Dawson COLLEGE

Department of Chemistry & Chemical Technology
GENERAL CHEMISTRY (202-SN1-RE)

PRACTICE FINAL EXAMINATION

This practice examination is intended to reflect the structure and timing of the upcoming General Chemistry (202-SN1-RE) Final Examination; however, the question style, difficulty level, and exact formatting on the actual exam may differ.

Furthermore, while this practice test contains no laboratory questions, *the actual final exam will include one question dedicated to laboratory work*. Students are advised to review all in-semester laboratory exercises and lab reports as part of their final preparation.

Recommended Preparation Guidelines

- **Simulate Authentic Test Conditions:** Complete this practice exam in a quiet room without distractions. Strictly enforce the **3-hour** time limit to gauge your pacing.
- **Work Closed-Book:** To get the maximum learning benefit, treat this sample exam as a true practice. Attempt all questions without checking textbooks, instructors' notes or solutions mid-exam. Once completed, grade your exam using the answer key attached at the end.

General Examination Instructions

- Non-programmable calculators are permitted.
- A periodic table and important constants are attached to the back of the exam booklet.
- Show all work for calculation questions, ensuring all numerical answers are expressed with the correct number of significant figures and appropriate units.

1. a) Fill in the blank spaces (isotope symbols, number of protons, number of electrons, number of neutrons, mass number, and charge) indicated with question marks in the boxes of the table below.

Isotopes	Protons	Electrons	Neutrons	Mass number	Net charge
	35			80	1-
	88	86		226	

b) An element forms an ionic compound with nitrate anions, having the formula $X(NO_3)_4$. If the ion of element X has a mass number of 207 and has 78 electrons,

i. What is the identity of the element? _____

ii. How many neutrons does the element have? _____

2. The following multiple-choice questions and statements pertain to elements and their periodic properties. There is only one correct answer for each question or statement. Use a tick (√) to select your answer.

a. Which group 1 element is not a metal?

i. H □
ii. Li □
iii. Fr □
iv. K □

b. All of the following elements are nonmetals except:

i. Sb □
ii. C □
iii. Cl □
iv. O □

c. The horizontal rows of the periodic table are known as the

i. Periods □
ii. Groups □
iii. Metalloids □
iv. Nonmetals □

d. When a metal and a nonmetal react, the ___ tends to lose electrons and the ___ tends to gain electrons. Fill in the blank spaces from the choices below.

i. metal, nonmetal □
ii. nonmetal, nonmetal □
iii. metal, metal □
iv. nonmetal, metal □

e. Gaseous elements characterized as being unreactive are found in which group of the periodic table?

i. 5A (or Group 15) □
ii. 6A (or Group 16) □
iii. 7A (or Group 17) □
iv. 8A (or Group 18) □

f. Which of the following elements is classified as a metalloid?

i. calcium □
ii. germanium □
iii. fluorine □
iv. uranium □

3. a) Name the compounds whose formulae are given below:

i) Mg_3P_2 _____

ii) SF_6 _____

iii) $HClO(aq)$ _____

iv) NaH_2PO_4 _____

v) $FeSO_4 \bullet 2H_2O$ _____

b) Write formulae for the compounds whose names are given below:

i) Mercury(I) chloride _____

ii) Aluminum nitride _____

iii) Sulfurous acid _____

iv) Disulfur dichloride _____

v) Sodium dichromate _____

4. a) Calculate the percent composition by mass of all elements in Na_2CO_3.

b) An unknown compound has the composition 28.03 % magnesium, 21.60% silicon, and 1.16 % hydrogen atoms with the remainder of the mass as oxygen atoms. The molar mass of this unknown compound is 520.8 g/mol. Determine the empirical and molecular formulas of the unknown compound.

5. a) Complete and balance the following chemical equations:

i. Precipitation:

K_3PO_4 *(aq)* + $MnCl_2$ *(aq)* →

ii. Oxidation-Reduction:

$C_5H_{10}O_5$ *(s)* + O_2 *(g)* → CO_2 *(g)* + H_2O *(l)*

iii. Acid-base:

HNO_3 *(aq)* + $NaHCO_3$ *(aq)* →

b) Which of the following is considered a strong electrolyte?

i. $HC_2H_3O_2$
ii. $C_6H_{12}O_6$
iii. $PbBr_2$
iv. NH_4NO_3
v. CH_3OH

c) Rank the following solutions A, B, and C, from lowest to highest number of moles of chloride ions.

Solution A 10.0 mL 0.075 mol/L $SnCl_4$
Solution B 20.0 mL 0.100 mol/L $CaCl_2$
Solution C 25.0 mL 0.050 mol/L $AlCl_3$

_____ < _____ < _____

Lowest Highest

6. a) A sample of gas has a mass of 0.576 g. At a temperature of 85.0 °C and a pressure of 753 torr it has a volume of 117 mL. What is the molar mass of the gas?

b) Titanium is a strong, lightweight, corrosion-resistant metal that is used in rockets, aircraft, jet engines, and bicycle frames. It is prepared by the following reaction at temperatures between 950°C and 1150°C:

$$TiCl_4\ (g) + 2\ Mg\ (l) \rightarrow Ti\ (s) + 2\ MgCl_2\ (l)$$

i. If 4.97×10^3 g of $TiCl_4$ is reacted with 1.125×10^3 g of Mg, what is the theoretical yield of Ti (in grams)?

ii. If 985 g of Ti is obtained, what is the percent yield?

7. a) You mix an aqueous solution of $AgNO_3$ with an aqueous solution of Na_2S.

i. Write the formula equation.

ii. Write the complete ionic equation.

iii. Write the net ionic equation.

iv. Give the spectator ions.

b) How many grams of Na_2S are required to react completely with 27.8 mL of 0.163 M $AgNO_3$?

c) How many grams of precipitate are obtained from the reaction?

8. a) Predict whether each of the following compounds is soluble or insoluble in water.

	Soluble	Insoluble (Precipitate)
Na_2SO_4	☐	☐
KOH	☐	☐
$BaCO_3$	☐	☐
$AgNO_3$	☐	☐

b) For the following oxidation-reduction reaction:

$$MnO_4^-{}_{(aq)} + C_2O_4^{2-}{}_{(aq)} \rightarrow MnO_2{}_{(s)} + CO_3^{2-}{}_{(aq)}$$

i. Identify the oxidizing agent, the reducing agent, the species (atom) being oxidized and the species (atom) being reduced.

Oxidizing agent: _____ Element being oxidized: _____

Reducing agent: _____ Element being reduced: _____

ii. Balance this oxidation-reduction reaction that occurs in acidic solution using the half-reaction method. Clearly show your work.

9. a) Which of the following electron transitions on the hydrogen atom will produce a photon with the longest wavelength?

i. n = 8 to n = 6
ii. n = 4 to n = 1
iii. n = 3 to n = 6
iv. n = 3 to n = 1

b) Using the Bohr model, determine the wavelength of the photon emitted when an electron in an ion of Li^{2+} moves from the orbit with n = 4 to the orbit with n = 2. Show your calculations.

c) One of the radiographic devices used in a dentist's office emits an X-ray with a wavelength of 2.090×10^{-11}m.

i) What is the energy, in joules, for one photon of this X-ray?

ii) What is the frequency for one photon of this X-ray?

10. a) Write the full ground-state electron configuration of the following:

i. Br : _____

ii. Mg^{2+}: _____

b) Provide the abbreviated (or noble gas) ground-state electronic configuration for each ion.

Cu^{+} _____

Se^{2-} _____

c) Which ion with a 1+ charge has the electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$?

d) Determine if the following ground state electron configurations are correct:

		correct	incorrect
i.	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 4d^{10}\ 4p^6\ 5s^1$	☐	☐
ii.	$[Xe]\ 6s^2\ 4f^{10}$	☐	☐
iii.	$1s^2\ 2s^2\ 2p^6\ 3s^3$	☐	☐
iv.	$[Ar]\ 4s^1\ 3d^{10}$	☐	☐

e) How many electrons could exist in an atom with the following quantum numbers?

i. $n = 3,\ \ell = 1$: _____

ii. $n = 5,\ m_\ell = 3$: _____

f) For each combination of quantum numbers, indicate if it is allowed or not

		allowed	not allowed
i.	$n = 3, \ell = 2, \ m_\ell = -3, m_s = -\frac{1}{2}$	☐	☐
ii.	$n = 0, \ell = 0, \ m_\ell = 0, \ m_s = +\frac{1}{2}$	☐	☐
iii.	$n = 3, \ell = 1, \ m_\ell = -1, m_s = +\frac{1}{2}$	☐	☐
iv.	$n = 4, \ell = -3, m_\ell = -1, m_s = +1$	☐	☐

11. a) Circle your choice. Which of the following elements has the lowest ionization energy?

Mg Ba Cs

b) Down a group of the periodic table

		increases	decreases
i.	Electron affinity	☐	☐
ii.	Atomic radius	☐	☐

c) For each set of bonds, circle the bond that is more polar.

C-H vs C-N Si-O vs C-O N-H vs P-H

d) Assign the partial charge δ^+ or δ^- to each atom in each bond.

H - F S - O P - Cl

e) Rank the following ions in order of increasing size.

Cl^- K^+ Ca^{2+} S^{2-}

(Smallest) _____< _____< _____< _____ (Largest)

12. a) Provide the Lewis structure, the name of the molecular geometry and the bond angle(s) around the central atom for the following molecules or ions.

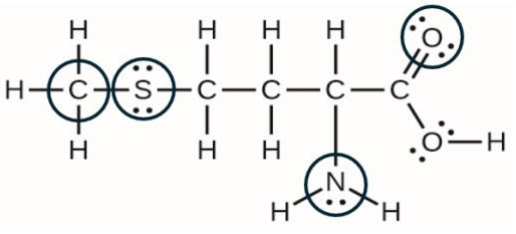
Formula	SO_2	BrF_5
Lewis Structure		
Geometry		
Bond angle(s) (°)		

b) Circle ALL molecules or ions that are non-polar.

SO_2 BF_4^- BrF_5

c) Provide <u>the most important</u> resonance structure of the SO_3^{2-} ion. Draw all lone-pair electrons and write the formal charge of atoms in the structure.

13. a) Methionine, $CH_3SCH_2CH_2CH(NH_2)CO_2H$, is an amino acid found in proteins. The Lewis structure of this compound is shown below. Give the hybridizations of the atoms that are circled.



Hybridization of circled C : _____

Hybridization of circled S : _____

Hybridization of circled N : _____

Hybridization of circled O : _____

b) What hybridization would you expect for the central atoms (in bold) in each of the following?

i. $H_2\mathbf{C}O$ _____

ii. $\mathbf{B}H_4^-$ _____

iii. $\mathbf{Xe}OF_4$ _____

c) The bonds between hydrogen and oxygen atoms within a water molecule can be characterized as :

a. hydrogen bonds
b. London dispersion forces
c. intermolecular forces
d. intramolecular bond
e. dispersion forces

d) The molecules in a sample of solid SO_2 are attracted to each other by a combination of:

a. London dispersion forces and H-bonding
b. H-bonding and ionic bonding
c. covalent bonding and dipole-dipole interactions
d. London dispersion forces and dipole-dipole interactions
e. none of these combinations

e) Make a sketch to show the hydrogen bonding between two methanol molecules (CH_3OH).

f) Arrange the following in order of highest to lowest boiling point temperature: CH_4, CH_3CH_3, CH_3CH_2Cl, CH_3CH_2OH.

i. CH_3CH_2OH, CH_3CH_2Cl, CH_3CH_3, CH_4

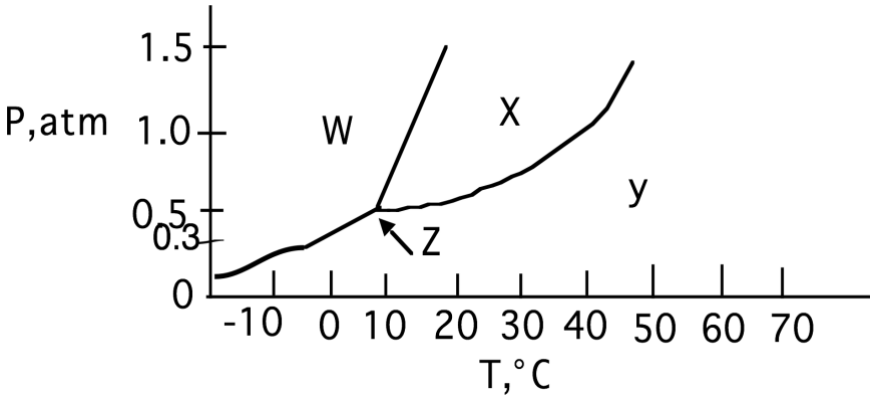
ii. CH_4, CH_3CH_3, CH_3CH_2Cl, CH_3CH_2OH

iii. CH_3CH_2Cl, CH_3CH_2OH, CH_3CH_3, CH_4

iv. CH_3CH_2Cl, CH_3CH_2OH, CH_4, CH_3CH_3

v. CH_3CH_2Cl, CH_3CH_3, CH_3CH_2OH, CH_4

14. For this phase diagram of compound A:



a) Which letter represents the following:

i) Gas phase _____
ii) Liquid _____
iii) Solid _____
iv) Triple Point _____

b) What is the normal boiling point of A?

i) 20°C ii) 40°C iii) 65°C iv) 80°C

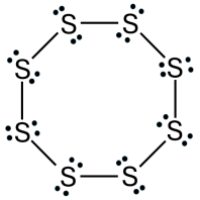
c) What is the normal melting point of A?

i) 12°C ii) 40°C iii) 80°C

d) What is the process called when we move:

i. from X to W _____
ii. from W to Y _____

15. a) The most common elemental form of sulfur is octasulfur (S_8). Combustion of this form of sulfur with oxygen produces sulfur dioxide (SO_2). Using the given bond energies, calculate the bond energy (in kJ/mol) for one sulfur-sulfur single bond in S_8 molecules.



$+ \ 8\ O_2(g) \rightarrow 8\ SO_2(g)$ $\Delta H = -2374$ kJ/mol of S_8

Bond	Bond energy (kJ/mol)
O=O	495
S=O	461

b) When 5.21 g of potassium nitrate (KNO_3) is dissolved in a coffee cup calorimeter containing 200.00 mL of water, the temperature of the solution decreased from 24.80 °C to 22.73 °C. Calculate the enthalpy of dissolution of KNO_3 (in kJ/mol). The specific heat of the solution is 4.184 J/g·°C and assume the density of the solution is 1.00 g/mL.

15. c) Using the equations below,

(i) $2\ NO_2(g) \rightarrow N_2(g) + 2\ O_2(g)$ $\Delta H_{rxn} =$ -66.40 kJ

(ii) $N_2(g) + 3\ O2(g) + H_2(g) \rightarrow 2\ HNO3(l)$ $\Delta H_{rxn} =$ -347.12 kJ

(iii) $2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$ $\Delta H_{rxn} =$ -483.64 kJ

(iv) $H_2O(g) \rightarrow H_2O(l)$ $\Delta H_{rxn} =$ -40.66 kJ

Calculate the ΔH for the following reaction:

$$4\ HNO_3(l) \rightarrow 2\ H_2O(l) + 4\ NO_2(g) + O_2(g)$$

Periodic Table of the Elements

	1A												3A	4A	5A	6A	7A	8A
1	1 H 1.008	2A																2 He 4.003
2	3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31	3B	4B	5B	6B	7B	8B	9B	10B	1B	2B	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc 98.00	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.3	57 La* 138.9	72 Hf 178.5	73 Ta 181.0	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po 209.0	85 At 210.0	86 Rn 222.0
7	87 Fr 223.0	88 Ra 226.0	89 Ac[a] 227.0	104 Rf 261.0	105 Db 262.0	106 Sg 263.0	107 Bh 262.0	108 Hs 265.0	109 Mt 266.0	110 Uun 269.0	111 Uuu 272.0	112 Uub 277.0						

= metalloid

*Lanthanides	58 Ce 140	59 Pr 141	60 Nd 144	61 Pm 145	62 Sm 150	63 Eu 152	64 Gd 157	65 Tb 159	66 Dy 163	67 Ho 165	68 Er 167	69 Tm 169	70 Yb 173	71 Lu 175
[a]Actinides	90 Th 232	91 Pa 231	92 U 238	93 Np 237.1	94 Pu 244	95 Am 243	96 Cm 247	97 Bk 247	98 Cf 251	99 Es 252	100 Fm 257	101 Md 258	102 No 259	103 Lr 260

Miscellaneous Constants and Conversions

Avogadro's number = 6.022×10^{23} mol^{-1}

Gas constant R = 0.08206 $L \cdot atm \cdot mol^{-1} \cdot K^{-1}$ = 8.314 $L \cdot kPa \cdot mol^{-1} \cdot K^{-1}$ = 8.314 $J \cdot mol^{-1} \cdot K^{-1}$

Standard pressure = 1.000 atm = 760.0 mm Hg = 101.3 kPa = 760.0 torr = 14.69 psi

Bohr constant (B) = 2.178×10^{-18} J

Rydberg constant (R_H) = 1.0974×10^{7} m^{-1}

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

Speed of light (c) = 2.998×10^{8} m s^{-1} 0°C = 273.15 K 1 J = kg m^2 s^{-2}

PRACTICE FINAL EXAMINATION 202-SN1-RE — ANSWER KEY

1. a) Row 1: Isotope: $^{80}_{35}Br^-$ | Protons: 35 | Electrons: 36 | Neutrons: 45 | Mass Number: 80 | Net Charge: 1−
   Row 2: Isotope: $^{226}_{88}Ra^{2+}$ | Protons: 88 | Electrons: 86 | Neutrons: 138 | Mass Number: 226 | Net Charge: 2+
   b) Identity of element: Lead (Pb) 207 82 Pb4+
   Number of neutrons: 125

2. • Group 1 element that is not a metal: H
   • All nonmetals except: Sb
   • Horizontal rows: Periods
   • Reaction behavior: metal (loses electrons), nonmetal (gains electrons)
   • Unreactive gaseous group: 8A
   • Element classified as a metalloid: germanium

3. a) i) Mg_3P_2 : Magnesium phosphide
   ii) SCl_6 : Sulfur hexachloride
   iii) HClO(aq) : Hypochlorous acid
   iv) NaH_2PO_4 : Sodium dihydrogen phosphate
   v) $FeSO_4 \cdot 2H_2O$: Iron(II) sulfate dihydrate

   b) i) Mercury(I) chloride : Hg_2Cl_2
   ii) Aluminum nitride : AlN
   iii) Sulfurous acid : H_2SO_3(aq)
   iv) Disulfur dichloride : S_2Cl_2
   v) Sodium dichromate : $Na_2Cr_2O_7$

4. a) Na = 43.38%, C = 11.33%, O = 45.29%
   b) Empirical Formula: $Mg_3Si_2H_3O_8$
   Molecular Formula: $Mg_6Si_4H_6O_{16}$

5. a) Balanced Equations:
   • Precipitation: $2\ K_3PO_4(aq) + 3\ MnCl_2(aq) \rightarrow Mn_3(PO_4)_2(s) + 6\ KCl(aq)$
   • Oxidation-Reduction: $C_5H_{10}O_5(s) + 5\ O_2(g) \rightarrow 5\ CO_2(g) + 5\ H_2O(l)$
   • Acid-Base: $HNO_3(aq) + NaHCO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l) + CO_2(g)$
   b) Strong Electrolyte: NH_4NO_3
   c) Chloride Ion Concentration Ranking: Solution C < Solution B < Solution A

6. a) Molar Mass = 146 g/mol
   b) i. Theoretical yield of Ti = 1.11×10^3 g (or 1108 g)
   ii. Percent yield = 88.9%

7. a) i. CME: $2\ AgNO_3(aq) + Na_2S(aq) \rightarrow Ag_2S(s) + 2\ NaNO_3(aq)$
   ii. CIE: $2\ Ag^+(aq) + 2\ NO_3^-(aq) + 2\ Na^+(aq) + S^{2-}(aq) \rightarrow Ag_2S(s) + 2\ Na^+(aq) + 2\ NO_3^-(aq)$
   iii. NIE: $2\ Ag^+(aq) + S^{2-}(aq) \rightarrow Ag_2S(s)$
   iv. Spectator Ions: Na^+ and NO_3^-

   b) Mass of Na_2S required = 0.177 g
   c) Mass of precipitate (Ag_2S) = 0.562 g

8. a) Solubility: Na_2SO_4 (Soluble) | KOH (Soluble) | $BaCO_3$ (Insoluble) | $AgNO_3$ (Soluble)

b) i. Oxidizing agent: MnO_4^- | Element reduced: Mn

Reducing agent: $C_2O_4^{2-}$ | Element oxidized: C

ii. Balanced Redox Equation: $2\ MnO_4^-(aq) + 3\ C_2O_4^{2-}(aq) + 2\ H_2O(l) \rightarrow 2\ MnO_2(s) + 6\ CO_3^{2-}(aq) + 4\ H^+(aq)$

9. a) Longest Wavelength: n = 8 to n = 6

b) Wavelength (λ): 5.40×10^{-8} m (54.0 nm)

c) Energy: 9.51×10^{-15} J; Frequency: 1.43×10^{19} Hz (or s^{-1})

10. a) Br: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$; Mg^{2+}: $1s^2\ 2s^2\ 2p^6$

b) Cu^+: [Ar] $3d^{10}$; Se^{2-}: [Ar] $4s^2\ 3d^{10}\ 4p^6$ (or [Kr])

c) Ion: Rb^+

d) Correctness: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 4d^{10}\ 4p^6\ 5s^1$ (incorrect) | [Xe] $6s^2\ 4f^{10}$ (correct) | $1s^2\ 2s^2\ 2p^6\ 3s^3$ (incorrect) | [Ar] $4s^1\ 3d^{10}$ (correct)

e) Electron capacity: n = 3, ℓ = 1 → 6 electrons | n = 5, mℓ = 3 → 4 electrons

f) Allowed checks: (n=3, ℓ=2, mℓ=-3, ms =-½) : Not allowed | (n=0, ℓ=0, mℓ=0, ms = +½) : Not allowed | (n=3, ℓ=1, mℓ=-1, ms =+½) : Allowed | (n=4, ℓ=-3, mℓ=-1, ms = +1) : Not allowed

11. a) Lowest IE: Cs

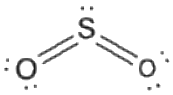
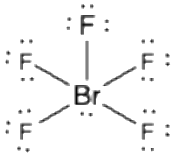
b) Down a group: Electron affinity (decreases) | Atomic radius (increases)

c) More polar bond: C-N | Si-O | N-H

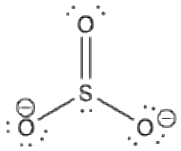
d) Partial charges: δ+ H - F δ- | δ+ S - O δ- | δ+ P - Cl δ-

e) Ionic Size: $Ca^{2+} < K^+ < Cl^- < S^{2-}$

12. a) SO_2: Geometry = Bent | Bond angle = < 120°; BrF_5: Geometry = Square pyramidal | Bond angle = ~90°

Formula	SO_2	BrF_5
Lewis Structure		

b) Non-polar species: BF_4

c) 

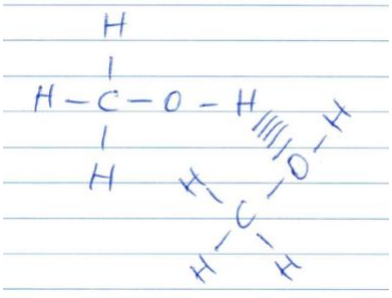
13. a) Circled C ($-CH_3$): sp^3 | Circled S: sp^3 | Circled N: sp^3 | Circled O (=O): sp^2

b) Central atoms: C in H_2CO (sp^2) | B in BH_4^- (sp^3) | Xe in $XeOF_4$ (d^2sp^3)

c) d. intramolecular bond

d) d. London dispersion forces and dipole-dipole interactions

e) Sketch showing O-H···:O-H hydrogen bonding between two CH_3OH molecules.



f) Boiling point order: $CH_3CH_2OH > CH_3CH_2Cl > CH_3CH_3 > CH_4$

14. a) Phases: Gas = Y | Liquid = X | Solid = W | Triple Point = Z
    b) Normal boiling point: 40 °C
    c) Normal melting point: 12 °C
    d) Processes: X to W (Freezing) | W to Y (Sublimation)

15. a) S-S Bond energy: 130.3 kJ/mol
    b) Enthalpy of dissolution (KNO_3): +34.5 kJ/mol
    c) ΔH for net reaction: +262.08 kJ