QUESTION 1: **[28 MARKS]**

1.1. When methanol, CH_3OH, is dissolved in water, a non-conducting solution result. When acetic acid, CH_3COOH, dissolves in water, the solution is weakly conducting and acidic in nature. Describe what happens upon dissolution in the two cases, and account for the different results. (5)

1.2. You come across a beaker that contains water, aqueous ammonium acetate, and a precipitate of calcium phosphate.

1.2.1. Write the balanced molecular equation for a reaction between two solutions containing ions that could produce this solution. Include all phase labels (states). (4)

1.2.2. Write the net ionic equation for the reaction in part 1.2.1. Include all phase labels (states). (3)

1.3. The formula for the arsenate ion is AsO_4^{3-}. What is the formula for arsenous acid? (2)

1.4. Give the name and formula of the acid corresponding to the chlorite ion. (2)

1.5. Using the net ionic equation below

1.5.1. Determine the volume (in mL) of a 0.150 M $FeCl_3$ solution needed to react completely with 20.0 mL of 0.0450 M $AgNO_3$ solution. (3)

1.5.2. How many grams of AgCl will be formed? (3)

The net ionic equation for the reaction is

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(aq)$

1.6. How can you tell if a reaction is an oxidation-reduction reaction? (2)

1.7. The following is a redox reaction

$2KCl + MnO_2 + 2H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + Cl_2 + 2H_2O$

Indicate which element is oxidized, which element is reduced as well as the oxidizing and the reducing agents. (4)

QUESTION 2: (START ON A NEW PAGE) **[24 MARKS]**

2.1. A child receives a helium balloon in a shopping mall. When he goes outside during a snowstorm, the balloon decreases in size. Which gas law is this an example of? (2)

2.2. Consider a container of gas under a particular P, V, T set of conditions. Describe how the pressure would change if the volume were doubled while the absolute temperature was increasing by a factor of two. (3)

2.3. Rank the following gases from least dense to most dense at 1.00 atm and 298 K: SO_2, HBr, CO_2. Explain. (2)

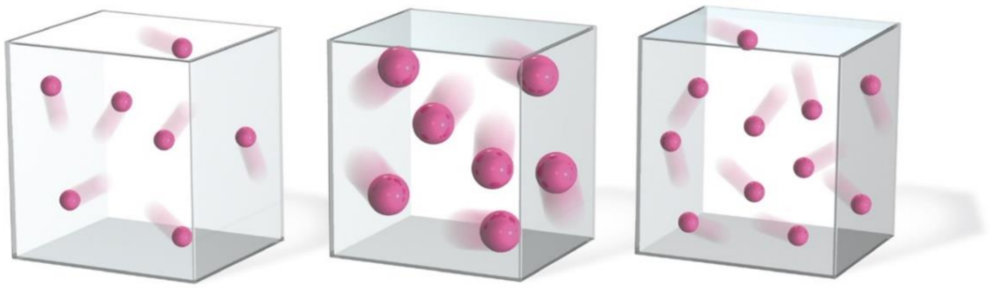
2.4. Given the following reaction, $2\ CH_3CHO(g) + O_2(g) \rightarrow 2\ CH_3CO_2H(l)$, calculate the final pressure in a 250-mL reaction vessel if 0.566 moles of CH_3CHO are allowed to react with 0.323 moles of O_2 at 42 °C. (4)

2.5. A mixture of 8.0 mol Ne and 8.0 mol Xe are at STP in a rigid container. Both gases have the same average kinetic energy. True/False. (2)

2.6. A sample gas mixture occupies a volume of 1.446 liters when the temperature is 159.0 °C and the pressure is 694.8 torr. The mole fraction of methane in this sample is 0.3260. How many methane molecules are there in the sample? _______ (4)

2.7. Using the kinetic molecular theory, explain how raising the temperature of a gas causes it to expand at constant pressure. (2)

2.8. Which sample of an ideal gas has the greatest pressure? Justify your answer. Assume that mass of each particle is proportional to its size and that all the gas samples are at the same temperature. (5)



QUESTION 3: (START ON A NEW PAGE) **[27 MARKS]**

3.1. Explain why the anion, HCO_3^-, can be classified as an amphoteric species. (3)

3.2. For the system

$H_3PO_4(aq) + COOH^-(aq) \rightleftharpoons HCOOH(aq) + H_2PO_4^-(aq)$

the position of equilibrium lies to the right. What is the strongest base in this reaction? (2)

3.3. You can obtain the pH of a 0.100 M HCl solution by assuming that the entire H_3O^+ ion comes from the HCl, in which case the pH equals - log 0.100 = 1.00. But if you want the pH of a solution that is 1.00×10^{-7} M HCl, you need to account for any H_3O^+ ion coming from water. Why? (2)

3.4. Calculate the concentration of an aqueous solution of $Ca(OH)_2$ that has a pH of 10.05. (4)

3.5. Which weak acid solution: 0.100 M $HC_2H_3O_2$, 0.500 M $HC_2H_3O_2$ and 0.0100 M $HC_2H_3O_2$

3.5.1. has the greatest percent ionization? Explain your choice. (2)

3.5.2. has the lowest pH? Explain your choice. (2)

3.6. Which has the highest pH; a 0.1 M solution of a base with pKb = 4.5 or one with pKb = 6.5? Explain your choice. (3)

3.7. Ethylamine, $CH_3CH_2NH_2$, has a strong, pungent odor like that of ammonia. Like ammonia, it is a Brønsted base. A 0.10 M solution has a pH of 11.87. Using the ICE table, calculate the K_b and pK_b values for ethylamine. (6.5)

3.7.1. What is the percentage ionization of the ethylamine? (2.5)

QUESTION 4: (START ON A NEW PAGE) **[26 MARKS]**

4.1. Why do atoms only emit certain wavelengths of light when they are excited? (Why do line spectra exist?) (3)

4.2. Describe Niels Bohr's model of the structure of the hydrogen atom. (2)

4.3. According to the quantum mechanical model for the hydrogen atom, which electron transition would produce light with the longer wavelength: 4p→2s or 4p→2p. Explain your choice. (4)

4.4. The maximum number of electrons in an atom that can have the following exact same set of quantum numbers is ________. (2)

$n = 4 \quad l = 3 \quad m_l = -2 \quad m_s = +1/2$

4.5. Explain why every shell contains an s subshell? (2)

4.6. The 1st three ionization energies for calcium are 590, 1145, and 4912 kJ/mol, representing the formation of Ca^+, Ca^{2+}, and the Ca^{3+} ions. Briefly explain why there is such a large change in the required energy to remove the 3rd electron from calcium. (2)

4.7. What is the Pauli exclusion principle? What effect does it have on the filling of orbitals by electrons? (3)

4.8. What is meant by the term effective nuclear charge? (2)

4.8.1. How does the effective nuclear charge experienced by the valence electrons of an atom vary going from left to right across a period of the periodic table? Explain. (2)

4.9. Write an orbital diagram for the ground state of the zinc atom. (2)

4.9.1. Is the atomic substance diamagnetic or paramagnetic? Explain your answer. (2)

END OF PAPER

PSFT0B1 DATA

N_A (Avogadro's number) = 6.022×10^{23}

1 atm = 760 mmHg = 760 torr = 1.01×10^5 Pa = 1.013 bar

R (gas constant) = 0.08206 L·atm/K·mol

R (gas constant) = $8.314\,4621(75) \times 10^{-2}$ L bar K^{-1} mol^{-1}

P_{H2O} (25 °C) = 23.78 mmHg

h (Planck's constant) = 6.626×10^{-34} J.s

F (Faraday constant) = 9.6485×10^4 C/mol e^-

c (speed of light) = 2.998×10^8 m/s

b = 2.18×10^{-18} J

SOLUBILITY TABLE:

Anion	Solubility rule
Mostly soluble	
Acetates, nitrates and perchlorates	All cations form *soluble* compounds. ($KClO_4$ and $AgC_2H_3O_2$ slightly soluble)
Chlorides, bromides, iodides	All cations form *soluble* compounds except Hg_2^{2+}, Ag^+ and Pb^{2+} ($PbCl_2$ and $PbBr_2$ slightly soluble)
Sulfates	All cations form *soluble* compounds except Pb^{2+}, Ba^{2+} and Sr^{2+} (Ca^{2+} and Ag^+ form slightly soluble compounds)
Mostly insoluble	
Carbonates and phosphates	All cations form *insoluble* compounds except Group IA metals and NH_4^+
Sulfides	All cations form *insoluble* compounds except Group IA and IIA metals and NH_4^+
Hydroxides	All cations form *insoluble* compounds except Group IA metals, Ba^{2+} and Sr^{2+} and NH_4^+ [$Ca(OH)_2$ is slightly soluble]

TABLE 4.1 Solubility Rules for Ionic Compounds in Water

Soluble Compounds

1. All compounds of the alkali metals (Group 1A) are soluble.
2. All salts containing NH_4^+, NO_3^-, ClO_4^-, ClO_3^-, and $C_2H_3O_2^-$ are soluble.
3. All chlorides, bromides, and iodides (salts containing Cl^-, Br^-, or I^-) are soluble except when combined with Ag^+, Pb^{2+}, and Hg_2^{2+} (note the subscript "2").
4. All sulfates (salts containing SO_4^{2-}) are soluble except those of Pb^{2+}, Ca^{2+}, Sr^{2+}, Hg_2^{2+}, and Ba^{2+}.

Insoluble Compounds

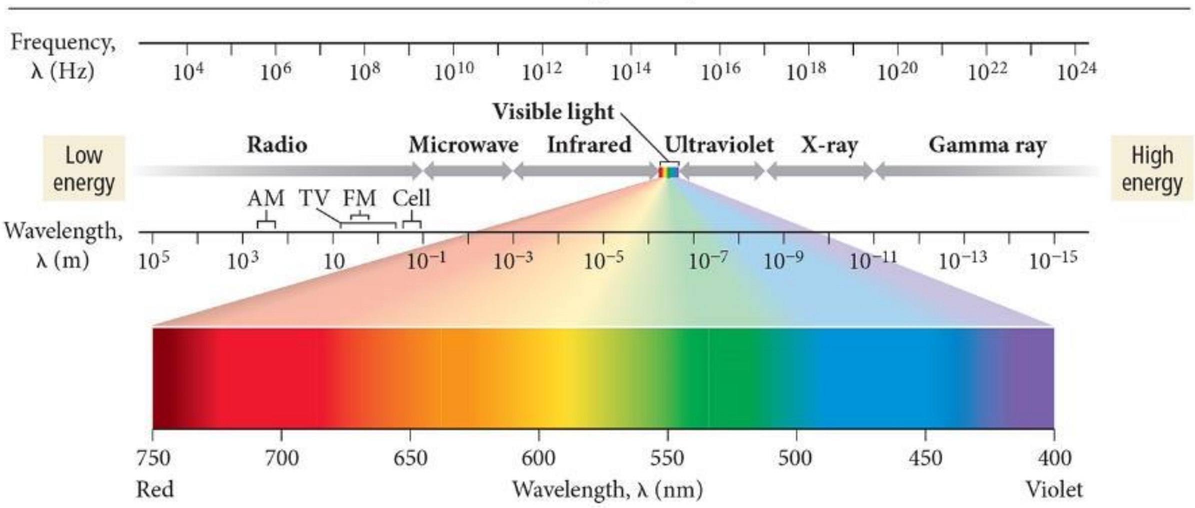
5. All metal hydroxides (ionic compounds containing OH^-) and all metal oxides (ionic compounds containing O^{2-}) are insoluble except those of Group 1A and of Ca^{2+}, Sr^{2+}, and Ba^{2+}.

   When metal oxides do dissolve, they react with water to form hydroxides. The oxide ion, O^{2-}, does not exist in water. For example,

$$Na_2O(s) + H_2O \longrightarrow 2NaOH(aq)$$

6. All salts that contain PO_4^{3-}, CO_3^{2-}, SO_3^{2-}, and S^{2-} are insoluble, except those of Group 1A and NH_4^+.

The Electromagnetic Spectrum



IUPAC Periodic Table of the Elements

Key:

atomic number
Symbol
name
conventional atomic weight
standard atomic weight

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.008 [1.0078, 1.0082]																	2 He 4.0026
3 Li 6.94 [6.938, 6.997]	4 Be 9.0122											5 B 10.81 [10.806, 10.821]	6 C 12.011 [12.009, 12.012]	7 N 14.007 [14.006, 14.008]	8 O 15.999 [15.999, 16.000]	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305 [24.304, 24.307]											13 Al 26.982	14 Si 28.085 [28.084, 28.086]	15 P 30.974	16 S 32.06 [32.059, 32.076]	17 Cl 35.45 [35.446, 35.457]	18 Ar 39.95 [39.792, 39.963]
19 K 39.098	20 Ca 40.078(4)	21 Sc 44.956	22 Ti 47.867	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845(2)	27 Co 58.933	28 Ni 58.693	29 Cu 63.546(3)	30 Zn 65.38(2)	31 Ga 69.723	32 Ge 72.630(8)	33 As 74.922	34 Se 78.971(8)	35 Br 79.904 [79.901, 79.907]	36 Kr 83.798(2)
37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.224(2)	41 Nb 92.906	42 Mo 95.95	43 Tc	44 Ru 101.07(2)	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60(3)	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57-71 lanthanoids	72 Hf 178.49(2)	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23(3)	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38 [204.38, 204.39]	82 Pb 207.2	83 Bi 208.98	84 Po	85 At	86 Rn
87 Fr	88 Ra	89-103 actinoids	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm	62 Sm 150.36(2)	63 Eu 151.96	64 Gd 157.25(3)	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97
89 Ac	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

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