

11U CHEMISTRY EXAM REVIEW QUESTIONS (Part 1)
UNITS 1 -3

- Classify the following as chemical or physical properties of matter.

a) silver does tarnish	d) Oxygen supports combustion
b) gold is a conductor of electricity	e) Mercury metal can evaporate
c) potassium iodide dissolves in water	f) Lead has a high density
- Identify each of the changes below as chemical (C) or physical (P).

a) twisting copper wire into a coil	c) melting wax
b) burning coal	d) fermentation of grapes to make wine
- Classify each of the following as a pure substance (PS), a solution (S), or a mechanical mixture (MM).

a) plain Jell-0	d) Raisin Bran cereal
b) neon gas	e) calcium carbonate
c) vinegar	f) air
- What is the name for the smallest particle of a covalently-bonded compound ? _____
 - Only pure substances are homogeneous, true or false?
 - Explain your answer to b).
- J.J Thomson discovered the _____ while studying _____ rays
 - Very briefly describe Thomson's model of the atom.
- Rutherford did an experiment that disproved Thomson's model of the atom. In his experiment Rutherford shot a beam of _____ particles at a thin sheet of _____. Most of these particles _____ and this proved that 99.99% of the atom consists of _____. However, about 1 in 20,000 of the particles were _____ and this proved the existence of a very dense _____ in the atom. Based on the fact that the positive charge on an atom is a multiple of that of hydrogen, Rutherford concluded that this structure contained a subatomic particle called the _____. Later experiments by _____ discovered the last major subatomic particle found in this region, the _____.
- Consider the following atoms : $^{16}_8\text{X}$, $^{17}_9\text{Y}$, $^{17}_8\text{Z}$
Which of these atoms (if any) are isotopes of the same element? Explain your reasoning.
- Draw Bohr-Rutherford diagrams for the following isotopes of each element.
 - Sodium-23 ($^{23}_{11}\text{Na}$)
 - Phosphorus-31 ($^{31}_{15}\text{P}$)
- Use the information given below to calculate the average atomic mass of element X.

ISOTOPE	ABUNDANCE (%)	MASS (u or g/mol)
^{302}X	12.64	302.04
^{304}X	18.23	304.12
^{306}X	69.13	305.03

10. Element X is composed of isotopes X-56 (mass 56.0 u) and X-59 (mass 59.0 u). What is the % abundance of each of these isotopes if the atomic mass is 57.3 u?

11. On the periodic table label the following groups of elements: metals, nonmetals, transition metals, halogens, noble gases, alkali metals, metalloids and the inner transition metals.

12. How many valence electrons are there in:

- a) calcium b) fluorine

13. Which of the following ions would you expect to, be stable?

- Mg^{2+} K^- B^{3-} S^{2-}

14. Consider the following electron configurations for neutral atoms:

- i) 2, 8, 7 ii) 2, 8 iii) 2, 8, 8 iv) 2, 8, 8, 2 v) 2, 8, 8, 1

- a) Which of these would you expect to be an alkaline earth metal? _____
 b) List the five atoms in order of increasing atomic radii. _____
 c) Which of these would you expect to have the lowest ionization energy? _____
 d) Which of these would you expect to be in the same family? _____

15. All of the trends on the periodic table can be explained by two factors :

- i) the number of occupied Bohr energy levels (orbits)
 ii) the number of protons in the nucleus (nuclear charge)

Answer each of the following using these 2 factors.

- a) Circle which is bigger and state why . Li or Na
 b) Circle which is bigger and state why . O or F
 c) Circle which is bigger and state why . Na^+ or Ne
 d) Circle which has the higher ionization energy and state why. S or Cl
 e) Circle which has the greater electron affinity and state why. Cl or Br
 f) Circle which has the greatest metallic character and state why. Be or Mg

16. Write the ionization equations for each of the following:

- a) calcium b) phosphorus

17. Explain the difference between covalent, polar covalent and ionic bonds.

18. Determine the bond type in the following compounds using electronegativity values.

- a) BrCl b) NI₃ c) Al₂O₃ d) Cs₂S

19. Draw a Lewis structure and name the shape for these compounds.

Compound Name	Formula	Lewis Structure	Name of 3D Shape
carbon tetrachloride			
nitrogen			
nitrogen trifluoride			
hydrogen peroxide			
boron trihydride			

20. Complete the following table.

Formula	Name	Name	Formula
PbS		nitric acid	
SnSO ₃ •6H ₂ O		plumbic iodide	
SF ₆		neon	
Na ₂ O ₂		iron (III) hydrogen sulfite	
Al(IO ₄) ₃		chlorine	
NaBrO		potassium bromite	
Ca(HCO ₃) ₂		copper (II) sulfate nonahydrate	
MgO		manganese (II) hypochlorite	
KH ₂ PO ₃		hydrosulfuric acid	
HClO ₂ (aq)		BaO ₂	
KNO ₂		manganese dioxide	
Fe(OH) ₂ (ous/ic)		sulfuric acid	
NH ₄ OH		silver sulfide	
Mg ₃ P ₂		lead (IV) perchlorate	
Au(OH) ₃		magnesium phosphite	

21. Balance the following reactions.

- a) $\text{Ca}_3(\text{PO}_4)_2 + \text{SiO}_2 + \text{C} \rightarrow \text{P}_4 + \text{CaSiO}_4 + \text{CO}$
b) $\text{Al}_2\text{C}_6 + \text{H}_2\text{O} \rightarrow \text{Al}(\text{OH})_3 + \text{C}_2\text{H}_2$
c) $\text{FeCl}_2 + \text{KMnO}_4 + \text{HCl} \rightarrow \text{FeCl}_3 + \text{KCl} + \text{MnCl}_2 + \text{H}_2\text{O}$
d) $\text{BiO}_2 + \text{H}_2 \rightarrow \text{Bi} + \text{H}_2\text{O}$
e) $\text{FeS} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3 + \text{SO}_2$

22. For each pair of reactants, predict if a reaction will occur. If a reaction occurs, complete the word equation. If no reaction occurs, write NR.

- a) tin (s) + potassium iodide (aq)
b) fluorine (aq) + sodium chloride (aq)
c) sodium (s) + calcium sulfate (aq)

23. a) Complete these reactions and write a balanced chemical equation.
b) Write the type of reaction (synthesis, decomposition, single displacement, double displacement, combustion, neutralization) beside each.

- a) water \rightarrow
b) silver + copper (II) sulfate \rightarrow
c) carbon dioxide + water \rightarrow
d) barium chloride + sodium sulfate \rightarrow
e) butane (C_4H_{10}) + oxygen \rightarrow
f) sodium hydroxide + nitric acid \rightarrow

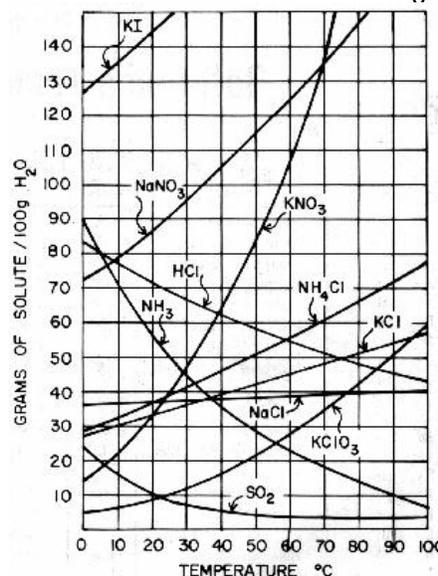
24. In an experiment designed to compare the reactivity of 4 metals (W, X, Y and Z), the metallic elements were reacted with aqueous solutions containing their ions and an anion A. The following observations were made when the single displacement reactions were attempted:

- a) $\text{Z (s)} + \text{WA (aq)} \rightarrow \text{ZA (aq)} + \text{W (s)}$
b) $\text{W (s)} + \text{XA (aq)} \rightarrow \text{NR}$
c) $\text{X (s)} + \text{ZA (aq)} \rightarrow \text{XA (aq)} + \text{Z (s)}$
d) $\text{W (s)} + \text{YA (aq)} \rightarrow \text{WA (aq)} + \text{Y (s)}$

11U CHEMISTRY EXAM REVIEW (Part 2)
UNITS 4 – 6

1. Explain the difference between each pair of terms.
 - a) limiting and excess reactant
 - b) actual yield and theoretical yield
 - c) empirical and molecular formula
2. Calculate the percent composition of sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$.
3. A compound contains 56.6% potassium, 8.7% carbon and 34.7% oxygen. Determine the empirical formula for the compound.
4. Hydroquinone is an organic compound commonly used as a photographic developer. It has a molecular mass of 110 g/mol and a composition of $\text{C}_3\text{H}_3\text{O}$. Calculate the molecular formula of hydroquinone.
5. Use this equation for all three problems:
$$2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$$
 - a) How many moles of water are produced when 5.00 moles of oxygen are consumed?
 - b) If 3.00 moles of water are produced, how many grams of oxygen must be consumed?
 - c) How many moles of hydrogen must be used, given the data in problem b)?
 - d) What mass of water is produced when 105 g of hydrogen reacts with excess oxygen?
6. What mass of sodium carbonate must be used to produce 10.36 L of carbon dioxide at 24°C and 103 kPa according to the following neutralization reaction?
$$\text{HCl}(\text{aq}) + \text{Na}_2\text{CO}_3(\text{s}) \rightarrow \text{NaCl}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$$
7. If 129.0 g of oxygen gas and 300 g of propane are mixed and allowed to react as shown below, determine the volume of water vapour formed at 116 kPa and 120°C .
$$\text{C}_3\text{H}_8(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$$
8. One of the reactions used in the smelting of copper ores to produce copper and sulfur dioxide involves reacting copper(I) oxide with copper(II)sulfide. When 250 kg of copper (I) oxide is heated with 129 kg of copper (II) sulfide, 285 kg of copper is recovered.
$$\text{Cu}_2\text{O}(\text{s}) + \text{Cu}_2\text{S}(\text{s}) \rightarrow \text{Cu}(\text{s}) + \text{SO}_2(\text{g})$$
 - a) Write the balanced chemical equation for the reaction.
 - b) Determine the limiting reagent.
 - c) Calculate the theoretical yield using stoichiometry.
 - d) Determine the percentage yield of copper.
9. A crystal is added to a solution. Describe what happens to the crystal if the solution is unsaturated, saturated or supersaturated.

10. Consider the solubility curves.
- What is the solubility of KNO_3 at 50°C ?
 - What must the temperature be to create a saturated solution of KNO_3 using 50 g of the salt and 50 g of water?
 - What mass of KNO_3 will precipitate if a saturated solution of KNO_3 at 60°C is cooled to 10°C ?
11. 25 g of sodium chloride is dissolved in 100 g of water. What is the concentration in:
- % by mass
 - mol / L (the density of the solution is 1.15 g/mL)



12. What mass of lead (II) nitrate is required to make 4.5 L of a 0.75 mol/L solution?
13. Concentrated phosphoric acid has a concentration of 18.0 mol/L. What volume of concentrated phosphoric acid is needed to make 8.00 L of a 1.50 mol/L solution?
14. The water in a swimming pool has tested positive for lead at a concentration of 4.8 ppm. What mass of lead would there be in a 60,000 L pool? ($D_{\text{H}_2\text{O}} 1.00 \text{ g/mL} = 1.00 \text{ kg/L}$.)
15. Calculate the concentration (in mol/L) of a 7.50% by mass acetic acid solution if the density is 1.02 g/mL. Acetic acid has the formula CH_3COOH .
16. Write a balanced dissociate equation for these salts.
- KCl(s)
 - $\text{Ba(NO}_3)_2 \text{ (s)}$
 - $\text{(NH}_4)_2\text{SO}_3 \text{ (s)}$
17. Complete balanced chemical equations for the following precipitation reactions. Include the state symbols s or aq.
- $\text{Fe(NO}_3)_2 \text{ (aq)} + \text{NaOH (aq)} \rightarrow$
 - $\text{AgNO}_3 \text{ (aq)} + \text{MgCl}_2 \text{ (aq)} \rightarrow$
18. Complete a balanced chemical equation, total ionic equation and net ionic equation.
- $$\text{Pb(NO}_3)_2 \text{ (aq)} + \text{KI (aq)} \rightarrow$$
19. Complete balanced neutralization reactions.
- $\text{H}_3\text{PO}_4 \text{ (aq)} + \text{NaOH (aq)} \rightarrow$
 - $\text{HI (aq)} + \text{Na}_2\text{CO}_3 \text{ (aq)} \rightarrow$
20. Describe the Arrhenius and Bronsted-Lowry definitions of an acid and a base.

21. Write an equation that demonstrates how pure water can act as both a Bronsted Lowry acid and Bronsted Lowry base at the same time.
22. List 3 properties of acids and 3 properties of bases.
23. Distinguish between strong acids (e.g. hydrochloric acid) and weak acids (e.g. acetic acid)
24. What is the pH of the following solutions?
 - a) 0.05 mol/L HCl
 - b) 0.00005 mol/L HNO₃
25. What is the hydronium concentration of a hydrobromic acid solution with a pH of 0.7?
26. What volume of 0.520 mol/L sodium hydroxide are needed to neutralize 100 mL of a 7.50 mol/L hydrobromic acid solution? Include a balanced chemical equation.
27. 10.0 mL of 0.150 mol/L sulfuric acid is titrated with 24.5 mL of sodium hydroxide. Calculate the concentration of the base. Include a balanced chemical equation.
28. Sketch and label a heating curve for benzene (mp = 5.5°C; bp = 80.1°C).
29. A sample of gas has a volume of 54.0 L at 22°C. What will the volume be if it is cooled to -50°C?
30. The pressure on 925 mL of a gas is 120 kPa. What must the pressure be changed to in order to make the volume 0.700 L?
31. A 50.0 L volume of gas is at 42°C and 87 kPa. Calculate the temperature required to change the volume to 42.5 L at a pressure of 770 mmHg.
32. A student collects 300 mL of hydrogen over water at 105.7 kPa and 25°C. How many moles of the gas were collected? Be sure to correct for vapour pressure.
33. Calculate the number of:
 - a) moles in 45.2 L of carbon dioxide gas at STP
 - b) molecules in 45.2 L of carbon dioxide gas at STP
 - c) moles in 45.2 L of oxygen gas at STP.
 - d) moles in 45.2 L of carbon dioxide gas at 25°C and 97.0 kPa.
 - e) 45.2 g of carbon dioxide gas at STP
 - f) 45.2 g of oxygen gas at STP.
34. What is the density of methane gas (CH₄) at 20°C and 115 kPa?
35. If 1.00 L of nitrogen gas is reacted with 3.0 L of fluorine gas, how many litres of nitrogen trifluoride gas will be produced, assuming pressure and temperature remain constant?
36. A student collected 245 mL of an unknown gas X over water at an atmospheric pressure of 108.6 kPa and a temperature of 20°C. If the mass of the gas is 0.7802 g, determine the molar mass.

EXAM REVIEW ANSWERS: Parts 1

1. C, P, P, C, P, P 2. P, C, P, C 3. S, PS, S, MM, PS, S 4. Molecule
5. electron, cathode 6. nucleus, alpha (helium nuclei), gold, passed through, empty space, deflected, nucleus, proton, Chadwick, neutron
7. X and Z; same atomic # but different mass # 8. Na^{11}_{12} 2, 8, 1 P^{15}_{16} 2, 8, 5
9. 304.49 u 10. 57% X-56; 43% X-59 11. See *Chemistry 11* textbook back cover
12. 2, 7 13. Mg^{2+} and S^{2-} are stable ions (they have a full valence shell)
14. a) iv b) smallest ii < iii < I < iv < v largest c) v d) ii and iii are both noble gases
15. a) Na (more energy levels) b) O (less nuclear charge)
c) Ne (same electronic arrangement; lower nuclear charge)
d) Cl (more nuclear charge) e) Cl (electrons closer to nucleus since fewer energy levels)
f) Mg (weaker hold on electrons since more energy levels)
16. a) $\text{Ca} \rightarrow \text{Ca}^{2+} + 2\text{e}^-$ b) $\text{P} + 3\text{e}^- \rightarrow \text{P}^{3-}$
17. Covalent: equal sharing of a pair of electrons (no partial charges or dipoles); Polar covalent: unequal sharing of a pair of electrons (have bond dipoles); Ionic: electrostatic attraction between charged ions.
18. a) 0.2; C b) 0.5 C/PC c) 2.0, I d) 1.8, I
19. CCl_4 , tetrahedral; N_2 , linear; NF_3 , trigonal pyramidal; H_2O_2 , angular; BH_3 , trigonal planar
- 20.

Formula	Name	Name	Formula
PbS	Lead (II) sulfide	nitric acid	HNO_3
$\text{SnSO}_3 \cdot 6\text{H}_2\text{O}$	Tin (II) sulfite hexahydrate	plumbic iodide	PbI_4
SF_6	Sulfur hexafluoride	Neon gas	$\text{Ne}_{(g)}$
Na_2O_2	Sodium peroxide	iron (III) hydrogen sulfite	$\text{Fe}(\text{HSO}_3)_2$
$\text{Al}(\text{IO}_4)_3$	Aluminum periodate	Chlorine gas	$\text{Cl}_{2(g)}$
NaBrO	Sodium hypobromite	potassium bromite	KBrO_2
$\text{Ca}(\text{HCO}_3)_2$	Calcium hydrogen carbonate	copper (II) sulfate nonahydrate	$\text{CuSO}_4 \cdot 9\text{H}_2\text{O}$
MgO	Magnesium oxide	manganese (II) hypochlorite	$\text{Mn}(\text{ClO})_2$
KH_2PO_3	Potassium dihydrogen phosphate	hydrosulfuric acid	$\text{H}_2\text{S}_{(aq)}$
$\text{HClO}_2_{(aq)}$	Chlorous acid	BaO_2	Barium peroxide
KNO_2	Potassium nitrite	manganese dioxide	MnO_2
$\text{Fe}(\text{OH})_2_{(ous/ic)}$	ferrous hydroxide	sulfuric acid	$\text{H}_2\text{SO}_4_{(aq)}$
NH_4OH	Ammonium hydroxide	silver sulfide	Ag_2S
Mg_3P_2	Magnesium phosphide	lead (IV) perchlorate	$\text{Pb}(\text{ClO}_4)_4$
$\text{Au}(\text{OH})_3$	Gold (III) hydroxide	magnesium phosphite	$\text{Mg}_3(\text{PO}_3)_2$

21. a) 2,6,10 → 1, 6, 10 b) 1,6 → 2,3 c) 3,1,8 → 5,1,1,4 d) 1,3 → 2,3 e) 4, 7 → 2, 4
22. a) NR b) → sodium fluoride + chlorine c) → calcium + sodium sulfate
23. a) $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$ (decomp) b) $2\text{Ag} + \text{CuSO}_4 \rightarrow \text{Ag}_2\text{SO}_4 + \text{Cu}$ (single displ)
 c) $\text{CO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{CO}_3$ (syn) d) $\text{BaCl}_2 + \text{Na}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + 2\text{NaCl}$ (double displ)
 e) $\text{C}_4\text{H}_{10} + 15/2 \text{O}_2 \rightarrow 5\text{CO}_2 + 5\text{H}_2\text{O}$ (combust)
 f) $\text{NaOH} + \text{HNO}_3 \rightarrow \text{NaNO}_3 + \text{H}_2\text{O}$
24. $X > Z > W > Y$

EXAM REVIEW ANSWERS: Parts 2

1. see notes 2. 29.1% Na; 40.6% S, 30.3% O 3. K_2CO_3 4. $\text{C}_6\text{H}_6\text{O}_2$
5. a) 10.0 mol b) 1.50 mol c) 3.00 mol d) 937 g 6. 45.8 g 7. 91.0 L
8. a) 2, 1 → 6,1 b) $\text{LR} = \text{Cu}_2\text{S}$ c) 309 kg Cu d) 92.2% yield
9. unsat. = disappears; no change = sat.; grows = supersat.
10. a) 84g/100g water b) 57°C c) 84 g precipitates out (all answers approximate)
11. a) 20% m/m b) 3.9 mol/L 12. 1.1 kg 13. 0.67 L 14. 288 g 15. 1.28mol/L
16. a) $\text{KCl} \rightarrow \text{K}^+ + \text{Cl}^-$ b) $\text{Ba}(\text{NO}_3)_2 \rightarrow \text{Ba}^{2+} + 2\text{NO}_3^-$ c) $(\text{NH}_4)_2\text{SO}_3 \rightarrow 2\text{NH}_4^+ + \text{SO}_3^{2-}$
17. a) $\rightarrow 2\text{NaNO}_3(\text{aq}) + \text{Fe}(\text{OH})_2(\text{s})$ b) $\rightarrow 2\text{AgCl}(\text{s}) + \text{Mg}(\text{NO}_3)_2(\text{aq})$
18. a) $\text{Pb}(\text{NO}_3)_2(\text{aq}) + 2\text{KI}(\text{aq}) \rightarrow \text{PbI}_2(\text{s}) + 2\text{KNO}_3(\text{aq})$
 b) $\text{Pb}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + 2\text{K}^+(\text{aq}) + 2\text{I}^-(\text{aq}) \rightarrow \text{PbI}_2(\text{s}) + 2\text{K}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq})$
 c) $\text{Pb}^{2+}(\text{aq}) + 2\text{I}^-(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$
19. a) $\rightarrow \text{Na}_3\text{PO}_4(\text{aq}) + 3\text{H}_2\text{O}$ b) $\rightarrow 2\text{NaI}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
20. - 22 Refer to notes. 23. Strong acids dissociate 100%; weak acids do not.
24. a) 1.30 b) 4.3 25. 0.20 mol/L 26. 1.44 L 27. 0.122 mol/L
28. Refer to notes. 29. 40.8 L 30. 159 kpa 31. 316 K 32. 0.0124mol
33. a) 2.02 mol b) 1.22×10^{24} molecules c) 2.02 mol d) 1.77 mol e) 1.02 mol f) 1.41 mol
34. 0.757 g/L 35. 2 L 36. 73 g/mol

Studying Suggestions

- Create a word list for each unit
- Create a formula sheet for the course
- Create problem solving maps for major problems (e.g. limiting reactant questions)
- Make corrections to your unit tests and review multiple choice errors
- Study only 1 or 2 units at one sitting; take breaks while studying to avoid burning out
- Do the practice exam (available on-line at IRHS Science SC3U → Exam Review)
- Get lots of sleep and be sure to eat breakfast the morning of the exam
- AVOID excessive caffeine (e.g. coffee, Red Bull) the night before or day of the exam
- AVOID pulling an “all nighter”!