THE COPPERBELT UNIVERSITY
SCHOOL OF MATHEMATICS AND NATURAL SCIENCES
DEPARTMENT OF CHEMISTRY

CHEMISTRY (CH 110/ FO 130) TUTORIAL SHEET 1 TERM ONE YEAR 2024

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1. Explain the difference between:
   (i) law and theory (ii) hypothesis and theory
   (iii) precision and accuracy (iv) homogeneous and heterogeneous mixture
2. Many atomic dimensions are expressed in angstroms. What is the angstrom equal to in terms of the SI Units nanometer (nm) and picometer (pm)
3. Convert each of following:

   (a) 1.5 gcm^{-3} to gdm^{-3} (b) 100 cm^3 to dm^3 (c) 0.36 m to Å (d) 16 cm^3 to m^3

4. Define each of the following terms:
   (a) Quantitative measurement (b) Precision
   (c) Qualitative measurement (d) Systematic error
5. The density of air at ordinary atmospheric pressure and 25 ^oC is 1.19 g/L. What is the mass of air in kilograms in a room that measures 3.0 m x 4.0 m x 5.0 m?
6. Convert the following temperatures to Kelvin:
   (a) The freezing point of water 0 ^oC.
   (b) Melting temperature of Neon -248 ^oC.
   (c) Body temperature of 98.6^oF.
7. Cesium atoms are the largest naturally occurring atoms. Each has a radius of 2.62 Å. How many atoms of cesium would have to be laid side by side to a row of cesium atoms 1.0 mm long, assuming the atoms are spherical?
8. Perform the following mathematical operations and express each result in to the correct number of significant figures.
   (a) $(9.04 – 8.23 + 21.954 + 81.0) \div 3.1416$
   (b) $0.1654 + 2.07 – 2.114$
   (c) $8.27(4.987 – 4.962)$
   (d) $485 \div 9.231$
9. State the postulates of John Dalton’s theory of the atom.

10. State the law of constant composition.
11. Indicate the number of protons, neutrons and electrons in each of the following species.
    (a) Ca^{2+} (b) Cu^{2+} (c) P^{3-} (d) Cl^{-} (e) Ni^{2+} (f) Fe (g) Fr^{+}
12. If an element consists of 47.50 % atoms with a mass of 120.052 amu each and 52.50 % atoms with a mass of 122.855 amu each, what is the atomic mass of the element?
13. Provide the systematic name of each of the following compounds:
    (a) $AlCl_3$ (b) $Cu(NO_3)_2$ (c) $K_2Cr_2O_7$ (d) IF_5 (e) P_2O_5 (f) HClO (g) H_3PO_4
14. Provide the chemical formula of each of the following compounds:
    (a) Potassium chromate (b) Iron (III) hydroxide (c) hydrogen cyanide
    (d) Potassium permanganate (e) tetraphosphorus hexasulphide (f) Iodic acid

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THE COPPERBELT UNIVERSITY

DEPARTMENT OF CHEMISTRY

CH110: TUTORIAL SHEET 2

STOICHIOMETRY

Term I (2023-2024)

1. (i) Calculate the number of moles that contain 4.50 × 10 atoms of zinc. (ii) What amount of H is there in 5.6 mol of (hydrazine)?

   (iii) Find the amount of penicillin, , that contains 0.10 mol C.

   (iv) What mass of contains 5.10 × 10 molecules of ?

2. Ethylene glycol used in automobile antifreeze, contains only carbon, hydrogen and oxygen. Combustion analysis of a 23.46 mg sample yields 20.42 mg of H_2O and 33.27 mg CO_2. What is the empirical formula and molecular formula of the ethylene glycol if it has a molecular mass of 62.0 *amu.*

3. Barium peroxide, BaO_2, reacts with hydrochloric acid, HCl, to give hydrogen peroxide, H_2O_2, and barium chloride, $BaCl_2$. In a particular experiment, it was planned to treat 1.45 g of barium peroxide with 25.5 mL of hydrochloric acid solution containing 0.0277 g of HCl per mL

   (a) Write a balanced chemical equation for the reaction. (b) Which is the limiting reactant

   (c) What mass in grams of hydrogen peroxide (H_2O_2) should result?

   (d) How many grams of the reactant in excess will be left over?

   (e) If 1.34 g of $BaCl_2$ were produced, calculate the percentage yield for $BaCl_2$

4. Silver (Ag) is a precious metal used mainly in jewelry. What is the mass (in grams) of one Ag atom?

5. How many hydrogen atoms are present in 43.8 g of urea [$(NH_2)_2CO$], which is used as a fertilizer, in animal feed, and in the manufacture of polymers? The molar mass of urea is
   60.06 g.

6. Phosphoric acid (H_3PO_4) is a colorless, syrupy liquid used in detergents, fertilizers, toothpastes, and in carbonated beverages for a "tangy" flavor. Calculate the percent composition by mass of H, P, and O in this compound.

7. Ascorbic acid (vitamin C) contains 40.92% carbon, 4.58% hydrogen and 54.50% oxygen by mass. What is the empirical formula of ascorbic?

8. Iron (Fe), the main component of steel, is the most important metal in industrial society. How many Fe atoms are in 95.8 g of Fe?

9. Ammonium carbonate is a white solid that decomposes with warming. Among its many uses, it is a component of baking powder, fire extinguishers, and smelling salts. How many formula units are in 41.6 g of ammonium carbonate?

10. In mammals, lactose (milk sugar) is metabolized to glucose ($C_6H_{12}O_6$), the key nutrient for generating chemical potential energy.
(a) What is the mass percent of each element in glucose?
(b) How many grams of carbon are in 16.55 g of glucose?

11. Write a balanced equation for the reaction between aqueous strontium chloride and aqueous lithium phosphate to form solid strontium phosphate and aqueous lithium chloride.

12. Aspirin, acetylsalicylic acid ($C_9H_8O_4$), is the most commonly used pain reliever in the world. It is produced by the reaction of salicylic acid ($C_7H_6O_3$) and acetic anhydride ($C_4H_6O_3$) according to the following equation:

$$C\,H\,O \quad l \quad + C\,H\,O \quad l \quad \rightarrow C\,H\,O \quad s \quad + HC\,H\,O\,l$$

In a certain aspirin synthesis, 104.8 g of salicylic acid and 110.9 g of acetic anhydride are combined. Calculate the percent yield of the reaction if 105.6 g of aspirin are produced.

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Tutorial 2

1. (i) $\left[\frac{1 \text{ mole Zinc}}{6.02\times10^{23}\text{ atoms}}\right] \times 4.50\times10^{24}\text{ atoms}$

$= \underline{7.47 \text{ moles}}$

(ii) Amount $= 4 \times 5.6\text{mol}$

$= \underline{22 \text{ moles}}$

(iii) 1 mol $C_{16}H_{18}N_2O_4S$ —— 16 mol of C

x —— 0.10

Thus $x = \frac{0.10\times1}{16} = \underline{0.0063 \text{ mol penicillin}}$

(iv) $\left[\frac{1\text{ mol}}{6.02\times10^{23}\text{ atoms}}\right] \times 5.10\times10^{24}$

$= 8.47 \text{ mol}$

MASS $=$ mol $\times$ mm

$= 8.47\text{mol} \times 44.01\text{ g/mol}$

$= \underline{371.75\text{ g}}$

2. Combustion of ethylene glycol in excess oxygen

$$C_xH_yO_z + O_2 \longrightarrow CO_2 + H_2O$$

Data

$C_xH_yO_z$	CO_2	H_2O
mass = 23.46 mg	mass = 33.27 mg	mass = 20.42 mg
molar mass = 62.0 g/mol	mm = 44.01 g/mol	mm = 18.0 g/mol

→ for C and H we know only their molar masses

C = 12.0 g/mol

H = 1.0 g/mol

Note : The fractional composition of an element in a compound can be obtained from ratio of the molar mass of an element to the molar mass of compound containing it

$$\frac{M_C}{M_{CO_2}} = \frac{m_C}{m_{CO_2}} \quad \text{and} \quad \frac{M_H}{M_{H_2O}} = \frac{m_H}{m_{H_2O}}$$

$$m_C = \frac{M_C \times m_{CO_2}}{M_{CO_2}} = \frac{12.0\ \text{amu} \times 33.27\ \text{mg}}{44.0\ \text{amu}} = 9.07\ \text{mg}$$

$$m_H = \frac{M_H \times m_{H_2O}}{M_{H_2O}} = \frac{1.0\ \text{amu} \times 20.42\ \text{mg}}{18.0\ \text{amu}} = 2.15\ \text{mg}$$

3. (a) $BaO_2(s) + 2HCl(aq) \longrightarrow H_2O_2(aq) + BaCl_2(aq)$

(b) finding the limiting reactant

- mass of HCl that reacts : $(0.0277 g/mL) \times (25.5 mL)$
$= \underline{0.70635 g}$.

mol HCl available $= 0.70635 g / 36.5 g/mol = 0.01937$ mol

mol BaO_2 available $= 1.45 g / 169.3 g/mol = 0.008565$ mol

→ for limiting reagent: divide the moles of each reactant by their respective coefficient in balanced equation

BaO_2 : 0.008565 mol / 1 = 0.008565

HCl : 0.01937 mol / 2 = 0.009685

Since $\underline{BaO_2}$ is less it is the limiting reagent.

(c) 0.008565 mol × 34 g/mol = $\underline{0.291 g}$

(d) g of the reactant in excess

- moles of HCl used = 0.008565 × 2 = 0.01713 mol

- moles of HCl remaining = 0.01937 mol − 0.01713 mol = 0.00224 mol

mass of HCl = 0.00224 mol × 36.46 g/mol = $\underline{0.0817 g}$

(e) $\dfrac{1.34 g}{1.78 g} \times 100\% = 75.13\%$

$BaCl_2$ mass
= 0.008565 mol × 208.23 g/mol
= $\underline{1.78 g}$

Since the mass of $C_xH_yO_z$ is the sum of the individual masses of the atoms we have:

$$m_{C_xH_yO_z} = m_C + m_H + m_O$$

$$m_O = m_{C_xH_yO_z} - m_C + m_H$$
$$= 23.46 - (9.07 + 2.15)$$
$$= 12.24 \text{ mg}$$

→ Converting the mass from mg to g we have

	C	H	O
m =	0.00907 g	0.00215 g	0.01224 g
n =	$\frac{0.00907\,g}{12.0\,g/mol} = 7.55 \times 10^{-4}$ mol	$\frac{0.00215\,g}{1.0\,g/mol} = 2.13 \times 10^{-3}$ mol	$\frac{0.01224\,g}{16\,g/mol} = 7.65 \times 10^{-4}$ mol
dividing moles by the smallest	$\frac{7.55 \times 10^{-4}}{7.55 \times 10^{-4}}$	$\frac{2.13 \times 10^{-3}}{7.55 \times 10^{-4}}$	$\frac{7.65 \times 10^{-4}}{7.55 \times 10^{-4}}$
	1	3	1

Empirical formula = CH_3O with 31.0 amu

$$\frac{\text{mm of molecular formula}}{\text{mm of empirical formula}} = \frac{62.0 \text{ amu}}{31.0 \text{ amu}} = 2$$

Thus molecular formula is given by $(CH_3O)_2 = C_2H_6O_2$

4. Since 1 mole of Ag atom = 6.022×10^{23} Ag atoms

1 mole of Ag = 107.9 g

mass of one Ag atom $= \left[\frac{1 \cancel{\text{mol Ag}}}{6.022 \times 10^{23} \cancel{\text{atoms}}}\right]\left[\frac{107.9\,g}{1 \cancel{\text{mol Ag}}}\right] \times 1 \cancel{\text{atom}}\ Ag$

$= \underline{1.792 \times 10^{-22}\,g}$

5. Data

- 43.8 g of $(NH_2)_2CO$
- 1 mol of $(NH_2)_2CO$ = 60.06 g
- 1 mol of $(NH_2)_2CO$ = 4 mol of H
- 1 mol of H = 6.022×10^{23} H atoms

$$43.8\,\cancel{g\,(NH_2)_2CO} \times \left[\frac{1\,\cancel{mol\,(NH_2)_2CO}}{60.06\,\cancel{g\,(NH_2)_2CO}}\right]\left[\frac{4\,\cancel{mol\,H}}{1\,\cancel{mol\,(NH_2)_2CO}}\right]\left[\frac{6.022 \times 10^{23}\,H\ atoms}{1\,\cancel{mol\,H}}\right]$$

$= \underline{1.76 \times 10^{24}\ H\ atoms}$

(6). The molar mass of H_3PO_4 is 97.99g Thus the % by mass is given by $\frac{\text{mass of atom}}{\text{mass of Compound}} \times 100\%$

$$\% H = \frac{3(1.008g)\,H}{97.99g\ H_3PO_4} \times 100\% = \underline{\underline{3.086\%}}$$

$$\% P = \frac{30.97g\ P}{97.99g\ H_3PO_4} \times 100\% = \underline{\underline{31.61}}\%$$

$$\% O = \frac{4(16.00g)\,O}{97.99g\ H_3PO_4} \times 100\% = \underline{\underline{65.31}}\%$$

(7). We first get the number of moles.

$$C = \frac{40.92g}{12.01\,g/mol} = 3.407\,mol \quad H = \frac{4.58g}{1.008\,g/mol} = 4.54\,mol$$

$$O = \frac{54.50g}{16.00\,g/mol} = 3.406\,mol$$

Dividing with the smallest number of moles 3.406 mol

$$\begin{array}{ccc} C & H & O \\ \frac{3.407}{3.406} & \frac{4.54}{3.406} & \frac{3.406}{3.406} \\ = 1 & = 1.33 & = 1 \end{array} \Rightarrow CH_{1.33}O$$

Converting 1.33 into an integer by multiplying by 3

$= 3.99 = 4$

$\Rightarrow \underline{\underline{C_3H_4O_3}}$ empirical formula

(8.) - find moles of Fe then convert them into atoms.

$$\text{moles of Fe} = \frac{95.8\ g}{55.85\ g/mol} = 1.72\ mol\ Fe$$

$$\text{atoms of Fe} = 1.72\ \cancel{mol\ Fe}\left[\frac{6.022\times10^{23}\ \text{atoms Fe}}{1\ \cancel{mol\ Fe}}\right]$$

$$= \underline{1.04\times10^{24}\ \text{atoms Fe}}$$

(9.) → first find molar mass of $(NH_4)_2CO_3$

$= (2\times14.01\ g/mol\ N) + (8\times1.008\ g/mol\ H) + 12.01\ g/mol\ C$

$+ (3\times16.00\ g/mol\ O)$

$= 96.09\ g/mol\ (NH_4)_2CO_3$

→ moles of $(NH_4)_2CO_3 = \frac{41.6\ g}{96.09\ g/mol} = 0.433\ mol\ (NH_4)_2CO_3$

$$\text{formula units} = 0.433\ \cancel{mol}\left[\frac{6.022\times10^{23}\ \text{formula units}\ (NH_4)_2CO_3}{1\ \cancel{mol}}\right]$$

$$= \underline{2.61\times10^{23}\ \text{formula units}\ (NH_4)_2CO_3}$$

10. (a) First we find molar mass of $C_6H_{12}O_6$

$$= (6 \times 12.01) + (12 \times 1.008) + (6 \times 16.00)$$

$$= 180.16\ g/mol\ C_6H_{12}O_6$$

for:

C mass $= 6\ mol \times 12.01\ g/mol = 72.06\ g\ C$

mass % of C $= \frac{72.06\ g}{180.16\ g} \times 100\% = \underline{40\ mass\ \%\ C}$

H mass $= 12\ mol \times 1.008\ g/mol = 12.09\ g\ H$

mass % of H $= \frac{12.09\ g}{180.16\ g} \times 100\% = \underline{6.714\ mass\ \%\ H}$

O mass $= 6\ mol \times 16.00\ g/mol = 96\ g\ O$

mass % of O $= \frac{96\ g}{180.16\ g} \times 100\% = \underline{53.29\ mass\ \%\ O}$

(b) mass (g) of Carbon $\frac{40\%}{100\%} = 0.4000\ g$ mass

∴ mass of C = mass of glucose × mass of fraction of C

$$16.55\ \cancel{g\ glucose} \times \frac{0.4000\ g\ C}{1\ \cancel{g\ glucose}} = \underline{6.620\ g\ C}$$

11. $3SrCl_{2(aq)} + 2Li_3PO_{4(aq)} \rightarrow Sr_3(PO_4)_{2(s)} + 6LiCl_{(aq)}$

12. Klo find the number of moles for the two reactants.

$$C_7H_6O_{3(l)} + C_4H_6O_{3(l)} \rightarrow C_9H_8O_{4(s)} + HC_2H_3O_{2(l)}$$

$\rightarrow C_7H_6O_3 = \dfrac{104.8g}{138.12g/mol} = 0.7588 \text{ mol } C_7H_6O_3$

$\rightarrow C_4H_6O_3 = \dfrac{110.9g}{102.09g/mol} = 1.086 \text{ mol } C_4H_6O_3$

from the balanced equation the ratio is 1:1, since $C_7H_6O_3$ is less it is a limiting reagent. Thus theoretical yield is 0.7588 mol

mass = 0.7588 mol × 180.15g/mol = 136.7g $C_9H_8O_4$

Theoretical yield = 136.7g

Actual yield = 105.6g

% yield = $\dfrac{105.6g}{136.7g} \times 100\%$

= 77.25% yield

A gossip machine at rest

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THE COPPERBELT UNIVERSITY

DEPARTMENT OF CHEMISTRY

CH110: TUTORIAL SHEET 3

REACTIONS IN AQUEOUS SOLUTIONS

Term I (2023-2024)

1. The compounds K_2CO_3, Na_2CO_3, KCl and NaCl are soluble in water, but $CaCO_3$ is not.

   Given the following molecular equation:

   $$CaCl_2(aq) + Na_2CO_3(aq) \rightarrow CaCO_3(s) + 2NaCl(aq)$$

   (a) Write *molecular equation* for the reaction between solutions of $CaCl_2$ andK_2CO_3

   (b) Write the *ionic equation* for the reaction between solutions of $CaCl_2$ andK_2CO_3

   (c) What is meant by the term *spectator ion*

   (d) Write down the spectator ion(s)for this reaction if any

   (e) Write the *net ionic equation* for the reaction between solutions of $CaCl_2$ and K_2CO_3

2. What volume of 16 *M* sulfuric acid must be used to prepare 1.5 L of a 0.10-*M* H_2SO_4 solution?

3. Assign oxidation numbers to all the elements in the following compounds and ion:
   (a) Na_2O, (b) HNO_2, (c) $Cr_2O_7^{-2}$

4. Define molarity. Calculate the molarity of a solution prepared by dissolving 1.56 g of gaseous HCl in enough water to make 26.8 mL of solution.

5. Glycine (H_2NCH_2COOH) is the simplest amino acid. What is the molarity of an aqueous solution that contains 0.715 mol of glycine in 495 mL?

6. Specialized cells in the stomach release HCl to aid digestion. If they release too much, the excess can be neutralized with an antacid to avoid discomfort. A common antacid contains magnesium hydroxide, $Mg(OH)_2$, which reacts with the acid to form water and magnesium chloride solution. As a government chemist testing commercial antacids, you use 0.10 *M* HCl to simulate the acid concentration in the stomach. How many liters of “stomach acid” react with a tablet containing 0.10 g of $Mg(OH)_2$?

7. Predict whether a reaction occurs when each of the following pairs of solutions are mixed. If a reaction does occur, write balanced molecular, total ionic, and net ionic equations, and identify the spectator ions.
   (a) Potassium fluoride(aq) + strontium nitrate(aq) →
   (b) Ammonium perchlorate(aq) + sodium bromide(aq) →

8. You perform an acid-base titration to standardize an HCl solution by placing 50.00 mL of HCl in a flask with a few drops of indicator solution. You put 0.1524 M NaOH into the buret, and the initial reading is 0.55 mL. At the end point, the buret reading is 33.87 mL. What is the concentration of the HCl solution?

9. Identify the oxidizing agent and reducing agent in each of the following:
   (a) $\mathrm{2Al\ s + 3H\ SO\ aq \rightarrow Al\ SO\qquad aq + 3H\ g}$
   (b) $\mathrm{PbO\ s + CO\ g \rightarrow Pb\ s + CO\ g}$
   (c) $\mathrm{2H\ g + O\quad g \rightarrow 2H\ O\ g}$

10. Permanganate ion is a strong oxidizing agent, and its deep purple color makes it useful as an indicator in redox titrations. It reacts in basic solution with the oxalate ion to form carbonate ion and solid manganese dioxide. Balance the skeleton ionic equation for the reaction between $NaMnO_4$ and $Na_2C_2O_4$ in basic solution:

$$\mathrm{MnO\quad aq + C\ O\qquad aq \rightarrow MnO\quad s + CO\quad aq}$$

11. Balance the redox reaction between dichromate ion and iodide ion to form chromium(III) ion and solid iodine, which occurs in acidic solution.

$$+ \qquad \rightarrow \qquad + \quad !$$

CH110: Gases Tutorial Sheet 4 2024

a) Write down the mathematical expressions of the following gas laws:
   i. Boyle's law
   ii. Charles' law

b) Consider a mixture of two gases, A, and B, confined in a closed vessel. A quantity of a third gas, C, is added to the same vessel at the same temperature. State whether each of the following will increase, decrease, or remain the same.

   i. The partial pressure of gas A
   ii. The total pressure in the vessel
   iii. The mole fraction of gas B

c) A diver releases a bubble of air which rises to the surface of the sea. The volume of the bubble increases from 4.0 cm^3 to 10.0 cm^3 during this process.
   i. Why does the bubble expand as it rises?
   ii. Just below the surface, the pressure acting on the bubble is one atmosphere. What is the pressure (in atmospheres) acting on it before it starts to rise?

   iii. Which gas law is being applied in your calculation in c(ii) above?

d) An aerosol spray can with a volume of 250 ml contains 2.30 g of propane gas (C_3H_8) as a propellant.
   i. If the can is at 23^0C, what is the pressure in the can?
   ii. What volume would the propane occupy at STP?
   iii. The can's label says that exposure to temperatures above 130^0F may cause the can to burst. What is the pressure in the can at this temperature?

e) When 1.77 g of a gas was stored in a 1.500 L flask at 17 ºC, it exerted a pressure of 508 Torr. What is the molar mass of the gas?

f) A container of volume 5 dm^3 was filled with 2.02 g of $H_{2(g)}$ and 71 g of $Cl_{2(g)}$. The mixture was then sparked leading to the formation of $HCl_{(g)}$ and then cooled to a temperature of 25 ºC. At the end of the reaction, determine the
   (i) Partial pressure
   (ii) Partial pressure of each component gas
   (iii) Total pressure in the container.
   (iv) Mole fraction of each gas.

g) Between helium and ammonia gas, which one would you expect to exhibit behaviour close to ideal gas behaviour at low pressure? Provide brief explanation for your answer.

h) State the Van der Waals equation and define each of the terms involved.

1. The internal energy of a system increased by 982 J when it absorbed 492 J of heat. Was work done by or on the system? How much work was done?

2. Aluminum metal reacts with chlorine with spectacular display of sparks:

   $2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)\ \Delta H = -1408\ kJ$

   How much heat (in kilojoules) is released on reaction of 5.00 g Al?

3. Given the thermochemical equation for photosynthesis,
   $6H_2O(l) + 6CO_2(g) \rightarrow C_6H_{12}O_6(s) + 6O_2(g)$ $\Delta H = +2803$ kJ/mol
   Calculate the solar energy required to produce 75.0 g of $C_6H_{12}O_6$

4. A 1.000 g sample of octane (C_8H_{18}) is burned in a bomb calorimeter containing 1200 grams of water at an initial temperature of 25.00ºC. After the reaction, the final temperature of the water is 33.20 ºC. The heat capacity of the calorimeter (also known as the "calorimeter constant") is 837 J/ºC. The specific heat of water is 4.184 J/g ºC. Calculate the heat of combustion of octane in kJ/mol.

5. The combustion of 1 mole of benzene, $C_6H_6(l)$, to produce $CO_2(g)$ and $H_2O\ (l)$ liberates 3271 kJ when the products are returned to $25\ ^0C$ and 1 atm. What is the standard heat of formation of $C_6H_6(l)$ expressed in $kJ\ mol^{-1}$. Standard enthalpies of formation of $CO_2(g)$ and $H_2O(l)$ are $-394\ kJ\ mol^{-1}$ and $-286\ kJ\ mol^{-1}$ respectively

6. Consider these reactions, where M represents a generic metal.

   (i) $2M(s) + 6HCl(aq)$ Cl→ $2M_3(aq) + 3$ $\Delta H_1 = -619.0\ kJ$ $H_2(g)$
   $HCl(g) \rightarrow HCl(aq)$ $\Delta H_2 = -74.8\ kJ$
   $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$ $\Delta H_3 = -1845.0$
   (ii) $MCl_3(s) \rightarrow MCl_3(aq$) $\Delta H_4 = -132.0\ kJ$

   (iii)   kJ (iv)

   Use the information above to determine the enthalpy of the following reaction

   $2M(s) + 3Cl_2(g) \rightarrow MCl_3(s)$ $\Delta H_{rxn} = ?$

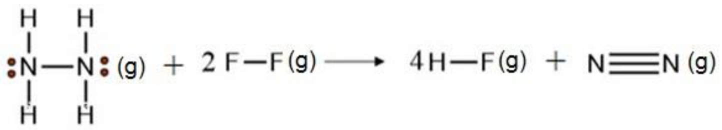
7. Calculate the enthalpy change for the following reaction
   $N_2O(g) + NO_2(g) \rightarrow 3NO(g)$

   Given

   (i) $N_2(g) + O_2(g) \rightarrow 2NO(g)$ $\Delta H^0 = +180.7\ kJ\ mol^{-1}$
   $2NO(g) + O_2(g) \rightarrow 2NO_2(\ g)$ $\Delta H^0 = -113.1\ kJ\ mol^{-1}$
   (ii) $2N_2O(g) \rightarrow 2N_2(g) + O_2(g\ )$ $\Delta H^0 = -163.2\ kJ\ mol^{-1}$
   (iii)

8. Use the provided bond energies to calculate the approximate enthalpy change for the following reaction. Relevant Lewis structures are provided.

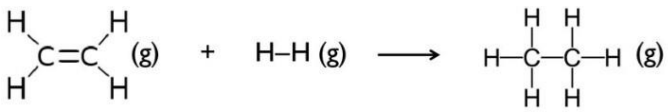
$$N_2H_4(g) \quad + \; 2F_2(g) \quad \rightarrow \quad 4HF(g) \quad + \quad N_2(g)$$



Bond	Bond Energy (kJ/mol)
H – H	432
N – H	391
C – N	305
C ≡ N	891
C – H	412
N ≡ N	941
H – F	565
N – N	160
C = C	612
F – F	154

9. The enthalpy change for the following reaction is –137 kJ. Using bond energies, estimate the C–C bond energy in $C_2H_6(g)$. Lewis structures for all reactants and products are shown.

$$C_2H_4(g) \quad + \quad H_2(g) \quad \rightarrow \quad C_2H_6(g)$$



10. The standard enthalpy of combustion for liquid acetone (C_3H_6O) is -1658 kJ mol^{-1}. Calculate the standard enthalpy of formation of liquid acetone (C_3H_6O), in kJ mol^{-1}.

For $H_2O(g), \Delta H_f^0 = -241.8$ kj mol^{-1}; for $CO_2(g), \Delta H_f^0 = -393.5$ kj mol^{-1}

11. Calculate the Enthalpy of reaction, ΔH^0_{rxn}, for the following reaction.

$$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$$

Given the following thermochemical reactions

(i)	$2CH_4(g) \rightarrow C_2H_4(g) + 2H_2(g)$	$\Delta H^0_{rxn} = +201.9$ kJ
(ii)	$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$	$\Delta H^0_{rxn} = -890.3$ kJ
(iii)	$C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$	$\Delta H^0_{rxn} = -137.0$ kJ
(iv)	$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$	$\Delta H^0_{rxn} = -285.8$ kJ

Tutorial sheet 5

Q1) $\Delta E = q + W$ ---- (i)

Data

$E = 982\,J$ $\quad q = 492\,J$

- Since q was absorbed, it will be positive

$\therefore$ Using equation (i) $\quad 982\,J = 492\,J + W$

$$982\,J - 492\,J = W$$

$$490 = W$$

$\therefore$ 490 J ~~was do~~ of work was done on the system

Q2) - For every 2 mol of the metal, aluminum the heat released is equal to 1408 KJ.

$$q = \frac{n\,\Delta H}{\text{Coefficient of Al}}$$

n = number of moles

ΔH = The enthalpy of the reaction

q = heat energy required.

$$q = \frac{5.00\,g}{26.815\,g/mol} \times \frac{-1408\,KJ}{2\,mol}$$

$$q = \cancel{5.00\,g} \; -130.47\,KJ \text{ of Al}$$

Q3) Molar Mass of $C_6H_{12}O_6 = 180.2 g/mol$

$$n = 75.0g \left[\frac{1 mol}{180.2g}\right] = 0.4162 \text{ mol of } C_6H_{12}O_6$$

$$\text{Solar energy required} = 0.4162 \cancel{mol} \times 2803 \frac{KJ}{\cancel{mol}}$$
$$= 1,166.61 \text{ KJ solar energy}$$

Q4) <u>Data</u>

- mass of octane $M = 1.00 g$
- mass of water $m_{water} = 1200g$
- Initial temp of water $T_1 = 25.00°C$
- Heat capacity of the calorimeter $C_{calorimeter} = 837 J/°C$
- Specific heat capacity of water $C_{water} = 4.814 J/(g°C)$

→ mole expression of Octane

$$n = \frac{M}{M_{octane}} = \frac{1.00g}{114.23 g/mol} = \frac{1.000}{114.23} mol$$

→ Heat gained by the calorimeter

$$Q_{calorimeter} = C_{calorimeter}(T_2 - T_1)$$
$$= 837 J/°C (33.20°C - 25.00°C)$$
$$= 6863.4 J$$

Q5 solution

- Enthalpy of combustion of 1 mol of benzene

$C_6H_{6(l)} + \frac{15}{2}O_2 \rightarrow 6CO_{2(g)} + 3H_2O_{(l)}; \Delta_f H^\circ = -3271\ kJ/mol$ (i)

- Enthalpy of formation of 1mol of $CO_2(g)$

$C + O_2 \rightarrow CO_2 : \Delta F H^\circ = -394\ kJ/mol$ --(ii)

- Enthalpy of formation of 1 mol of $H_2O_{(l)}$:

$H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(l)}; \Delta_c H^\circ = -286\ kJ/mol$ (iii)

formation reaction of benzene is given by

$6C_{(graphite)} + 3H_2(g) \longrightarrow C_6H_{6(l)}\ \Delta_f H^\circ = ?$ ---(iv)

With reference to equation (iv) we multiply equation (ii) by 6 and equation (iii) by 3 giving us the following

$\Rightarrow 6C + 6O_{2(g)} \longrightarrow 6CO_{2(g)} : \Delta_F H^\circ = -2,364\ kJ/mol$

$\Rightarrow 3H_{2(g)} + \frac{3}{2}O_2(g) \longrightarrow 3H_2O_{(l)} : \Delta_F H^\circ = -858\ kJ/mol$

Summing up the above equations

$6C_{(graphite)} + 3H_{2(g)} + \frac{15}{2}O_{2(g)} \rightarrow 6CO_{2}(g) + 3H_2O_{(L)}\ \Delta F H^\circ = -3222\ kJ/mol$

----(v)

Q6.) Overal reaction

$$2M_{(s)} + 3Cl_{2(g)} \longrightarrow 2MCl_3 \qquad \Delta H_{rxn}$$

- M represents a generic metal

steps

(i) $2M_{(s)} + 6HCl_{(aq)} \rightarrow 2MCl_{3(aq)} + 3H_{2(g)}$ $\Delta H_1 = -619.0\,kJ$

(ii) $HCl_{(g)} \longrightarrow HCl_{(aq)}$ $6 \times \Delta H_2 = -74.8\,kJ$

(iii) $H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$ $3 \times \Delta H_3 = -1845.0\,kJ$

(iv) $MCl_{3(s)} \longrightarrow MCl_{3(aq)}$ $2 \times \Delta H_4 = -132.0\,kJ$

- We use Hess's law

$$\Delta H_{net} = \sum \Delta H_r = \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4$$

- step (iv) has the product on the reactant side thus reverse and balance
- Whatever you do when balancing the equation affects the ΔH_r of that reaction step.

(iv) $2MCl_{3(aq)} \longrightarrow 2MCl_{3(s)}$ $\Delta H_4 = 2x + 132.0\,kJ$

(iii) $3H_{2(g)} + 3Cl_{2(g)} \longrightarrow 6HCl_{(aq)}$ $\Delta H_3 = -5535\,kJ$

(ii) $6HCl_{(g)} \longrightarrow 6HCl_{(g)}$ $\Delta H_2 = -448.8\,kJ$

(i) $2M_{(s)} + 6HCl_{(aq)} \rightarrow 2MCl_{3(aq)} + 3H_{2(g)}$ $\Delta H_1 = -619\,kJ$

$$\cancel{2MCl_{3(aq)}} + \cancel{3H_{2(g)}} + 3Cl_{2(g)} + \cancel{6HCl_{(g)}} + 2M_{(s)} + \cancel{6HCl_{(aq)}} \longrightarrow 2MCl_{3(s)} + \cancel{6HCl_{(aq)}} + \cancel{6HCl_{(g)}} + \cancel{2MCl_{3(aq)}} + \cancel{3H_{2(g)}}$$

$$2M_{(s)} + 3Cl_{2(g)} \longrightarrow 2MCl_3$$

$$\Delta H_{rxn} = \Delta H_1 + \Delta H_2 + \Delta H_3 + \Delta H_4$$

$$= -619\,kJ + -448.8\,kJ + 5535\,kJ + 264$$

$$= \underline{-6,338.8\,kJ}$$

Q7.) Given

(i) $N_2(g) + O_2(g) \rightarrow 2NO(g)$ $\quad \Delta H^\circ = +180.7\ kJ\,mol^{-1}$

(ii) $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$ $\quad \Delta H^\circ = -113.1\ kJ\,mol^{-1}$

(iii) $2N_2O(g) \rightarrow 2N_2(g) + O_2(g)$ $\quad \Delta H^\circ = -163.2\ kJ\,mol^{-1}$

reaction $\quad N_2O(g) + NO_2(g) \rightarrow 3NO(g)$

- Reverse and balance

(i) $\cancel{N_2} + O_2 \rightarrow 2NO$ $\quad \Delta H^\circ = +180.7\ kJ/mol$

(ii) $NO_2 \rightarrow NO + \cancel{\frac{1}{2}O_2}$ $\quad \Delta H^\circ = 56.65\ kJ/mol$

(iii) $N_2O \rightarrow \cancel{N_2} + \cancel{\frac{1}{2}O_2}$ $\quad \Delta H^\circ = -81.6\ kJ/mol$

Overal

$$\Delta H^\circ_{rxn} = \Delta H^\circ + \frac{1}{2}\Delta H^\circ_2 + \frac{1}{2}\Delta H^\circ_3$$

$$= +180.7 + 56.55 - 81.6$$

$$= +155.65\ kJ/mol.$$

$$= \underline{155.7\ kJ/mol}$$

Q8.) $\Delta H_{rxn} = \sum$ bond breaking $- \sum$ bond formation

or

$\Delta H_{rxn} = \sum$ reactants bond energy $- \sum$ product's bond energy

$\Delta H_{rxn} = \sum \Delta H_r - \sum \Delta H_p$ (Substitute bond energy values)

$\Delta H_{rxn} = ((N-N) + (N-H) \times 4 + 2(F-F) - 4(H-F) + N \equiv N$

$= 2032\ kj/mol - 3201\ kj/mol$

$= \underline{-1169\ kj/mol}$

Q9.) $\Delta H_{rxn} = \sum$ bond broken $- \sum$ bond formed

Given $\Delta H_{rxn} = -137\ kj$

$-137\ kj = (612\ kj/mol + 432\ kj/mol) - (x + 824\ kj/mol)$

$x = 220\ kjmol^{-1} + 137\ kj$

$C-C = 357\ kj/mol$

prove: $1044 - 357 - 824 = \underline{-137}$

Q 10.) Combustion rxn

$C_3H_6O_{(l)} + 4O_2 \longrightarrow 3CO_{2(g)} + 3H_2O_{(l)}$

We know that

$\Delta H_c^\circ = \Delta H_f^\circ$ (products) $- \Delta H_f^\circ$ (reactants)

Q 10) $\Delta H^{\circ}_{combustion} = -1658$ kJ/mol

$\Delta H^{\circ}_f (H_2O) = -241.8$ kJ/mol $\times 3$

$\Delta H^{\circ}_f (CO_2) = -393$ kJ/mol $\times 3$

Equation $C_3H_6O_{(l)} + 4O_{2(g)} \longrightarrow 3H_2O_{(l)} + 3CO_{2(g)}$ -1658 kJ/mol

Reverse and balance

$\cancel{3CO_{2(g)}} + \cancel{3H_2O_{(l)}} \longrightarrow C_3H_6O_{(l)} + \cancel{4O_2}$ $\Delta H^{\circ}_a = +1658$ kJ/mol

$3H_{2(g)} + \cancel{\frac{3}{2}O_{2(g)}} \longrightarrow \cancel{3H_2O_{(l)}}$ $\Delta H^{\circ}_f = -725.4$ kJ/mol

$3C_{(graphite)} + \cancel{3O_{2(g)}} \longrightarrow \cancel{3CO_{2(g)}}$ $\Delta H^{\circ}_f = -1179$ kJ/mol

---

$3C_{(s)} + 3H_{2(g)} + O_{(g)} \longrightarrow C_3H_6O_{(l)}$ $\Delta H^{\circ}_f = -246.4$ kJ/mol

~~ΔH°_{rxn}~~ =

$\Delta H^{\circ}_{rxn} = \underbrace{[c\,\Delta H^{\circ}_f(C) + d\,\Delta H^{\circ}_f(D)]}_{\text{products}} - \underbrace{[a\,\Delta H^{\circ}_f(A) + b\,\Delta H^{\circ}_f(B)]}_{\text{Reactants}}$

(16) $\Rightarrow$ Heat gained by the water

$$Q_{water} = mc_{water}(T_2 - T_1)$$
$$= (1200g)(4.814 J/(g°C))(33.20°C - 25.00°C)$$
$$= 41170.56 J$$

$\rightarrow$ Heat of combustion of octane.

$$\Delta H_{rxn} = -\frac{Q_{calorimeter} + Q_{water}}{n}$$

Substitute known values

$$\Delta H_{rxn} = -\frac{6863.4 J + 41170.56 J}{\frac{1.000}{114.23} mol}$$
$$= -5486919.2508 J/mol \left[\frac{1 KJ}{1000 J}\right]$$
$$= -5487 KJ/mol$$

$\therefore$ Heat of combustion of octane is $-5487 KJ/mol$

$Q = MC\Delta T$

Q 11.) Solution

– Combining equation (i) and (ii)

(i) $C_2H_{4(g)} + 2H_{2(g)} \longrightarrow \cancel{2CH_{4(g)}}$ $\Delta H^\circ_{rxn} = -201.9\,kJ$

(ii) $\cancel{2CH_4(g)} + 2O_{2(g)} \longrightarrow 2CO_2(g) + \cancel{2}4H_2O_{(l)}$ $\Delta H^\circ_{rxn} = -890.3 \times 2$
$= -1{,}780.6$

→ $\cancel{C_2H_{4(g)}} + \cancel{2}H_{2(g)} + 2O_{2(g)} \longrightarrow 2CO_{2(g)} + 4H_2O_{(l)}$
$\Delta H^\circ_{rxn} = -1982.5\,kJ$

(iii) $C_2H_{6(g)} \longrightarrow \cancel{C_2H_4(g)} + \cancel{H_2(g)}$ $\Delta H^\circ_{rxn} = +137.0\,kJ$

→ $C_2H_{6(g)} + \cancel{2}O_{2(g)} \cancel{+H_{2(g)}} \longrightarrow 2CO_2(g) + {}^3\cancel{4}H_2O$ $\Delta H^\circ_{rxn} = -1845.5$

(iv) $\cancel{H_2O_{(l)}} \longrightarrow \cancel{H_2(g)} + \cancel{\frac{1}{2}O_2}g$ $\Delta H^\circ_{rxn} = +285.8\,kJ$

$\rightarrow C_2H_{6(g)} + 7O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O$

$\Delta H^\circ_{rxn} = -1559.7\,kJ$

multiplying the above equation by 2

$2C_2H_6 + \frac{7}{2}O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O$

$\rightarrow 2C_2H_{6(g)} + 7O_{2(g)} \rightarrow 4CO_{2(g)} + 6H_2O_{(l)}$

$\Delta H^\circ_{rxn} = -1559.7\,kJ \times 2$

$= -3119.4\,kJ$

Reversing equation (i) we get

$6CO_{2(g)} + 3H_2O_{(l)} \rightarrow C_6H_{6(l)} + \frac{15}{2}O_2 : \Delta_f H^\circ = +3271\ kJmol^{-1}$ --- (vi)

Adding equation (v) and (vi) we get

$6C_{(graphite)} + 3H_{2(g)} \longrightarrow C_6H_{6(l)}$ --- (iv)

$(\Delta H^\circ)_{Benzene} = -3222\ kJmol^{-1} + 3271\ kJmol^{-1}$

$= 49\ kJmol^{-1}$

$(\Delta_f H^\circ)_{Benzene} = 49\ kJmol^{-1}$

I learn from the mistakes of people who take my advice

THE COPPERBELT UNIVERSITY
SCHOOL OF MATHEMATICS AND NATURAL SCIENCES
DEPARTMENT OF CHEMISTRY

CHEMISTRY (CH 110/ FO 130) TUTORIAL SHEET 6 TERM TWO YEAR 2024

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1. Fill in the gaps in the following table.

Symbol	$^{59}Co^{3+}$			
Protons		34	76	80
Neutrons		46	116	120
Electron		36		78
Net Charge			2+	

2. Calculate the energy in joules of an individual photon of red light of frequency 3.57 x 10^{14} sec^{-1}
3. Calculate the wavelength and frequency of the light emitted when there is an electron transition in the hydrogen atom from n = 4 to n = 2.
4. Calculate the energy required to move an electron from n = 2 to n = 3 principal energy levels of the hydrogen atom.
5. State the postulates of the Bohr Theory and the major deficiency of the theory.
6. What is the de Broglie wavelength of a 50 kg person jogging at a speed of 2.70 m/s?
7. Differentiate orbits and orbitals. Which atomic model are they derived from?
8. (a) List the four quantum numbers, tell what each identifies and state the values that each may assume
   (b) On a single Cartesian coordinate axes, sketch the shapes of the three p orbitals. Label them p_x, p_y and p_z.
   (c) Sketch the shapes of s and p orbitals. How are p orbitals oriented in relation to one another?
9. Which of the following electronic configurations represent the ground state and which ones represent the excited state electronic configuration? In each case justify your answer.
   (a) $1s^2\ 2s^2\ 2p^3$
   (b) $1s^2\ 2s^2\ 3p^1$
   (c) $1s^1\ 2s^2\ 2p^4\ 3s^2$
   (d) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5\ 3d^{10}\ 4s^2$
10. Give the values of the quantum numbers (n, l, m_l) associated with the following orbitals: (a) 2s (b) 3p (c) 5d
11. List all the possible subshells and orbitals associated with the principal quantum number, n = 3.
12. An element x belongs to the fourth period of the Periodic Table. You are asked to compare in terms of all the quantum numbers of the following electrons in x:

(a) the second and the eleventh electrons
(b) the fifth and the eighth electrons
(c) the ninth and the sixteenth electrons
(d) the eighteenth and nineteenth electrons

13. Give both the electronic configurations and orbital diagrams for each of the following: (a) Zn (b) S^{2-} (c) Cr (d) Cr^{2+} (e) Cu^{1+}
14. Write the ground state electronic configuration (spdf) notation for Palladium (Pd) and determine:
    (a) the last energy level
    (b) the types of orbitals in the last energy level
    (c) the total number of orbitals in the last energy level
    (d) the number of paired and unpaired electrons in the last energy level (e) the sum of electron spin numbers in the last energy level.
15. Arrange the following species into groups of iso-electronic species.
    Sc^{3+}, N^{3-}, B^{3+}, Rb^{+}, P^{3-}, Ca^{2+}, C^{4+}, Be^{2+}, Mg^{2+}, O^{2-}.
16. Explain why the ground state electronic configuration for molybdenum (Z=42) and silver (Z= 47) are different from what we might expect.
17. Which atom/ion in each of the following pairs would have lower first ionization energy?
    (i) Al or B (ii) F or N (iii) Mg or Na (iv) Br or Cl (v) Na^{+} or Mg^{2+} (vi) Fe^{2+} or Fe^{3+}
18. Define each of the following terms and explain their variation in the periodic table:
    (i) Ionization energy (ii) Electron affinity (iii) Electronegativity
19. Explain the following terms;
    (i) Hunds rule,
    (ii) Heisenberg`s uncertainty principle,
    (iii) Pauli Exclusion principle, (iv) Paramagnetic substances, (v) Iso-electronic species.

THE ENDFM2024

Atomic Structure

Q1. → # of protons = # of electrons

mass # = protons + neutrons

→ −ve means gain of electrons and +ve means loss of electrons.

Symbol	$^{59}Co^{3+}$	Se^{2-}	Os^{2+}	Hg^{2+}
Protons	27	34	76	80
Neutrons	32	46	116	120
Electrons	24	36	74	78
Net Charge	3+	2−	2+	2+

Note: Isotopes have same proton number but different neutron number hence mass number may vary for the same element.

Q2. $E = h\nu$

$= 6.626 \times 10^{-34}\ J\cdot s \times 3.57 \times 10^{14}\ s^{-1}$

$= \underline{2.365 \times 10^{-19}\ J}$

Q3. method 1

$$\Rightarrow \frac{1}{\lambda} = R\left(\frac{1}{n_i^2} - \frac{1}{n_f^2}\right)$$

Note: We will use method 2 so that you are familiar with the relationships of equations

Method 2

- We first find change in energy ΔE then use it to find frequency and finally use frequency for wavelength.

$$\rightarrow \Delta E = R_H \left[\frac{1}{n_i^2} - \frac{1}{n_f^2} \right] \quad \text{------ (i)}$$

$$\rightarrow \nu = \frac{\Delta E}{h} \quad \text{------ (ii)} \quad \boxed{\Delta E = h\nu}$$

$$\rightarrow \lambda = \frac{c}{\nu} \quad \text{------ (iii)} \quad \boxed{c = \lambda\nu}$$

Data

$n_i = 4 \qquad n_f = 2$

$$\Delta E = R_H \left[\frac{1}{n_i^2} - \frac{1}{n_f^2} \right]$$

$$= 2.18 \times 10^{-18}\,J \left[\frac{1}{4^2} - \frac{1}{2^2} \right]$$

$$= -4.0875 \times 10^{-19}\,J$$

$\Delta E = h\nu$

$$\nu = \frac{\Delta E}{h} = \frac{4.0875 \times 10^{-19}\,J}{6.63 \times 10^{-34}\,J.s} = 6.165 \times 10^{14}\,s^{-1}$$

$$\lambda = \frac{c}{\nu}$$

$$\lambda = \frac{3 \times 10^{8}\,m/s}{6.165 \times 10^{14}\,s^{-1}} = 4.866 \times 10^{-7}\,m$$

Can convert to nm

Q4. $\Delta E = R_H \left[\frac{1}{n_i^2} - \frac{1}{n_f^2} \right]$

$= 2.18 \times 10^{-18} J \left[\frac{1}{2^2} - \frac{1}{3^2} \right]$

$= 2.18 \times 10^{-18} J \times 0.13889$

$= \underline{\underline{3.0278 \times 10^{-19} J}}$

Q5. Bohr theory

- Electrons revolve around the nucleus in a certain selected circular path called orbit.
- Electrons absorb energy from lower to higher energy level (vice versa)
- Each shell or an orbital has fixed energy, Circular orbitals are known as the orbital shells.
- Spectrum of light emitted from an electron is called the line spectrum

limits

- failed to explain the spectra of atoms with ~~than~~ more than one electron.
- Could not explain the shape of orbitals
- Violation of Heisenberg's Uncertainity principle (suggests that both position and direction of an electron can be determined)
- Cant explain the splitting of spectral lines in presence of magnetic field and electric field.

Q6. $\lambda = \frac{h}{mv}$

m = mass v = speed h = plank's constant

$\lambda = \frac{6.63 \times 10^{-34} \, J.s}{50 \, kg \times 2.70 \, m/s}$

$= 4.91 \times 10^{-36} \, Js^2/kgm$

$\boxed{1 J = 1 kg \, m^2/s^2}$

$= 4.91 \times 10^{-36} \, \cancel{J}\cancel{s^2}/\cancel{kg}m \cdot \frac{1 \cancel{kg} m^{\cancel{2}}/\cancel{s^2}}{\cancel{J}}$

$= 4.91 \times 10^{-36} \, m$

Q7. - An orbit is a fixed path around the nucleus in which electrons revolve while an orbital is an area where you can find an electron.

- Borh model and Quantum mechanic model respectively.

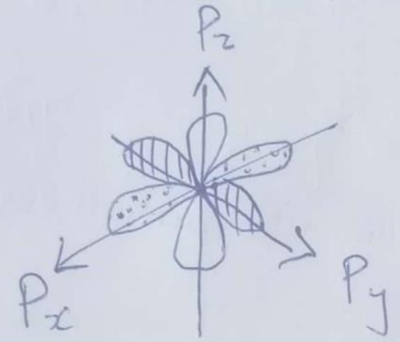
Q8. (a) n = energy and distance (1, 2, 3 ----)

l = shapes of orbitals l = n-1 (0, 1, 2, 3 --)

m_l = Orientation of orbitals in space (-l 0 l) depends on the value of l (2l+1)

m_s = Takes electron spin into account $+\frac{1}{2}$ or $-\frac{1}{2}$

(b)



(c) s = O P = 8 P = orientated perpenticular to each other.

Q9. → if the element is given, ~~writte~~ write the ground state configuration for the element and compare with the given configuration if they are different then the given E.C is in exacited state.

→ In this question no particular element is given so ignore the last orbital and check if the other orbitals have filled up electrons if they are not filled up as expected then it is exacited.

(a) ground state (b) ground state (c) exacited state
(d) exacited state.

Q 10. (a) 2s

$n = 2$

$l = 0$

$ml = 0$

$ms = \frac{1}{2}$ or $-\frac{1}{2}$

(b) 3p

$n = 2$

$l = 1$

$ml = -1 \ 0 \ 1$

$ms = \frac{1}{2}$ or $\frac{-1}{2}$

(c) 5d

$n = 5$

$l = 2$

$ml = -2 \ -1 \ 0 \ 1 \ 2$

$ms = \frac{1}{2}$ or $\frac{-1}{2}$

Q 11. Where $n = 3$

$l = 2$, with all ~~pos~~. values from 0, 1, 2

0 = s orbital
1 = p orbital
2 = d orbital

Since $n = 3$ subshells will be 3s, 3p & 3d

Total orbitals $s = 1$ $p = 3$ $d = 5$

Total $= 9$ or $n^2 = 3^2 = 9$

Total electrons : since each orbital can only have 2 max. number of electrons

$9 \times 2 = 18$ electrons

or

$$2n^2 = 2(3)^2$$
$$= 2 \times 9$$
$$= 18 \text{ electrons}$$

Q 12. To find element X we must know both the group and period where it belongs. But we are only given the period which is 4p.

- Thus we will use the max number of electrons in our question to get the electronic configuration for X. which is 19 from question d.

Solution

E.C for $X^{19} = 1s^2\ 2s^2\ 2p^6\ 3s^2 3p^6 4s^1$

Orbital diagram: $s = 0$ $p = 1$
$l = 0$ $l = 1$

1 ↓		1 ↓		1↓ 1↓ 1↓		1 ↓		1↓ 1↓ 1↓		1
$1s^2$		$2s^2$		$2p6$		$3s^2$		$3p6$		$4s^1$

(a)

2nd	11 th
$n = 1$	$n = 3$
$l = 0$	$L = 0$
$ml = 0$	$ml = 0$
$m_s = -\frac{1}{2}$	$m_s = \frac{1}{2}$

(b)

5 th	8 th
$n = 2$	$n = 2$
$l = 1$	$L = 1$
$ml = -1\ 0\ 1$	$ml = -1\ 0\ 1$
$m_s = \frac{1}{2}$	$m_s = -\frac{1}{2}$

(c)

9 th	16 th
$n = 2$	$n = 3$
$L = 1$	$l = 1$
$ml = -1\ 0\ 1$	$ml = -1\ 0\ 1$
$m_s = \frac{1}{2}$	$m_s = -\frac{1}{2}$

d

18th	19 th
$n = 3$	$n = 4$
$L = 1$	$l = 0$
$ml = -1 0 1$	$ml = 0$
$m_s = -\frac{1}{2}$	$m_s = \frac{1}{2}$

Note : Pauli's Exclusion principle : In an atom no two electrons will have the same values for all the 4 quantum numbers.

→ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow}$

$+\frac{1}{2}$ $-\frac{1}{2}$ $+\frac{1}{2}$

Q 13. Due to higher energy levels for transition metals the orbitals become degenerated; degenerate orbitals are orbitals with same energy.

∴ d - orbitals are more stable when filled in completely thus s orbitals donates electrons to the d-orbitals for their stability hence we write 3d 4s instead of 4s 3d.

(a) $Zn = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$

orbital diagram $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}$ $\boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}$ $\boxed{\uparrow}$

(b) $S^{2-} = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

O.D $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}$

(c) $Cr = 1s^2\ 2s^2\ 2p^6\ 3s^2\ \underbrace{3d^5\ 4s^1}$

O.D $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow|\uparrow\downarrow|\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow|\uparrow|\uparrow|\uparrow|\uparrow}$ $\boxed{\uparrow}$

from $4s^2\ 3d^4$
to $3d^5\ 4s^1$

(d) $Cr^{2+} = 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^4$

[↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑|↑|↑|↑|]

(e) Cu^+ = Expected ground state $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$

$4s^2 \xrightarrow{+1} 3d^9 = 4s^1\ 3d^{10}$

$= 3d^{10}\ 4s^1$

Correct $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

Thus for Cu^+ lost one electron will be

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$

[↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓] [↑↓|↑↓|↑↓] [↑↓|↑↓|↑↓|↑↓|↑↓]

Q 14. palladium (Pd) 46

Expected ground state

→ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^8$

∴ $5s^2 \xrightarrow{+2} 4d^8 \rightarrow 5s^0\ 4d^{10}$

$\rightarrow 4d^{10}\ 5s^0$ $\qquad 5s^0 = 0$

So we can do away with it.

E.C = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 4d^{10}$

→ full d orbitals are more stable than partially filled ones

(a) last energy level $n = 4$

(b) orbitals = s, p, d, d from 4s, 4p, 4d

(c) 9 orbitals

(d) paired = 9 unpaired = 0

s = 1, p = 3, d = 5

(e) all are paired so $+\frac{1}{2}$ and $-\frac{1}{2}$ for all orbitals will be 0.

Q15. Iso-electronic species (Same number of electrons)

Sc^{3+}	N^{3-}	B^{3+}	Rb^{+}	P^{3-}	Ca^{2+}	C^{4+}	Be^{2+}	Mg^{2+}	O^{2-}
↓	↓	↓	↓	↓	↓	↓	↓	↓	↓
18	10	2	36	18	18	2	2	10	10

(i) Sc^{3+}, P^{3-}, Ca^{2+}

(ii) N^{3-}, Mg^{2+}, O^{2-}

(iii) B^{3+}, C^{4+}, Be^{2+}

(iv) Rb^{+}

Q16. $Mo = 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^5 5s^1$

$Ag = 1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^1$

- Same reason given for transition metals in the previous questions.

Q 17. lower first ionization energy

(i) Al or B (closer) = Al

(ii) F or N = N

(iii) Mg or Na = Na

(iv) Br or Cl = Br

(v) Na^{+} or Mg^{2+} = Na^{+} (nuclear effective charge)

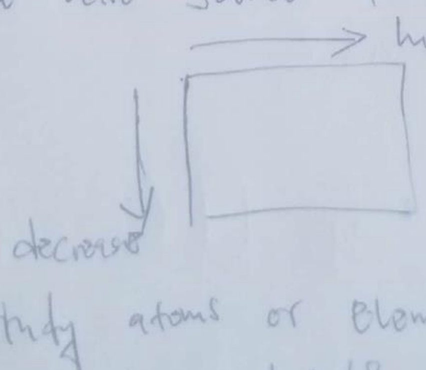
(vi) Fe^{2+} or Fe^{3+} = Fe^{2+}

Q 18. (i) Energy required to remove an outermost electron from neutral atom in gas phase.

(ii) Energy released when an electron is added to a neutral atom.

(iii) Ability of an atom to attract electrons towards itself.

all three have same trend on the periodic table



Note: Study atoms or elements with exceptions to the above trends.

Q 19 (i) Hund's rule: every atomic orbital within a sublevel is singly occupied before it is doubly occupied: All singly occupied orbitals posses electrons with same spin $\boxed{\uparrow\downarrow}$ $\boxed{\uparrow|\uparrow|\uparrow}$

(ii) Heisenberg's Uncertainty principle: It is impossible to measure both position and momentum of an object (wave - partical duality of matter)

(iii) Pauli Exclusion principle: In a single atom no two electrons will have same values for all 4 quantum numbers.

(iv) paramagnetic substances: Substances attracted by magnetic field and have unpaired electrons.

(v) Iso- electronic species; Species with same number of electrons.

END !!!


Today 14 girls asked me to go out!!
(I was in girls washroom)
7,776
188

THE COPPERBELT UNIVERSITY
SCHOOL OF MATHEMATICS AND NATURAL SCIENCES
DEPARTMENT OF CHEMISTRY

CHEMISTRY (CH 110/ FO 130) TUTORIAL SHEET 7 TERM THREE YEAR 2024

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1. Draw Lewis structures of the following molecules:

   a) $COCl_2$ b) CCl_2F_2 c) H_3PO_3 d) H_2SO_4 e) HNO_3 f) HCN

2. Draw Lewis structures of the following ions:

   a) NO_3^- b) NO^+ c) NO_2^- d) PO_4^{3-} e) I_3^- f) ClO_4^- g) $TiCl_2^+$

3. Write the Lewis structures of each of the following molecules and ions and assign formal charges to the atoms involved:

   a) PO_4^{3-} b) OH^- c) NO_2 d) H_2SO_4 f) OCl^-

4. Draw resonance structures of each of the following molecules and ions. Also show the formal charges on atoms:

   a) CO_3^{2-} b) N_2O_2 c) OCN^- d) NO_2^- e) SO_3 h) O_3

5. Use VSEPR theory to predict the molecular geometry of the following molecules and ions.

   In each case, state for whether the molecule/ion is polar or nonpolar.

   a) H_3O^+ b) BrF_5 c) NH_3 d) NH_4^+ e) H_2O f) CO_2 g) PF_5 h) IF_5

6. Describe the bonding scheme in the following molecules in terms of hybridization.

   a) BCl_3 b) AsH_3 c) CO_2 d) XeF_4 e) H_2O f) NH_3

7. Arrange the following bonds in order of increasing ionic character, justify your answer:

   N – H, F – H, C – H and O – H.

8. List the common types of intermolecular forces observed in various compounds.

9. What type of intermolecular forces will be in each of the following compounds:

   a) HCl b) CBr_4 c) Ethanol d) H_2O e) NH_3 f) CH_4

10. The equations for the heats of formation of copper (I) oxide, Cu_2O and copper (II) oxide, CuO are:

    $2Cu(s) + ½O_2(g) \rightarrow Cu_2O(s)$; $\Delta H^0_f = -166.7 kJ/mol$

    $Cu(s) + ½O_2(g) \rightarrow CuO(s)$; $\Delta H^0_f = -155.2 kJ/mol$

    The first and second ionization energies of copper are 750 and 2000kJ/mol respectively; its atomization energy is 339.3kJ/mol. The atomization energy of oxygen is 249.2kJ/mol.

The first and second electron affinity for oxygen are -141.4kJ/mol and 790.8kJ/mol

respectively. Draw the Born-Haber cycle and calculate the lattice energies (ΔH_{latt}) of the two oxides

11. Use the data provided below to calculate the lattice energy of RbCl.

    a. Electron affinity of Cl = –349 kJ/mol
    b. 1st ionization energy of Rb = 403 kJ/mol
    c. Bond energy of Cl_2 = 242 kJ/mol
    d. Sublimation energy of Rb= 86.5 kJ/mol V. ΔHf [RbCl (s)] = – 430.5 kJ/mol

    Rb (s) + ½ Cl_2 (g) ➔ RbCl (s)

THE END

Mg (IE) 2nd 1390 [illegible]
more [illegible] [illegible]

$Cu_{(s)} + \frac{1}{2} O_{2(g)} \xrightarrow{\Delta H_f = -155.2\ kJ/mol} CuO_{(s)}$

$\Delta H_a = 339\ kJ$ ↓ $Cu_{(g)}$

$\Delta H_a = 249.2\ kJ/mol$ ↓ $O_{(g)}$

1st IE 750 ↓ $Cu^{+}_{(g)}$

1st EA −141.4 ↓ $O^{-}_{(g)}$

2nd IE 2000 ↓ $Cu^{2+}_{(g)}$

2nd EA 790.8 ↓ $O^{2-}_{(g)}$

$Cu^{2+}_{(g)} + O^{2-}_{(g)} \xrightarrow{\Delta H_{latt}} CuO_{(s)}$

Hess's law

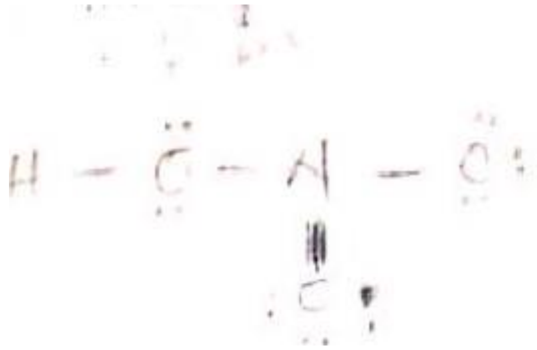
$$\Delta H_f = \sum \Delta H + \Delta H_{latt}$$

$$\Delta H_{lattice} = \Delta H_f - \sum \Delta H$$

$$= -155.2 - [339.1 + 249.2 + 750 + -141.4 + 2000 + 790.8]$$

$$= -155.2 - [3,987.9]$$

$$= -4,143.1\ kJ/mol$$



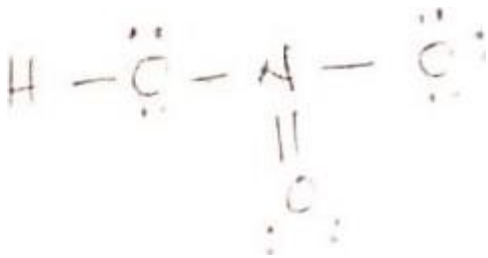
F.C

$N = 5 - 0 - \frac{8}{2}$

$= 5 - 4$

$= +1$

$C = -1$



(F) HCN . 10 v E

$$H - C - \ddot{N}:$$

$$\downarrow$$

$$H - C \equiv N:$$

F.C

N = -2

C = +2

H = 0

H — N — H (sp³)
|
H

Q7 C − H < N − H < O − H < F − H

- greater the electronegativity difference the greater the ionic character

Q8 – Hydrogen bonding $\overset{\text{differ electron}}{H - X}$

– London dispersion forces: due to formation of temporal dipole due to movement of electron ... btn molecules

– Dipole - Dipole forces (Strongest)
formed due to electrostatic interactions btn two or more dipoles

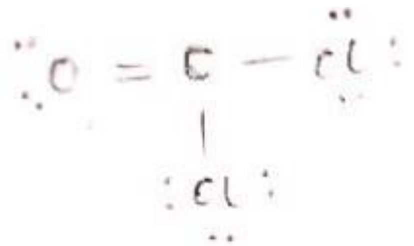
– Dipole - induced dipole forces: due to interaction btn an uncharged atom or molecules with a dipole

Q9 (a) HCl ; Dipole - Dipole, polar covalent bond btn hydrogen and chlorine

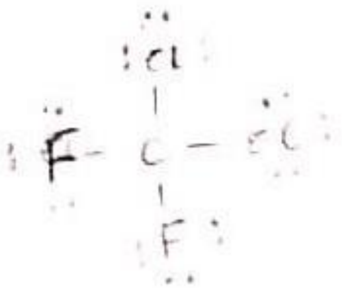
(b) CBr_4 non-polar so can't form dipole bi... or hydrogen

- London dispersion forces due to fluctuations in the electron cloud of atoms / molecules

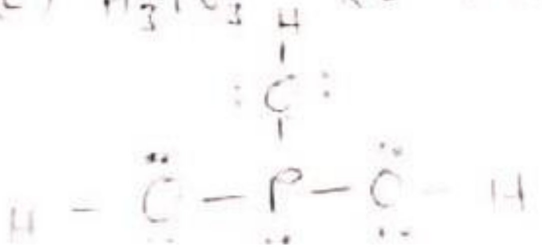
- the larger the ... the stronger the London ... to large Br ...



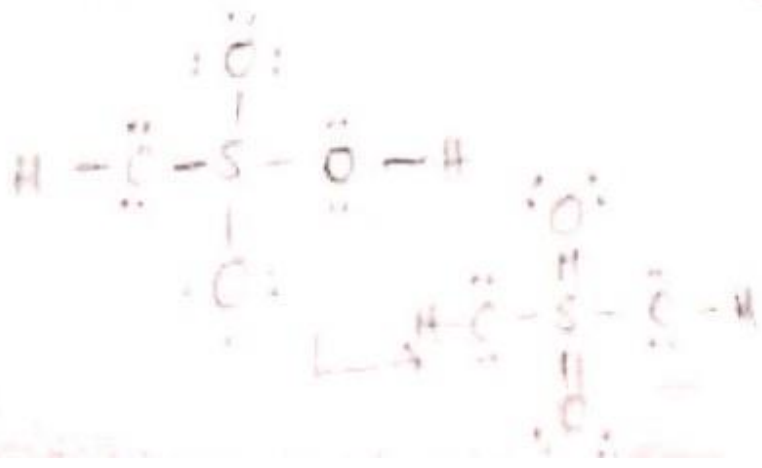
b) CCl_2F_2 32 V.E



(c) H_3PO_3 26 V.E



(d) H_2SO_4 32 V.E



F.C

S = +2

O = -1

O = 0

H = 0

$2Cu_{(s)} + \frac{1}{2}O_{2(g)} \xrightarrow{\Delta H_f = -166.7} Cu_2O_{(s)}$

2×339.2 ↓   ↓ 249.2

$Cu_{(g)}$ $O_{(g)}$

$2 \times 750 =$ ↓   ↓ -141.4

Cu^+ O^-

↓ 798.8

O^{2-}

lattice

$Cu^+ + O^{2-}$

$\Delta H_{lattice} = \Delta H_f - \sum \Delta H$

$= -166.7\ kJ/mol - [678.4 + 249.2 + 1500 + -141.4 + 798.8]$

$= -166.7 - [$

$= -3{,}293.7\ kJ/mol$

$[:\ddot{Cl} - Ti$

Q3. (a) PO_4^{3-}

F.C

$P = 5 - 0 - \frac{8}{2}$

$= +1$

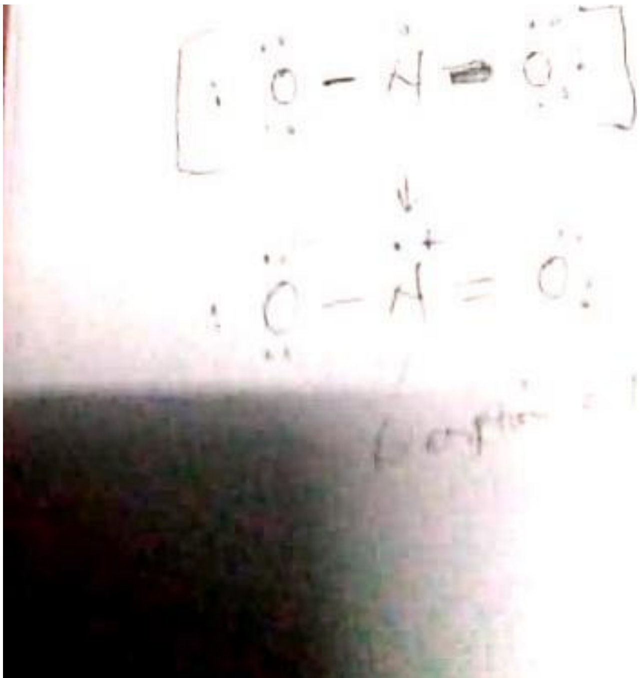
$O = 6 - 6 - \frac{2}{2}$

(b) OH^-

F.C

$O = -1$

$H = 0$



C

Hybridisation = $\frac{4+2}{2} = \frac{6}{2} = 3$

$= 3$

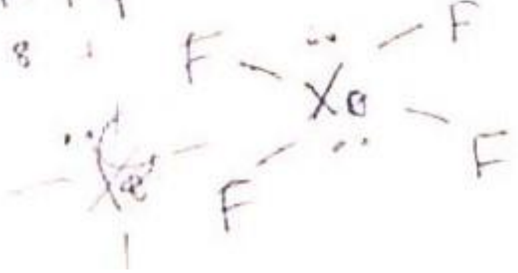
$= sp^2$ hybridisation

C $1s^2 2s^2 2p^2$

[↑↓] [↑|↑|]

XeF_4 Xenon tetrafluoride 6 sp^3d^2

8



$\frac{4+8}{2} =$

Q6 (a) BCl_3 Hybridization $= \frac{V.e + T.c}{2}$

B : E.C $1s^2 2s^2 2p^1$

V.e - valence electron of central atom

BCl_3

T.C - Total atoms attached to central atom

$= \frac{3+3}{2} = \frac{6}{2} = 3$

$\therefore BCl_3 = Sp^2$ hybridized

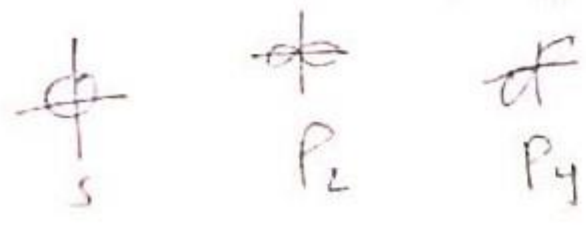
$4 = Sp^3$ $3 = Sp^2$
$2 = Sp$

$B \rightarrow 2s^2 2p^1$

g.s [1↓] [1 | |]

E.S [1] [1 | 1 |]

Hybride State [11 | 11 | 11] sp^2
Cl Cl Cl



(b) AsH_3

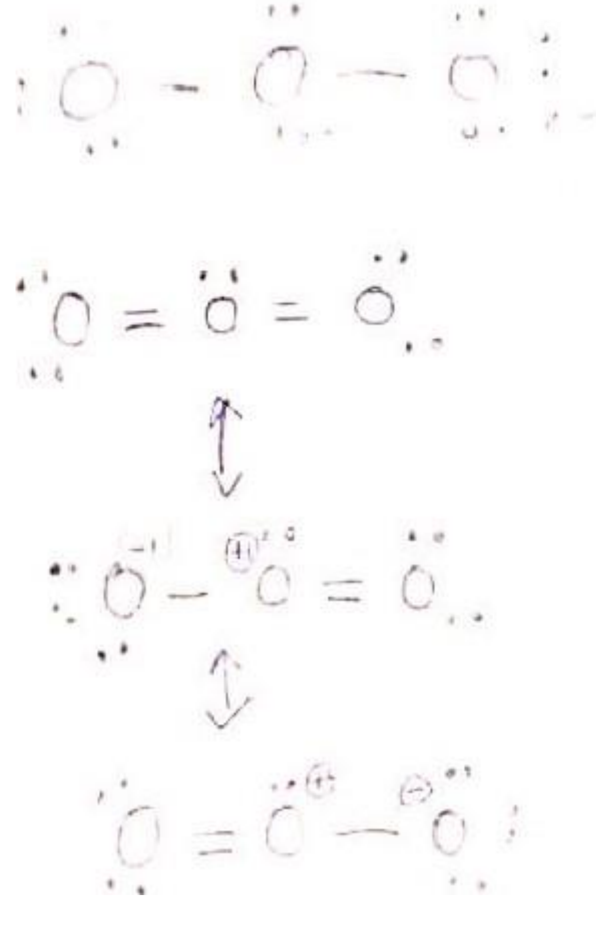
H - As - H
|
H

$= \frac{5+3}{2} = \frac{8}{2} = 4$

$= Sp^3$ hybridization

[Ar] $3d^{10} 4s^2 4p^3$

[11] [1 | 1 | 1]
H H H



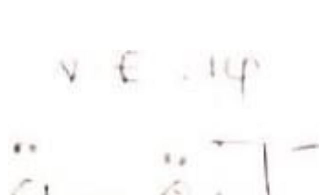
$O = 6 - 2 - \frac{4}{2}$
$= 6 - 4$
$= +2$

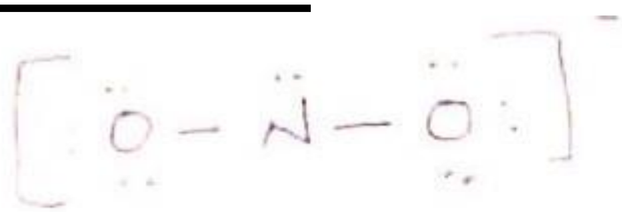
$O = 6 - 1$

central atom lone pair
- polar

number = 4
2 pair
trigonal pyramidal

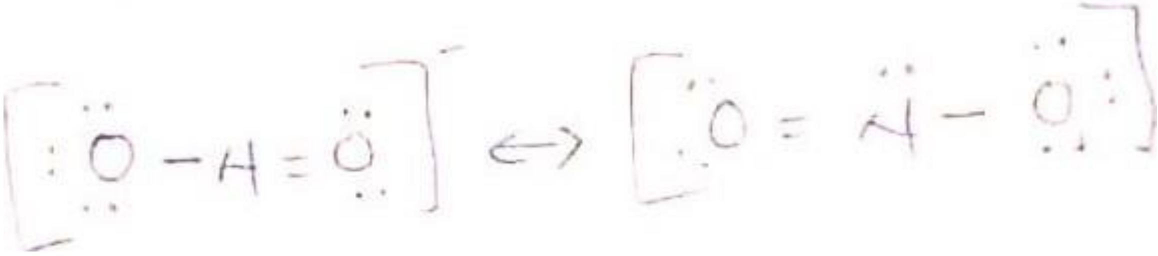
V E : 14





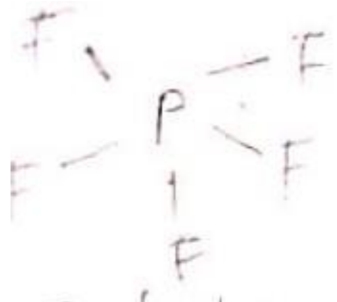
N = +1

O = -1



$\ddot{O} = C = \ddot{O}$ linear and non polar
dipole moments cancel out

more than 8 octets

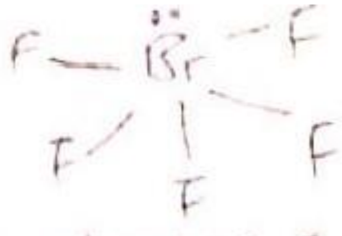


non-polar (symmetrical)
zero net dipole

5 bonds, no lone pair

Trigonal bipyramidal

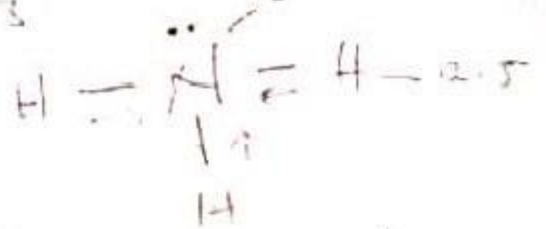
square pyramidal
total sites 6
1 lone pair



geometry: Octahedral
square pyramidal

Steric Number = 6
1 lone pair on central atom = polar
+ difference in electronegativity = unequal charge distribution on
more electronegative central atom

(c) NH_3 → polar



Trigonal pyramidal

(d) NH_4^+

Tetrahedral
non-polar (no lone pair on central atom)
highly symmetrical

H_2O → polar

Bent (tetrahedral)

Q 11 $Rb_{(s)} + \frac{1}{2} Cl_{2(g)} \xrightarrow{\Delta H_f = -430.5} RbCl_{(s)}$

$Rb_{(s)} \xrightarrow{86.5} Rb_{(g)} \xrightarrow{403} Rb^+_{(g)}$

$\frac{1}{2} Cl_{2(g)} \xrightarrow{\frac{242}{2}} Cl_{(g)} \xrightarrow{-349} Cl^-$

$Rb^+_{(g)} + Cl^- \xrightarrow{\Delta H_{latt}} RbCl_{(s)}$

$$\Delta H_{latt} = \Delta H_f - [86.5 + 121 + 403 + -349]$$

$$\Delta H_{latt} = -430.5 - [261.5]$$

$$= -692 \text{ KJ/mol}$$

Q 9 (c) Ethanol : polar molecule OH

H = O Forms
→ Hydrogen bonding (primary bonds))
→ dipole dipole
→ all molecules have london

H_2O

→ strong Hydrogen bonding

- polar → Hydrogen bonding
  → Dipole Dipole (polar)

CH_4 ⇒ London dispersion (non-polar)
temporal dipole

DRUG Nshima (4 lamps).

4	4		4

DATE

QTY 4 TDS

SPECIAL INSTRUCT

[redacted] Take 3 tamie daily

NAME Everyone EXPIRY DATE: N/A

CANCER DISEASES HOSPITAL PHARMACY

COPPERBELT UNIVERSITY

CHEMISTRY DEPATMENT

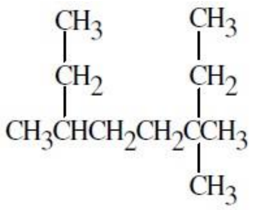
CH110 Tutorial Sheet - Organic Chemistry

2024 Academic Year

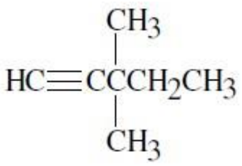
1. Give the IUPAC name of the following hydrocarbons:

   a) $CH_3CH_2C(CH_3)_2CH_2CH_2CH_2CH_3$

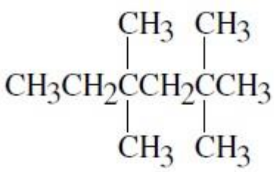
   b)



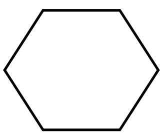
   c)

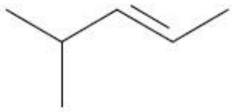


   d)

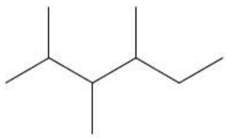


   e)

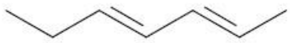




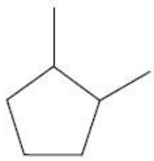
   f)



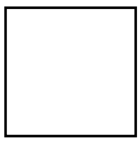
   g)



   h)



   i)



2. For each of the substances in question 1, identify whether it is saturated or unsaturated.

   Explain how you can test for unsaturation.

3. Draw the structures of each of the following:

   a) 2,3,5,6-tetramethyloctane.

   b) cis-3,4-dimethyl-3-hexene

   c) 1-ethyl-3-methylcycloheptane

d) 4,4-dimethyl-2-pentyne

e) 6-methyl-2,4-nonadiene

f) ethylcyclopentane

g) 2,3-dimethyl-1,3-butadience

h) trans-2-butene

i) 2,3,3-trimethyl-1-hexene

j) 2-tert-butylpentane

4. a) Draw nine (9) possible straight chain structures for the compounds with the formula

   C_5H_8

   b) Give the IUPAC name for each of the structures you have drawn in (a) above.

   c) Why are such compounds called isomers?

5. Draw and name all the structural isomers of each of the following:

   a) C_5H_{12}

   b) C_6H_{14}

   c) C_7H_{16}

6. Predict the products of the following reactions:

   a) Pentane + $\longrightarrow$ O_2

   b) 2- $\longrightarrow$ methylpropane + Br_2

   c) 2,3-dimethylhexane + O_2 $\xrightarrow{hv}$

   d) 3-ethyl-5-methylheptane + Cl_2 $\xrightarrow{hv}$

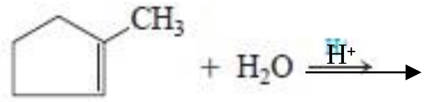
7. Reagents such as HCl, HBr, and HOH (H_2O) can add across carbon–carbon double and triple bonds, with H forming a bond to one of the carbon atoms in the multiple bond and Cl, Br, or OH forming a bond to the other carbon atom in the multiple bond. In some cases, two products are possible. For the major organic product, the addition occurs so that the hydrogen atom in the reagent attaches to the carbon atom in the multiple bond that already has the greater number of hydrogen atoms bonded to it. a) What is the name of the rule followed in this description?

   b) With this rule in mind, draw the structure of the major product in each of the following reactions:

   i) $CH_3CH_2CH{=}CH_2 + H_2O \xrightarrow{Pt}$ ii)

   $CH_3CH_2CH{=}CH_2 + HBr \rightarrow$ iii)

   $CH_3CH_2C{\equiv}CH + 2HBr \rightarrow$ iv)



8. Using appropriate reactants, alcohols can be oxidized into aldehydes, ketones, and/or carboxylic acids. Primary alcohols can be oxidized into aldehydes, which can then be oxidized into carboxylic acids. Secondary alcohols can be oxidized into ketones, while tertiary alcohols do not undergo this type of oxidation.

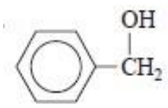
a) Draw the structure of the compound with three carbon atoms that belongs to each of the following families of organic compounds:

i) Alcohols ii) Aldehydes iii) Ketones iv) Carboxylic acids v) Esters vi) Ammines

b) Name each of the substances with the structures you have drawn in (a) above.

9. Give the structure of the product(s) resulting from the oxidation of each of the following alcohols.

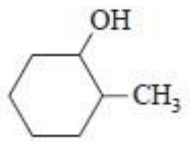
i) 3-methyl-1-butanol vi)



ii) 3-methyl-2-butanol

iii) 2-methyl-2-butanol

iv) Propanal vii)



v) 2,3-dimethyl pentanal

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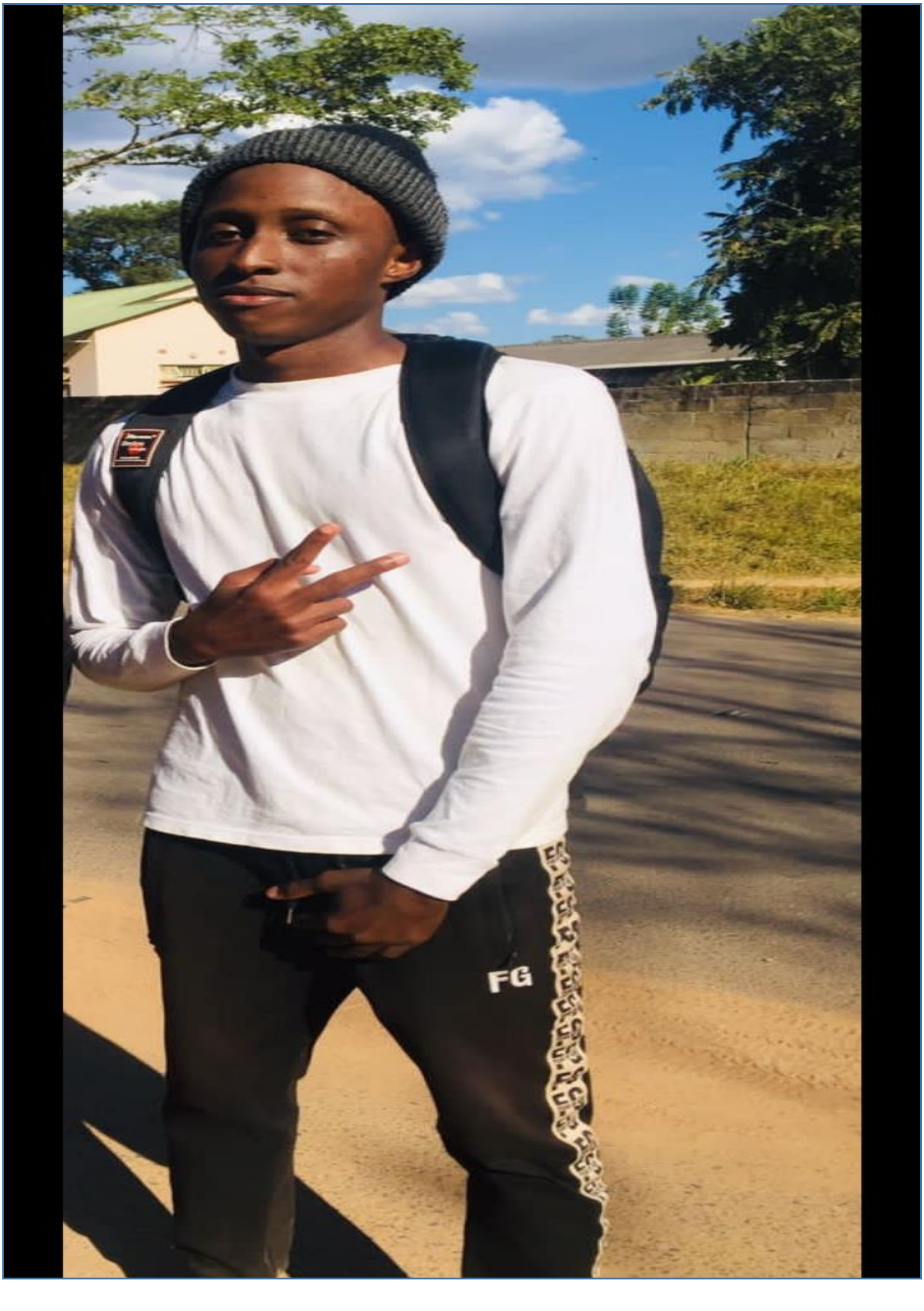

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