 points for product(s)

(c) A 0.1 *M* nitrous acid solution is added to the same volume of a 0.1 *M* sodium hydroxide solution.

$HNO_2 + OH^- \rightarrow H_2O + NO_2^-$	1 point for reactant(s), 2 points for product(s)

(d) Hydrogen iodide gas is bubbled into a solution of lithium carbonate.

$HI + CO_3^{2-} \rightarrow H_2O + I^- + CO_2$ *OR* $HI + CO_3^{2-} \rightarrow I^- + HCO_3^-$	1 point for reactant(s), 2 points for product(s)

(e) An acidic solution of potassium dichromate is added to a solution of iron(II) nitrate.

$Cr_2O_7^{2-} + Fe^{2+} + H^+ \rightarrow Fe^{3+} + Cr^{3+} + H_2O$	1 point for reactant(s), 2 points for product(s)

Question 4 (cont'd.)

(f) Excess concentrated aqueous ammonia is added to a solution of nickel(II) bromide.

$Ni^{2+} + NH_3 \rightarrow [Ni(NH_3)_6]^{2+}$	1 point for reactant(s), 2 points for product(s)

<u>Note:</u> complexes with numbers of NH_3 ligands less than 6 were also accepted.

(g) A solution of sodium phosphate is added to a solution of aluminum nitrate.

$PO_4^{3-} + Al^{3+} \rightarrow AlPO_4$	1 point for reactant(s), 2 points for product(s)

(h) Concentrated hydrochloric acid is added to a solution of sodium sulfide.

$H^+ + OH^- + HS^- \rightarrow H_2O + H_2S$ *OR* $H^+ + HS^- \rightarrow H_2S$	1 point for reactant(s), 2 points for product(s)

Question 5

In a laboratory class, a student is given three flasks that are labeled *Q*, *R*, and *S*. Each flask contains one of the following solutions: 1.0 *M* $Pb(NO_3)_2$, 1.0 *M* NaCl, or 1.0 *M* K_2CO_3. The student is also given two flasks that are labeled *X* and *Y*. One of these flasks contains 1.0 *M* $AgNO_3$, and the other contains 1.0 *M* $BaCl_2$. This information is summarized in the diagram below.

(a) When the student combined a sample of solution *Q* with a sample of solution *X*, a precipitate formed. A precipitate also formed when samples of solutions *Q* and *Y* were combined.

(i) Identify solution Q .

Solution *Q* is K_2CO_3	1 point for correct identification of solution *Q*

(ii) Write the chemical formulas for each of the two precipitates.

Ag_2CO_3 and $BaCO_3$	1 point each for correct formulas of carbonate precipitates

(b) When solution *Q* is mixed with solution *R*, a precipitate forms. However, no precipitate forms when solution *Q* is mixed with solution *S*.

(i) Identify solution *R* and solution *S*.

Solution *R* is $Pb(NO_3)_2$ and solution *S* is NaCl.	1 point each for identification of solutions *R* and *S*

(ii) Write the chemical formula of the precipitate that forms when solution *Q* is mixed with solution *R*.

$PbCO_3$	1 point for correct formula of the precipitate

Question 5 (cont'd.)

(c) The identity of solution *X* and solution *Y* are to be determined using only the following solutions: 1.0 *M* $Pb(NO_3)_2$, 1.0 *M* NaCl , and 1.0 *M* K_2CO_3 .

(i) Describe a procedure to identify solution X and solution Y.

The identities of solutions *X* and *Y* can be determined by adding a sample of NaCl to each solution. *OR* The identities of solutions *X* and *Y* can be determined by adding a sample of $Pb(NO_3)_2$ to solution *X* and solution *Y*.	1 point for a correct identification of a reagent that will differentiate between the two solutions

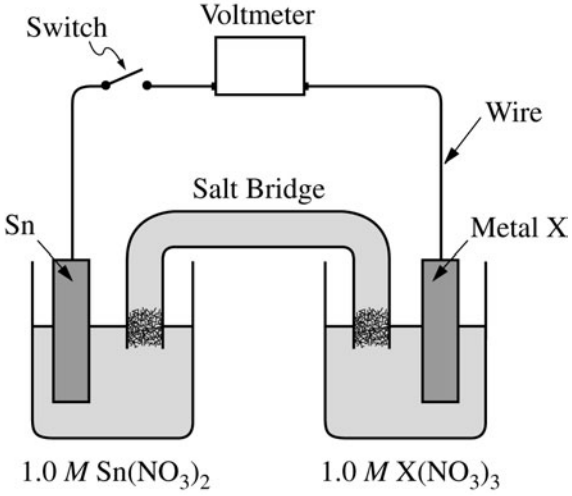
(ii) Describe the observations that would allow you to distinguish between solution *X* and solution *Y*.

NaCl will form a white precipitate when added to a solution of $AgNO_3$, but will not form a precipitate when added to a solution of $BaCl_2$. *OR* $Pb(NO_3)_2$ will form a white precipitate when added to a solution of $BaCl_2$, but will not form a precipitate when added to a solution of $AgNO_3$.	1 point for correct observation that a precipitate is formed in one solution but not the other

(iii) Explain how the observations would enable you to distinguish between solution X and solution *Y*.

When NaCl is added to solution *X*, a precipitate of AgCl forms if solution *X* is $AgNO_3$. If no precipitate forms, solution *X* must be $BaCl_2$. The same logic can be used to identify solution *Y*. *OR* When $Pb(NO_3)_2$ is added to solution *X*, a precipitate of $PbCl_2$ forms if solution *X* is $BaCl_2$. If no precipitate forms, solution *X* must be $AgNO_3$. The same logic can be used to identify solution *Y* .	1 point for correct identification of the solution that forms a precipitate

Question 6



An electrochemical cell is constructed with an open switch, as shown in the diagram above. A strip of Sn and a strip of an unknown metal, X, are used as electrodes. When the switch is closed, the mass of the Sn electrode increases. The half-reactions are shown below.

$$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s) \qquad E° = -0.14 \text{ V}$$
$$X^{3+}(aq) + 3e^- \rightarrow X(s) \qquad E° = ?$$

(a) In the diagram above, label the electrode that is the cathode. Justify your answer.

The Sn (tin) electrode is the cathode.	1 point for identifying Sn as the cathode
The increase in mass indicates that reduction occurs at the Sn electrode: $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$ Reduction occurs at the cathode.	1 point for reasoning based on increase in mass

(b) In the diagram above, draw an arrow indicating the direction of the electron flow in the external circuit when the switch is closed.

Diagram should have arrow showing electrons flowing from the anode towards the cathode.	1 point for correct direction of electron flow

Question 6 (cont')

(c) If the standard cell potential, E°_{cell}, is +0.60 V, what is the standard reduction potential, in volts, for the X^{3+}/X electrode?

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ $+0.60\ \text{V} = -0.14\ \text{V} - E^\circ_{anode}$ $E^\circ_{anode} = -0.74\ \text{V}$	1 point for correct potential with correct sign

(d) Identify metal X.

Cr	1 point for correct metal

(e) Write a balanced net-ionic equation for the overall chemical reaction occurring in the cell.

$3\ Sn^{2+} + 2\ Cr \rightarrow 3\ Sn + 2\ Cr^{3+}$	1 point for correctly balanced net-ionic equation

(f) In the cell, the concentration of Sn^{2+} is changed from 1.0 *M* to 0.50 *M*, and the concentration of X^{3+} is changed from 1.0 *M* to 0.10 *M*.

(i) Substitute all the appropriate values for determining the cell potential, E_{cell}, into the Nernst equation. (Do not do any calculations.)

$E_{cell} = E^\circ_{cell} - \frac{0.0592}{n} \log \frac{[Cr^{3+}]^2}{[Sn^{2+}]^3}$ $E_{cell} = +0.60\ \text{V} - \frac{0.0592}{6} \log \frac{[0.10]^2}{[0.50]^3}$	1 point for using $E^\circ_{cell} = +0.60$ V 1 point for using $n = 6$ 1 point for substituting correctly into the Q expression based on the equation in part (e)

(ii) On the basis of your response in part (f) (i), will the cell potential, E_{cell}, be greater than, less than, or equal to the original E°_{cell}? Justify your answer.

E_{cell} will be greater (more positive). Since the Q ratio is a number less than 1, the log of the ratio will be negative. A negative times a negative is positive. Thus, $-\frac{0.0592}{6} \log \frac{[0.10]^2}{[0.50]^3}$ increases E_{cell}	1 point for the correct prediction with an explanation based on Q

Question 7

Use appropriate chemical principles to account for each of the following observations. In each part, your response must include specific information about both substances.

(a) At 25°C and 1 atm, F_2 is a gas, whereas I_2 is a solid.

Both F_2 and I_2 are nonpolar, so the only intermolecular attractive forces are London dispersion forces. I_2 is solid because the electrons in the I_2 molecule occupy a larger volume and are more polarizable compared to the electrons in the F_2 molecule. As a result, the dispersion forces are considerably stronger in I_2 compared to F_2.	1 point for indicating that both molecules have dispersion forces as IMFs 1 point for indicating that I_2 molecules are more polarizable than F_2 molecules

(b) The melting point of NaF is 993°C, whereas the melting point of CsCl is 645°C.

Both NaF and CsCl are ionic compounds with the same charges on the cations and anions. The ionic radius of Na^+ is smaller than the ionic radius of Cs^+ and the ionic radius of F^- is smaller than the ionic radius of Cl^-. Therefore, the ionic centers are closer in NaF than in CsCl. Melting occurs when the attraction between the cation and the anion are overcome due to thermal motion. Since the lattice energy is inversely proportional to the distance between the ion centers (Coulomb's Law), the compound with the smaller ions will have the stronger attractions and the higher melting point.	1 point for indicating that NaF and CsCl are both ionic compounds (or are composed of M^+ and X^- ions) 1 point for indicating that the strength of these forces is determined by the distance between the ionic centers (or the size of the ions)

(c) The shape of the ICl_4^- ion is square planar, whereas the shape of the BF_4^- ion is tetrahedral.

The central iodine atom in ICl_4^- has four bonding pairs and two lone pairs of electrons on the central iodine atom, so the molecular geometry is square planar. BF_4^- has four bonding pairs and no lone pairs on the central boron atom, so the molecular geometry is tetrahedral.	2 points for indicating that ICl_4^- has two unshared electron pairs, but BF_4^- has no unshared pairs Note: 1 point earned if student gives incorrect numbers of unshared electron pairs but indicates that difference in number of unshared electron pairs determines difference in geometry.

Question 7 (cont'd.)

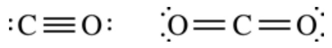
(d) Ammonia, NH_3, is very soluble in water, whereas phosphine, PH_3, is only moderately soluble in water.

Ammonia has hydrogen-bonding intermolecular forces, whereas phosphine has dipole-dipole and/or dispersion intermolecular forces. Water also has hydrogen-bonding intermolecular attractive forces. Ammonia is more soluble in water than phosphine because ammonia molecules can hydrogen-bond with water molecules, whereas phosphine molecules cannot hydrogen-bond with water molecules.	1 point for indicating that NH_3 can form hydrogen bonds but PH_3 cannot 1 point for indicating that NH_3 can form hydrogen bonds with water, but PH_3 cannot

Question 8

Answer the following questions about carbon monoxide, CO(*g*), and carbon dioxide, $CO_2(g)$. Assume that both gases exhibit ideal behavior.

(a) Draw the complete Lewis structure (electron-dot diagram) for the CO molecule and for the CO_2 molecule.

	1 point for each correct, complete Lewis structure

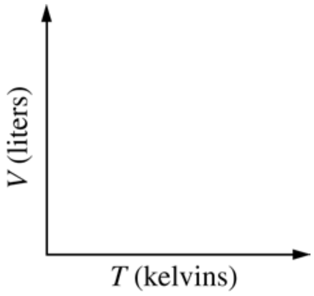
(b) Identify the shape of the CO_2 molecule.

CO_2 has a linear molecular geometry	1 point for correct molecular geometry

(c) One of the two gases dissolves readily in water to form a solution with a pH below 7. Identify the gas and account for this observation by writing a chemical equation.

The gas that produces a pH less than 7 when added to water is CO_2 . The reaction that accounts for this is $CO_2(g) + H_2O(l) \rightarrow HCO_3^-(aq) + H^+(aq)$ <u>OR</u> , $CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$	1 point for identifying CO_2 in a correct chemical equation

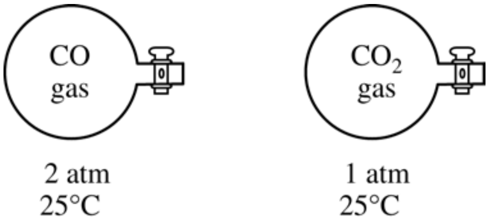
(d) A 1.0 mole sample of CO(*g*) is heated at constant pressure. On the graph below, sketch the expected plot of volume versus temperature as the gas is heated.



The graph should have a straight line with a positive slope.	1 point for drawing a correct line

Question 8 (cont'd.)

(e) Samples of CO(g) and CO_2(g) are placed in 1 L containers at the conditions indicated in the diagram below.



(i) Indicate whether the average kinetic energy of the CO_2(g) molecules is greater than, equal to, or less than the average kinetic energy of the CO(g) molecules. Justify your answer.

The average kinetic energy is the same for both samples because the temperature is the same for both samples. Average kinetic energy is proportional to temperature.	1 point for correct answer and explanation

(ii) Indicate whether the root-mean-square speed of the CO_2(g) molecules is greater than, equal to, or less than the root-mean-square speed of the CO(g) molecules. Justify your answer.

The root-mean-square speed for CO_2 is lower than the root-mean-square speed for CO. The molar mass of CO_2 is higher than the molar mass of CO. The root-mean-square speed is inversely proportional to the square root of the molar mass of the gas.	1 point for correct answer and explanation

(iii) Indicate whether the number of CO_2(g) molecules is greater than, equal to, or less than the number of CO(g) molecules. Justify your answer.

There are fewer CO_2 molecules than CO molecules. The CO_2 molecules exert half the pressure of the CO molecules (at the same T and V), so there must be half as many molecules present.	1 point for correct answer and explanation