Study Guide

For Placement into Chemistry 30 (CHEM 182)

Academic Upgrading Placement Test | Version: 9 Jan 2025

Important Information

The Chemistry Placement test is a free assessment designed for Academic Upgrading placement purposes only. No section of the test may be used for admission to any SAIT program other than Academic Upgrading. The Chemistry Placement Test is not accepted for admission to any other institution.

- The passing mark required for eligibility to register in CHEM 182 (Chemistry 30) is 60%.
- We aim to put students' passing marks on our system within 2 business days of successful completion of the test.
- Students who have been accepted into the Academic Upgrading program can register for the course they are placed into once we have granted them permission based on their passing grades.
- Students who have already taken and passed SAIT's Academic Upgrading courses in Math and Physics ARE NOT required to take a placement test.

Chemistry Placement Study Guide

This study guide is designed to prepare students for the Academic Upgrading Chemistry Placement test for entry into CHEM 182 (Chemistry 30). Please use the following practice material from Chemistry 20 to prepare for your online placement test to meet eligibility for CHEM 182. An answer key is included at the end of this guide.

This test is for placement into Chemistry 30 equivalency (CHEM 182):

- The study guide consists of 49 questions for practice. The actual test will consist of 20 questions. You may use the formula sheet, acid-base indicators and strengths, solubility table and periodic table of elements as provided at the end of the guide.
- **A data booklet including the periodic table will be provided.** Students will require a pen and paper for the calculation-based questions.
- Students should allow for 60 minutes to complete the test. An additional 30 minutes has been added to allow for accommodated time, for a total test time of 90 minutes.

- A passing mark of 60% or greater is required in this test for eligibility to register in CHEM 182.
- This test is to be written in the Testing Centre.
- You may choose to utilize the Chemistry Study Guide from the Calgary Public Library or bookstore for additional review.

SAIT Academic Upgrading Course Sequence

APSC 180
Science Prep

→

CHEM 181
Chemistry 20
Equivalent

→

CHEM 182
Chemistry 30
Equivalent

Copyright information

Chemistry I – CHEM 181 (Chemistry 20)

Course Content

Below is presented a list of the learning modules used to achieve the learning outcome(s) for this course….

1. Fundamentals of Chemistry and Matter

Learning Outcome:

Explain the basic components of chemistry.

Objectives:

1.1 Explain that the goal of science is knowledge about the natural world.

1.2 Explain that scientific knowledge and theories develop through hypotheses, the collection of evidence, investigation and the ability to provide explanations.

1.3 Explain that scientific knowledge is subject to change as new evidence becomes apparent and as laws and theories are tested and subsequently revised, reinforced or rejected.

1.4 Use appropriate International System of Units (SI) notation, fundamental and derived units and significant digits.

1.5 Convert between units using dimensional analysis.

1.6 Outline the properties of matter, including states and physical changes, components of mixtures, atoms, elements and compounds, and conservation of energy and mass.

1.7 Write empirical definitions of metals and non-metals.

1.8 Identify atoms and ions, charges, families, periods, representative elements, and transitional metals by using the periodic table.

1.9 Apply accepted IUPAC symbols and names of elements.

1.10 Define electron, proton, neutron, nucleus, atomic number, isotope, mass number, ion, cation, and anion.

2. The Diversity of Matter and Chemical Bonding

Learning Outcome:

Describe the role of modelling, evidence and theory in explaining and understanding the structure, chemical bonding and properties of ionic compounds.

Objectives:

2.1	Recall principles for assigning names to ionic compounds.
2.2	Explain why formulas for ionic compounds refer to the simplest whole-number ratio of ions that result in a net charge of zero.
2.3	Define valence electron, electronegativity, ionic bond and intramolecular force.
2.4	Use the periodic table and electron dot diagrams to support and explain ionic bonding theory.
2.5	Explain how an ionic bond results from the simultaneous attraction of oppositely charged ions.
2.6	Explain that ionic compounds form lattices and that these structures relate to the compounds’ properties; e.g., melting point, solubility, reactivity.

Describe the role of modelling, evidence and theory in explaining and understanding the structure, chemical bonding and properties of molecular substances.

Objectives:

2.7	Recall principles for assigning names to molecular substances.
2.8	Explain why formulas for molecular substances refer to the number of atoms of each constituent element.
2.9	Relate electron pairing to multiple and covalent bonds.
2.10	Draw electron dot diagrams of atoms and molecules, writing structural formulas for molecular substances and using Lewis structures to predict bonding in simple molecules.
2.11	Apply VSEPR theory to predict molecular shapes for linear, angular (V-shaped, bent), tetrahedral, trigonal pyramidal and trigonal planar molecules.
2.12	Illustrate, by drawing or by building models, the structure of simple molecular substances.
2.13	Explain intermolecular forces, London (dispersion) forces, dipole-dipole forces and hydrogen bonding.
2.14	Relate properties of substances (e.g., melting and boiling points, enthalpies of fusion and vaporization) to the predicted intermolecular bonding in the substances.
2.15	Determine the polarity of a molecule based on simple structural shapes and unequal charge distribution.

2.16 Describe bonding as a continuum ranging from complete electron transfer to equal sharing of electrons.

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3. Forms of Matter: Gases

Learning Outcome:

Explain molecular behaviour, using models of the gaseous state of matter.

Objectives:

3.1 Describe and compare the behaviour of real and ideal gases in terms of kinetic molecular theory.

3.2 Convert between the Celsius and Kelvin temperature scales.

3.3 Explain the law of combining volumes.

3.4 Illustrate how Boyle's and Charles's laws, individually and combined, are related to the ideal gas law ($PV = nRT$)

a) express pressure in a variety of ways, including units of kilopascals, atmospheres and millimetres of mercury

b) perform calculations, based on the gas laws, under STP, SATP and other defined conditions

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4. Matter as Solutions, Acids and Bases

Learning Outcome:

Investigate solutions, describing their physical and chemical properties.

Objectives:

4.1 Recall the categories of pure substances and mixtures and explain the nature of homogeneous mixtures.

4.2 Provide examples from living and nonliving systems that illustrate how dissolving substances in water is often a prerequisite for chemical change.

4.3 Explain dissolving as an endothermic or exothermic process with respect to the breaking and forming of bonds.

4.4 Differentiate between electrolytes and nonelectrolytes.

4.5 Express concentration in various ways; i.e., moles per litre of solution, percent by mass and parts per million.

4.6 Calculate, from empirical data, the concentration of solutions in moles per litre of solution and determine mass or volume from such concentrations.

4.7 Calculate the concentrations and/or volumes of diluted solutions and the quantities of a solution and water to use when diluting.

4.8 Define solubility and identify related factors; i.e., temperature, pressure and miscibility.

4.9 Explain a saturated solution in terms of equilibrium; i.e., equal rates of dissolving and crystallization.

4.10 Test for the formation of precipitates using a solubility table while recognizing factors, such as temperature, which affect the table's values.

4.11 Describe the procedures and calculations required for preparing and diluting solutions.

Describe acidic and basic solutions qualitatively and quantitatively.

Objectives:

4.12 Recall International Union of Pure and Applied Chemistry (IUPAC) nomenclature of acids and bases.

4.13 Recall the empirical definitions of acidic, basic and neutral solutions determined by using indicators, pH and electrical conductivity.

4.14 Calculate the pH of strong acids.

4.15 Use appropriate SI units to communicate the concentration of solutions and express pH and concentration answers to the correct number of significant digits; i.e., use the number of decimal places in the pH to determine the number of significant digits of the concentration.

4.16 Define Arrhenius (modified) acids as substances that produce H3O+ (aq) in aqueous solutions and recognize that the definition is limited.

4.17 Define Arrhenius (modified) bases as substances that produce OH- (aq) in aqueous solutions and recognize that the definition is limited.

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5. Quantitative Relationships in Chemical Changes

Learning Outcome:

Explain how balanced chemical equations indicate the quantitative relationships between reactants and products involved in chemical changes.

Objectives:

5.1 Predict the product(s) of a chemical reaction based upon the reaction type.

5.2 Recall the balancing of chemical equations in terms of atoms, molecules and moles.

5.3 Contrast quantitative and qualitative analysis.

5.4 Write balanced ionic and net ionic equations, including identification of spectator ions, for reactions taking place in aqueous solutions.

5.5 Calculate the quantities of reactants and/or products involved in chemical reactions, using gravimetric, solution or gas stoichiometry.

Use stoichiometry in quantitative analysis.

Objectives:

5.6 Explain chemical principles (i.e., conservation of mass in a chemical change), using quantitative analysis.

5.7 Identify limiting and excess reagents in chemical reactions.

5.8 Define theoretical yields and actual yields.

5.9 Explain the discrepancy between theoretical and actual yields.

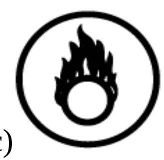
Self Assessment

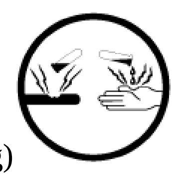
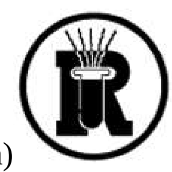
This assessment is only meant to give students an idea of what the questions will look like on the placement test. Note that the test questions will be in multiple choice format. Refer to the objectives to ensure you study all topic areas.

1. Classify each of the following as homogeneous or heterogeneous mixtures:
   a) blood b) chocolate chip cookies c) dissolved "Kool-Aid"
   d) homogenized milk e) bronze

2. Classify the following changes as physical or chemical:
   a) food spoils b) an icicle melts
   c) a nail rusts d) oil is pumped out of a well
   e) an egg is fried f) salt dissolves in water
   g) a window is broken

3. Identify the following lab safety symbols.

a)  b)  c)  d) 

e)  f)  g)  h) 

4. Explain the difference between Material Safety Data Sheets (MSDS), supplier labels, and workplace labels.

5. Explain the difference between endothermic and exothermic reactions. Include the effect of bond energies.

6. Describe the different types of intermolecular forces.

7. Fill in the following chart. Replace the X's with the proper atomic symbol. Note: some are ions, some are not.

Symbol	Atomic Mass	Atomic number	Protons	Neutrons	Electrons
			37	37	37
X^{2-}	16			8	
Cl^{1-}	35	17			
Ne	21				

8. Compare the particles that make up an atom by charge, mass and location in the atom.

9. Predict the monatomic ions formed by the following elements:
a) hydrogen b) magnesium c) sulfur d) iodine e) aluminum

10. Balance the following equations. If they need no coefficients, write "balanced."

a) $CF_{4(1)} \rightarrow C_{(s)} + F_{2(g)}$

b) $H_2SO_{4(aq)} + KOH_{(aq)} \rightarrow KHSO_{4(aq)} + H_2O_{(l)}$

c) $ZnCl_{2(aq)} + H_{2(g)} \rightarrow Zn_{(s)} + HCl_{(aq)}$

d) $SO_{2(g)} + H_2O_{(l)} + O_{2(g)} \rightarrow H_2SO_{4(aq)}$

e) $Li_{(s)} + H_2O_{(l)} \rightarrow LiOH_{(aq)} + H_{2(g)}$

f) $Cu_2O_{(aq)} + Cu_2S_{(aq)} \rightarrow Cu_{(s)} + SO_{2(g)}$

g) $Na_2SO_{4(aq)} + BaCl_{2(aq)} \rightarrow BaSO_{4(s)} + NaCl_{(aq)}$

h) $CH_3OH_{(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

11. Classify each of the above according to the 5 types of reactions. If no classification fits, write "other."

12. Write the formula for each material correctly and then balance the equation. For each reaction, tell what type of reaction it is. For some reactions, you will need to determine the products.

a) Sulfur trioxide and water combine to make sulfuric acid.
b) Lead (II) nitrate and sodium iodide react to make lead iodide and sodium nitrate.
c) Calcium fluoride and sulfuric acid make calcium sulfate and hydrogen fluoride (Hydrofluoric acid).
d) Calcium carbonate will come apart when you heat it to leave calcium oxide and carbon dioxide.
e) Propane burns (with oxygen).

f) Sodium hydroxide neutralizes carbonic acid.
g) Zinc sulfide and oxygen become zinc oxide and sulfur.
h) Lithium oxide and water make lithium hydroxide.
i) Aluminum hydroxide and sulfuric acid neutralize to make water and aluminum sulfate.
j) Zinc and copper (II) sulfate yield zinc sulfate and copper metal.
k) Barium hydroxide and sulfuric acid make water and barium sulfate.
l) Aluminum sulfate and calcium hydroxide become aluminum hydroxide and calcium sulfate.
m) Copper metal and silver nitrate react to form silver metal and copper (II) nitrate.
n) Sodium metal and chlorine react to make sodium chloride.
o) Calcium phosphate and sulfuric acid make calcium sulfate and phosphoric acid.

13. Describe the difference between ionic and molecular compounds. You may answer in point form.

14. Classify solids, liquids and gases according to the 3 types of motion in kinetic molecular theory.

15. Classify each of the following as ionic or molecular, and give the correct chemical name for each:

a) NI_3	g) $Zn(CH_3COO)_2$	m) $CaCrO_4$
b) NaI	h) SnF_4	n) CH_3OH
c) FeO	i) CrI_3	
d) K_2S	j) $LiCl \bullet 4H_2O$	
e) $GaBr_3$	k) P_2O_5	
f) S_4N_2	l) NH_3	

16. Give the correct chemical formula for each of the following:

1. hydrochloric acid	17. sodium sulfate decahydrate	33. lithium phosphate
2. sodium chloride	18. rubidium nitrite	34. lead (II) acetate – 3 – water
3. sodium hexafluoride	19. lead (II) sulfite	35. nitrogen dioxide
4. strontium nitrate	20. copper (I) sulfide	36. iron (III) oxide
5. calcium chloride	21. aluminum oxide	37. sodium peroxide
6. acetic acid	22. magnesium bromide	38. copper (II) oxide
7. phosphoric acid	23. sodium chlorate	39. liquid nitrogen

8. ammonia
9. chlorine
10. lithium sulfate
11. potassium chromate
12. calcium hydroxide
13. aluminum foil
14. ammonium sulfate
15. sulfuric acid
16. ammonium iodide
24. iron (II) chloride
25. hydrogen gas
26. silver chromate
27. zinc bicarbonate
28. barium oxide
29. aluminum nitrate
30. diphosphorus pentoxide
31. aluminum hydroxide
32. chromium (III) oxide
40. phosphorus pentafluoride
41. lead IV fluoride
42. iodine tribomide
43. carbonic acid
44. silver bisulfite
45. cupric hydroxide
46. nitric acid
47. mercury (II) bromide
48. hydrofluoric acid

17. Calculate the molecular weight (molar mass) of the following. Round to 2 decimal places.
a) CO_2 b) $Si(CH_3)_4$ c) $Na_2SO_4•6H_2O$

18. Predict the simplest molecular formula and write a balanced equation for the following combination reactions. Name the product. Note: the element listed first comes first in the name, too.
a) P_4 + F_2 →
b) H_2 + Br_2 →
c) C + Cl_2 →
d) Si + S_8 →

19. For the following combinations of solutions, predict if a precipitate will form. If it will, write a balanced equation showing the formation of the precipitate. If it won't, write N.R. (no reaction).
a) $CaCl_{2(aq)}$ + $AgNO_{3(aq)}$ →
b) $NaOH_{(aq)}$ + $Mg(NO_3)_{2(aq)}$ →
c) $NaCl_{(aq)}$ + $AgNO_{3(aq)}$ →
d) $Na_2SO_{4(aq)}$ + $Ba(NO_3)_{2(aq)}$ →
e) $NaCl_{(aq)}$ + $KNO_{3(aq)}$ →
f) $NaCl_{(aq)}$ + $Pb(NO_3)_{2(aq)}$ →

20. Problems on concentration and solution-making:
a) Explain how to make up five litres of a 0.175 M NaCl solution.
b) What volume of 0.86 M table sugar ($C_{12}H_{23}O_{12}$) has 50 grams of sugar in it?
c) How many grams of $KMnO_4$ would you get if you evaporated the water from 85.75 mL of 1.27 M solution?
d) To what volume must you dilute 15 grams of silver nitrate to make it 0.05 M?
e) What is the concentration of KCl if five grams of it are in 25.3 L?
f) How many moles of chlorine gas are in 17 L of 1.02 M solution?

g) How many grams of sulfuric acid are in 5 mL of 3.2 M acid?
h) I made up 500 ml of 0.1 M sodium hydroxide solution. Explain how I did it.
i) What is the concentration of silver nitrate if 15 grams of it are dissolved into 14.28 litres?

21. What type of electrons form bonds, paired, or unpaired?

22. Write and balance the chemical equation for those problems that need it. Show all your work.
    a) Sodium hydroxide and hydrochloric acid combine to make table salt and water. 14 mL of 0.1 M sodium hydroxide is added to an excess of acid. How many moles of table salt are made? How many grams of salt is that?
    b) 50 mL of 0.25 M copper II sulfate evaporates to leave $CuSO_4 \cdot 5H_2O$. (That is the pentahydrate crystal of copper II sulfate.) What is the mass of this beautiful blue crystal from the solution?
    c) Chlorine gas is bubbled into 100 mL of 0.25 M potassium bromide solution. This produces potassium chloride and bromine gas. The bromine (which dissolves in water) is taken from the solution and measured at 27˚C and 825 mmHg. What is the volume of bromine?
    d) 95 mL of 0.55 M sulfuric acid is put on an excess of zinc. This produces zinc sulfate and hydrogen. How many grams of zinc sulfate are made?
    e) 27.6 mL of a 0.19 M solution of silver nitrate and 15.4 mL of an unknown (but excess) amount of sodium chloride combine to make a white precipitate silver chloride and some dissolved sodium nitrate. (i) How many moles of silver chloride are made? (ii) How many grams of silver chloride is that? (iii) How many moles of sodium nitrate are made?

23. A solution of a sodium salt contains unknown anion(s) which may be Br^-, SO_4^{2-}, or CO_3^{2-}. Explain how to determine which anions are present.

24. A metal cylinder contains one mole of nitrogen gas at STP. What will happen to the pressure if another mole of gas is added to the cylinder but the temperature and volume do not change?

25. A gas in a container has a pressure of 510 kPa at 27˚C. What will the pressure be if the temperature is lowered to -173˚C?

26. A given mass of air has a volume of 3.0 L at 101 kPa. What volume will it occupy at 25.3 kPa if the temperature does not change?

27. 2.5 L of air at -50˚C are warmed to 100˚C. What is the new volume if the pressure remains constant?

28. A 5.0 L sample at a temperature of -50˚C has a pressure of 107 kPa. What will be the new pressure if the temperature is raised to 100˚C and the volume expands to 7.0 L? Answer to three significant figures.

29. What volume will 36.0 g of oxygen gas occupy at 25˚C and a pressure of 52.7 kPa?

30. Calculate the number of litres occupied at STP of:
    a) 2.5 mol of N_2
    b) 0.600 g of H_2
    c) 0.350 mol O_2

31. What pressure will be exerted by 0.450 mol of a gas at 25˚C if it is contained in a vessel whose volume is 6.5 L?

32. Determine the volume occupied by 0.582 mol of a gas at 15˚C if the pressure is 82.9 kPa.

33. No gas exhibits ideal behaviour at all temperatures and pressures. Explain the meaning of this statement.

34. Explain what a mole ratio is and where a person should look to find a particular mole ratio. Give an example.

35. Explain the difference between a coefficient number and a subscript number.

36. After determining the balanced equation, what is the first step in most stoichiometry problems?

37. Silicon dioxide reacts with carbon to form carbon monoxide and silicon monocarbide. What mass of carbon will react with 1.772 g of silicon dioxide?

38. Calcium oxide and water combine to form calcium hydroxide. If 4.0 g of CaO and 7.0 g of water are available,
    i) What is the limiting reagent?
    ii) What amount of calcium hydroxide can be formed?
    iii) If only 4.6 g are formed, what is the percent yield?

39. If 0.504 g of ammonium dichromate is decomposed according to:
    $(NH_4)_2Cr_2O_7 \rightarrow Cr_2O_3 + 4H_2O + N_2$
    what is the mass of chromium (III) oxide, Cr_2O_3, that would be obtained at the end of the experiment?

40. Yellow phosphorus (P_4) combines directly with chlorine gas to form phosphorus pentachloride. What mass of phosphorus pentachloride will be formed using 14.1 g of phosphorus and 20.0 L of chlorine at STP?

41. In the atmosphere, the air pollutant nitrogen dioxide reacts with water to produce nitric acid according to the unbalanced equation:
$NO_{2(g)} + H_2O_{(l)} \rightarrow HNO_{3(aq)} + NO_{(g)}$
What volume of NO_2 will react with 36.04 g of water at SATP?

42. Hydrogen sulfide combines with oxygen to form water and sulfur dioxide. If 3.41 g of hydrogen sulfide is mixed with 6.40 g of oxygen, what mass of sulfur dioxide will be produced?

43. When 21.0 g of tin metal was reacted with 30.0 g of chlorine according to the equation:
$Sn + 2\,Cl_2 \rightarrow SnCl_4$
the yield of tin(IV)chloride obtained was 37.6 g. What is the percent yield of this reaction?

44. Given that 37.5 mL of 0.100 M aluminum bromide solution reacts with silver nitrate solution, what is the mass of silver bromide precipitate?

45. In nature, copper is often found combined with sulfur in the ore chalcocite, Cu_2S. If 5.20 kg should theoretically be produced and only 3.60 kg are actually produced, what is the percent yield?

46. Given that 27.5 mL of 0.210 lithium iodide reacts completely with 0.133 M lead(II)nitrate solution, what volume of $Pb(NO_3)_2$ is required to complete precipitation?

47. List the four theoretical rules for determining the number and occupancy of valence orbitals.

48. Show the electron energy levels of the following. For the valence level, indicate paired and unpaired electrons.
    a) Mg
    b) O
    c) K
    d) C

49. For the following, indicate how many bonds each element can form.
    a) Cs
    b) Se
    c) C
    d) Sr

ANSWER KEY

1. a) heterogeneous b) heterogeneous c) homogeneous
   d) heterogeneous e) homogeneous

2. a) chemical b) physical c) chemical d) physical
   e) chemical f) physical g) physical

3. a) Class A, Compressed gas
   b) Class B, Flammable and combustible material
   c) Class C, Oxidizing material
   d) Class D, Materials causing immediate and serious toxic effects
   e) Class D, Materials causing other toxic effects
   f) Class D Biohazardous Infectious material
   g) Class E Corrosive material
   h) Class F Dangerously reactive material

4. MSDS have detailed information on physical and chemical properties, short and long-term effects, first aid, storage, handling and disposal of the chemical. Supplier labels contain hazard symbol(s), product identifier, risks, precautionary measures, first aid, supplier address, and reference to the MSDS. A "hatched" border is a key visual indicator of a controlled product and hence a hazardous chemical. Workplace labels only contain a product identifier, safe handling information and a reference to the MSDS.

5. Endothermic reactions remove energy from the surroundings, while exothermic reactions release energy into the surroundings. Bond energy is the energy required to break a chemical bond. It is also the energy released when a bond is formed. Overall, if a reaction is endothermic, the energy required to break the bonds of the reactants is greater than the energy released when the products are formed. Likewise, if a reaction is exothermic, the energy required to break the bonds of the reactants is less than the energy released when the products are formed.

6.

Type of force	Relative Strength	Description/Definition of Force
London	relatively weak	These are attractive forces that result when negative electrons in one molecule are attracted by the positive nuclei of atoms in nearby molecules.
Polar	usually weaker than London forces	Many covalent compounds do not share electrons equally between the two atoms. If one atom has higher electronegativity, the atom will pull the bonding electrons more closely to it than to the other atom. Polar molecules tend to line up so that the slightly positive end is near the slightly negative end of a nearby molecule.
Hydrogen bonding	relatively strong	Hydrogen bonds are special, relatively strong dipole-dipole forces between molecules containing F-H, O-

		H and N-H bonds. There are two parts to the current theory of hydrogen bonds. First, there is a large difference in electronegativities of fluorine, oxygen, and nitrogen when compared to hydrogen. This produces highly polar bonds. Second, the small size of the hydrogen atoms means that the positive pole is highly "concentrated," and therefore exerts a strong attraction on the negative pole of a nearby molecule.
Metallic	weak	A metallic solid can be pictured as a three-dimensional array of positive ions that remain fixed in the crystal lattice while the loosely held valence electrons move freely throughout the crystal.

7.

Symbol	Atomic Mass	Atomic number	Protons	Neutrons	Electrons
Rb	**74**	**37**	37	37	37
$\mathbf{O^{2-}}$	16	**8**	**8**	8	10
Cl^{1-}	35	17	**17**	**18**	**18**
Ne	21	**10**	**10**	**11**	**10**

8.

Particle	**Charge**	**Smallest to largest by mass**	**Location in atom**
Proton	+1	middle	inside nucleus
Neutron	0	largest	inside nucleus
Electron	-1	smallest	outside nucleus

9. a) H^+ b) Mg^{2+} c) S^{2-} d) I^- e) Al^{3+}

10.

e) $CF_{4(1)} \rightarrow C_{(s)} + 2F_{2(g)}$

f) $H_2SO_{4(aq)} + KOH_{(aq)} \rightarrow KHSO_{4(aq)} + H_2O_{(l)}$ balanced

g) $ZnCl_{2(aq)} + H_{2(g)} \rightarrow Zn_{(s)} + 2HCl_{(aq)}$

h) $SO_{2(g)} + H_2O_{(l)} + \frac{1}{2}O_{2(g)} \rightarrow H_2SO_{4(aq)}$ balanced

i) $2Li_{(s)} + 2H_2O_{(l)} \rightarrow 2LiOH_{(aq)} + H_{2(g)}$

j) $2Cu_2O_{(aq)} + Cu_2S_{(aq)} \rightarrow 6Cu_{(s)} + SO_{2(g)}$

k) $Na_2SO_{4(aq)} + BaCl_{2(aq)} \rightarrow BaSO_{4(s)} + 2NaCl_{(aq)}$

l) $2CH_3OH_{(l)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 4H_2O_{(g)}$

11. a) decomposition b) double replacement c) single replacement
d) formation e) single replacement f) other
g) double replacement h) complete combustion

12. a. SO_3 + H_2O → H_2SO_4 Formation
b. Pb(NO3)2 + 2NaI → PbI2(s) + 2NaNO3 Double Replacement
c. CaF2 + H2SO4 → CaSO4(s) + 2 HF Double Replacement
d. CaCO3 → CaO + CO2 Decomposition
e. C3H8 + 5 O2 → 4 H20 + 3 CO2 Combustion
f. 2 NaOH + H2CO3 → Na_2CO_3 + 2 H20 Double Replacement
g. 2 ZnS + O2 → 2 ZnO + 2 S Single Replacement
h. Li2O + H2O → 2 LiOH Formation
i. 2 Al(OH)3 + 3 H2SO4 → 6 H2O + Al2(SO4)3 Double Replacement
j. Zn + CuSO4 → ZnSO4 + Cu Single Replacement
k. Ba(OH)2 + H2SO4 → 2 H2O + BaSO4 Double Replacement
l. Al2(SO4)3 + 3 Ca(OH)2 → 2 Al(OH)3(s) + 3 CaSO4(s)
Double Replacement (both calcium sulfate and aluminum hydroxide are precipitates)
m. Cu + 2 AgNO3 → 2 Ag + Cu(NO3)2 Single Replacement
n. 2Na + Cl2 → 2 NaCl Formation
o. Ca3(PO4)2 + 3 H2SO4 → 3 CaSO4 + 2 $H3PO_4$ Double Replacement

13.

	Made up of:	**State at SATP**	**Conductivity in water**	**Theoretically**
Ionic	non-metal with metal	solid	conductive	electron(s) are transferred from one atom to another
Molecular	non-metal with non-metal	solid, liquid or gas	non-conductive	electrons are shared between atoms

14. Solids – mostly vibrational; Liquids – translational, rotational and vibrational; Gases – mostly translational

15.

a) molecular - nitrogen triiodide
b) ionic – sodium iodide
c) ionic – iron(II) oxide
d) ionic – potassium sulfide
e) ionic – gallium bromide
f) molecular – tetrasulfur dinitride
g) ionic – zinc acetate
h) ionic – tin (IV) fluoride
i) ionic – chromium (III) iodide
j) ionic – lithium chloride tetrahydrate
k) molecular – diphosphorus pentaoxide
l) molecular - ammonia
m) ionic –calcium chromate
n) molecular - methanol

16.

1. HCl
2. NaCl
3. NaF_6
4. $Sr(NO_3)_2$
5. $CaCl_2$
6. HCH_3COO
7. H_3PO_4
8. NH_3
9. Cl_2
10. Li_2SO_4
11. K2CrO4
12. Ca(OH)2
13. Al
14. $(NH_4)_2SO_4$
15. H_2SO_4
16. NH_4I
17. $Na_2SO_4 \bullet 10H_2O$
18. $RbNO_2$
19. $PbSO_3$
20. Cu_2S
21. Al_2O_3
22. $MgBr_2$
23. $NaClO_3$
24. $FeCl_2$
25. H_2
26. Ag_2CrO_4
27. $Zn(HCO_3)_2$
28. BaO
29. $Al(NO_3)_3$
30. P_2O_5
31. $Al(OH)_3$
32. Cr_2O_3
33. Li_3PO_4
34. $Pb(CH_3COO)_2 \bullet 3H_2O$
35. O_2
36. Fe_2O_3
37. Na_2O_3
38. CuO_2
39. N_2
40. PF_5
41. PbF_4
42. IBr_3
43. H_2CO_3
44. $AgHSO_3$
45. $Cu(OH)_2$
46. HNO_3
47. $HgBr_2$
48. HF

17. a) 44.01 g/mol
b) 88.25 g/mol
c) 250.16 g/mol

18.

a)	P_4 +	$6F_2$	$\rightarrow$	$4PF_3$	phosphorus trifluoride
b)	H_2 +	Br_2	$\rightarrow$	2HBr	hydrogen bromide
c)	C +	$2Cl_2$	$\rightarrow$	CCl_4	carbon tetrachloride
d)	4Si +	S_8	$\rightarrow$	$4SiS_2$	silicon disulfide

19.

a)	$CaCl_{2(aq)}$	+	$2\ AgNO_{3(aq)}$	$\rightarrow$	$2\ AgCl_{2(s)}$	+	$Ca(NO_3)_{2(aq)}$
b)	$2\ NaOH_{(aq)}$	+	$Mg(NO_3)_{2(aq)}$	$\rightarrow$	$Mg(OH)_{2(s)}$	+	$2\ NaNO_{3(aq)}$
c)	$NaCl_{(aq)}$	+	$AgNO_{3(aq)}$	$\rightarrow$	$AgCl_{(s)}$	+	$NaNO_{3(aq)}$
d)	$Na_2SO_{4(aq)}$	+	$Ba(NO_3)_{2(aq)}$	$\rightarrow$	$BaSO_{4(s)}$	+	$2\ NaNO_{3(aq)}$
e)	$NaCl_{(aq)}$	+	$KNO_{3(aq)}$	$\rightarrow$	N.R.		
f)	$2\ NaCl_{(aq)}$	+	$Pb(NO_3)_{2(aq)}$	$\rightarrow$	$PbCl_{2(s)}$	+	$2\ NaNO_{3(aq)}$

20. a) (i) Weigh out 51.2 grams of NaCl. (ii) Dissolve the solid in a small amount of water in a suitable volumetric device. (iii) Bring the solution to volume by adding water (q.s.) and mix to completely disburse.
b) 0.162 L

c) 17.2 g

d) 1.77 L

e) 2.65 m mols

f) 17.34 mols

g) 1.57

h) (i) Weigh out 2.00 grams of NaOH. (ii) Dissolve the solid in a small amount of water in a suitable volumetric device. (iii) Bring the solution to volume by adding water (q.s.) and mix to completely disburse.

i) 6.18 m molar

21. Unpaired.

22. a) 1.4×10^{-3} mols, 0.0819 g b) 3.12 g c) 284 ml
d) 8.44 g e) i) 5.24 E-3 mols ii) 0.752 g iii) 5.24×10^{-3} mols

23. One possible test would be: Add $Be(NO_3)_2$. If a precipitate forms, CO_3^{2-} is present. Continue to add $Be(NO_3)_2$ until no more precipitate is formed (so all the carbonate is used up). Second: add calcium or barium nitrate. If a precipitate forms, SO_4^{2-} is present. Third: add copper (I) nitrate. If a precipitate forms, Br^- is present.

24. It will double.

25. 170 kPa

26. 12 L

27. 4.2 L

28. 128 kPa

29. 52.8 L

30. a) 56 L b) 6.65 L c) 7.84 L

31. 170 kPa

32. 16.8 L

33. At low temperatures and/or high pressures:
- Gases may become liquids.
- The volume of molecules becomes significant not negligible.
- Ideal gas laws will not accurately predict behaviour.

34. The ratio of moles of reactants and products according to the coefficients in the balanced chemical equation. For example, in $2H_2 + O_2 \rightarrow 2H_2O$, the ratio is 2:1:2.

35. Coefficients are in front of each chemical formula, and indicate number of moles of the whole formula. Subscripts are part of the formula, and indicate numbers of atoms of each element in each molecule.

$$4\,Al + 3\,O_2 \rightarrow 2\,Al_2O_3$$

coefficients subscripts

36. Find the number of moles present.

37. 1.062 g of C

38. i) CaO is the limiting reagent. ii) 5.3 g of $Ca(OH)_2$ can be formed. iii) 87% yield.

39. 0.304 g of Cr_2O_3.

40. 74.4 g of PCl_5.

41. 148.8 L of NO_2.

42. 6.41 g of SO_2.

43. 81.6%

44. 2.11 g of AgBr

45. 69.2%

46. 21.7 mL $Pb(NO_3)_2$

47. There are four valence orbitals. An orbital may contain 0, 1 or 2 electrons. Electrons occupy empty orbitals before forming electron pairs. A maximum of 8 electrons can occupy orbitals in the valence level.

48.

a) Mg	b) O	c) K	d) C
$1e^-$, $1\ e^-$		$1e^-$	
$8\ e^-$	$2\ e^-$, $2\ e^-$, $1\ e^-$, $1\ e^-$,	$8\ e^-$	$1\ e^-$, $1\ e^-$, $1\ e^-$, $1\ e^-$
$2\ e^-$	$2\ e^-$	$8\ e^-$	$2\ e^-$
$12p^+$	$8p^+$	$2\ e^-$	$6p^+$
		$19p^+$	

49. a) 1 bond b) 2 bonds c) 4 bonds d) 2 bonds

Data Booklet
Placement test for entry into Chemistry 30 (CHEM 182)

Common Polyatomic Ions			
Ion	**Name**	**Ion**	**Name**
CH_3COO^-	acetate	H_3O^+	hydronium
NH_4^+	ammonium	OH^-	hydroxide
$C_6H_5COO^-$	benzoate	ClO^-	hypochlorite
BO_3^{3-}	borate	IO_3^-	iodate
C_2^{2-}	carbide	Hg_2^{+2}	mercury(I)
CO_3^{2-}	carbonate	NO_3^-	nitrate
ClO_3^-	chlorate	NO_2^-	nitrite
ClO_2^-	chlorite	$OOCCOO^{-2}$	oxalate
CrO_4^{2-}	chromate	ClO_4^-	perchlorate
CN^-	cyanide	MnO_4^-	permanganate
$Cr_2O_7^{2-}$	dichromate	O_2^{2-}	peroxide
$H_2PO_4^-$	dihydrogen phosphate	S_2^{2-}	persulfide
HCO_3^-	hydrogen carbonate (bicarbonate)	PO_4^{3-}	phosphate
$HOOCCOO^-$	hydrogen oxalate	PO_3^{3-}	phosphite
HPO_4^{2-}	hydrogen phosphate	SiO_3^{2-}	silicate
HSO_4^-	hydrogen sulfate (bisulfate)	SO_4^{2-}	sulfate
HSO_3^-	hydrogen sulfite (bisulfite)	SO_3^{2-}	sulfite
HS^-	hydrogen sulphide (bisulfide)	SCN^-	thiocyanate
		$S_2O_3^{2-}$	thiosulfate

Solubility of Some Common Ionic Compounds in Water at 25ºC								
Ion	**Group1 NH_4^+ H_3O^+,H^+**	**ClO_3^- NO_3^- ClO_4^-**	**CH_3COO^-**	**Cl^- Br^- I^-**	**SO_4^{2-}**	**S^{2-}**	**OH^-**	**PO_4^{3-} SO_3^{2-} CO_3^{2-}**
Solubility greater than or equal to 0.1 mol/L (very soluble)	all	all	most	most	most	Group1 Group2 NH_4^+	Group1 NH_4^+ Sr^{2+} Ba^{2+} Tl^+	Group1 NH_4^+
Solubility less than 0.1 mol/L (slightly soluble)	none	none	Ag^+ Hg^+	Ag^+ Pb^{2+} Hg^+ Cu^+ Tl^+	Ca^{2+} Sr^{2+} Ba^{2+} Ra^{2+} Pb^{2+} Ag^+	most	most	most

Chem 181 Formulas and Constants

1.000 atm= 101.325 kPa = 760.0 mm Hg = 760.0 torr = 14.69 psi

SATP: T= 25ºC and P = 100kPa; molar volume = 24.8 L/mol

STP: T=0ºC and P = 1 atm (101.325kPa); molar volume = 22.4 L/mol

Kelvins = Celsius + 273.15

1 mole = 6.022 x 10^{23}

$c = \frac{n}{v}$ $\quad c_1v_1 = c_2v_2$ $\quad ppm = \frac{m_{solute}}{m_{solvent}} x10^6$

$P_1V_1 = P_2V_2$ $\quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$ $\quad \frac{V_1}{n_1} = \frac{V_2}{n_2}$

$PV=nRT$ $\quad \frac{V_1P_1}{n_1T_1} = \frac{V_2P_2}{n_2T_2}$

$pH = - \log [H_3O^+]$ $\quad [H_3O^+] = 10^{-pH}$ $\quad pOH = - \log [OH^-]$ $\quad [OH^-] = 10^{-pOH}$

R = 8.31 kPa·L·mol^{-1}·K^{-1} = 0.08206 atm·L·mol^{-1}·K^{-1}

Acid–Base Indicators at 298.15 K

Indicator	Suggested Abbreviations	pH Range	Colour Change as pH Increases	K_a
methyl violet	$HMv_{(aq)}$ / $Mv^-_{(aq)}$	0.0 – 1.6	yellow to blue	~2 × 10^{-1}
cresol red	$H_2Cr_{(aq)}$/ $HCr^-_{(aq)}$ $HCr^-_{(aq)}$ / $Cr^{2-}_{(aq)}$	0.0 – 1.0 7.0 – 8.8	red to yellow yellow to red	~3 × 10^{-1} 3.5 × 10^{-9}
thymol blue	$H_2Tb_{(aq)}$/ $HTb^-_{(aq)}$ $HTb^-_{(aq)}$/ $Tb^{2-}_{(aq)}$	1.2 – 2.8 8.0 – 9.6	red to yellow yellow to blue	2.2 × 10^{-2} 6.3 × 10^{-10}
orange IV	$HOr_{(aq)}$/ $Or^-_{(aq)}$	1.4 – 2.8	red to yellow	~1 × 10^{-2}
methyl orange	$HMo_{(aq)}$ / $Mo^-_{(aq)}$	3.2 – 4.4	red to yellow	3.5 × 10^{-4}
bromocresol green	$HBg_{(aq)}$/ $Bg^-_{(aq)}$	3.8 – 5.4	yellow to blue	1.3 × 10^{-5}
methyl red	$HMr_{(aq)}$ / $Mr^-_{(aq)}$	4.8 – 6.0	red to yellow	1.0 × 10^{-5}
chlorophenol red	$HCh_{(aq)}$ / $Ch^-_{(aq)}$	5.2 – 6.8	yellow to red	5.6 × 10^{-7}
bromothymol blue	$HBb_{(aq)}$ / $Bb^-_{(aq)}$	6.0 – 7.6	yellow to blue	5.0 × 10^{-8}
phenol red	$HPr_{(aq)}$ / $Pr^-_{(aq)}$	6.6 – 8.0	yellow to red	1.0 × 10^{-8}
phenolphthalein	$HPh_{(aq)}$/ $Ph^-_{(aq)}$	8.2 – 10.0	colourless to pink	3.2 × 10^{-10}
thymolphthalein	$HTh_{(aq)}$/ $Th^-_{(aq)}$	9.4 – 10.6	colourless to blue	1.0 × 10^{-10}
alizarin yellow R	$HAy_{(aq)}$/ $Ay^-_{(aq)}$	10.1 – 12.0	yellow to red	6.9 × 10^{-12}
indigo carmine	$HIc_{(aq)}$ / $Ic^-_{(aq)}$	11.4 – 13.0	blue to yellow	~6 × 10^{-12}
1,3,5–trinitrobenzene	$HNb_{(aq)}$ / $Nb^-_{(aq)}$	12.0 – 14.0	colourless to orange	~1 × 10^{-13}

Table of Common Polyatomic Ions

acetate (ethanoate)	CH_3COO^-	chromate	CrO_4^{2-}	phosphate	PO_4^{3-}
ammonium	NH_4^+	dichromate	$Cr_2O_7^{2-}$	hydrogen phosphate	HPO_4^{2-}
benzoate	$C_6H_5COO^-$	cyanide	CN^-	dihydrogen phosphate	$H_2PO_4^-$
borate	BO_3^{3-}	hydroxide	OH^-	silicate	SiO_3^{2-}
carbide	C_2^{2-}	iodate	IO_3^-	sulfate	SO_4^{2-}
carbonate	CO_3^{2-}	nitrate	NO_3^-	hydrogen sulfate	HSO_4^-
hydrogen carbonate (bicarbonate)	HCO_3^-	nitrite	NO_2^-	sulfite	SO_3^{2-}
		oxalate	$OOCCOO^{2-}$	hydrogen sulfite	HSO_3^-
perchlorate	ClO_4^-	hydrogen oxalate	$HOOCCOO^-$	hydrogen sulfide	HS^-
chlorate	ClO_3^-	permanganate	MnO_4^-	thiocyanate	SCN^-
chlorite	ClO_2^-	peroxide	O_2^{2-}	thiosulfate	$S_2O_3^{2-}$
hypochlorite	ClO^- or OCl^-	persulfide	S_2^{2-}		

Legend for Elements

Solid	*Liquid*	Gas
Natural	Synthetic	

Note: The legend denotes the physical state of the elements at exactly 101.325 kPa and 298.15 K.

Key

26 55.85 3+, 2+ 1.8 Fe iron

Atomic number: 26; Atomic molar mass (g/mol)*: 55.85; Common ion charges (most common first): 3+, 2+; Electronegativity: 1.8; Symbol: Fe; Name: iron

* Based on $^{12}_{6}C$

( ) Indicates mass of the most stable isotope

Periodic Table

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 1.01 1+,1− 2.2 H hydrogen																	2 4.00 — — −269 −272† He helium
3 6.94 1+ 1.0 Li lithium	4 9.01 2+ 1.6 Be beryllium											5 10.81 — 2.0 B boron	6 12.01 — 2.6 C carbon	7 14.01 3− 3.0 N nitrogen	8 16.00 2− 3.4 O oxygen	9 19.00 1− 4.0 F fluorine	10 20.18 — — Ne neon
11 22.99 1+ 0.9 Na sodium	12 24.31 2+ 1.3 Mg magnesium											13 26.98 3+ 1.6 Al aluminium	14 28.09 — 1.9 Si silicon	15 30.97 3− 2.2 P phosphorus	16 32.07 2− 2.6 S sulfur	17 35.45 1− 3.2 Cl chlorine	18 39.95 — — Ar argon
19 39.10 1+ 0.8 K potassium	20 40.08 2+ 1.0 Ca calcium	21 44.96 3+ 1.4 Sc scandium	22 47.87 4+, 3+ 1.5 Ti titanium	23 50.94 5+, 4+ 1.6 V vanadium	24 52.00 3+, 2+ 1.7 Cr chromium	25 54.94 2+, 4+ 1.6 Mn manganese	26 55.85 3+, 2+ 1.8 Fe iron	27 58.93 2+, 3+ 1.9 Co cobalt	28 58.69 2+, 3+ 1.9 Ni nickel	29 63.55 2+, 1+ 1.9 Cu copper	30 65.39 2+ 1.7 Zn zinc	31 69.72 3+ 1.8 Ga gallium	32 72.64 4+ 2.0 Ge germanium	33 74.92 3− 2.2 As arsenic	34 78.96 2− 2.6 Se selenium	35 79.90 1− 3.0 *Br* bromine	36 83.80 — — Kr krypton
37 85.47 1+ 0.8 Rb rubidium	38 87.62 2+ 1.0 Sr strontium	39 88.91 3+ 1.2 Y yttrium	40 91.22 4+ 1.3 Zr zirconium	41 92.91 5+, 3+ 1.6 Nb niobium	42 95.94 6+ 2.2 Mo molybdenum	43 (97) 7+ 2.1 Tc technetium	44 101.07 3+, 4+ 2.2 Ru ruthenium	45 102.91 3+ 2.3 Rh rhodium	46 106.42 2+, 4+ 2.2 Pd palladium	47 107.87 1+ 1.9 Ag silver	48 112.41 2+ 1.7 Cd cadmium	49 114.82 3+ 1.8 In indium	50 118.71 4+, 2+ 2.0 Sn tin	51 121.76 3+, 5+ 2.1 Sb antimony	52 127.60 2− 2.1 Te tellurium	53 126.90 1− 2.7 I iodine	54 131.29 — 2.6 Xe xenon
55 132.91 1+ 0.8 Cs cesium	56 137.33 2+ 0.9 Ba barium	57-71 lanthanoids	72 178.49 4+ 1.3 Hf hafnium	73 180.95 5+ 1.5 Ta tantalum	74 183.84 6+ 1.7 W tungsten	75 186.21 7+ 1.9 Re rhenium	76 190.23 4+ 2.2 Os osmium	77 192.22 4+ 2.2 Ir iridium	78 195.08 4+, 2+ 2.2 Pt platinum	79 196.97 3+, 1+ 2.4 Au gold	80 200.59 2+, 1+ 1.9 *Hg* mercury	81 204.38 1+, 3+ 1.8 Tl thallium	82 207.21 2+, 4+ 1.8 Pb lead	83 208.98 1.9 Bi bismuth	84 (209) 2+, 4+ 2.0 Po polonium	85 (210) 1− 2.2 At astatine	86 (222) — — Rn radon
87 (223) 1+ 0.7 Fr francium	88 (226) 2+ 0.9 Ra radium	89-103 actinoids	104 (267) Rf rutherfordium	105 (268) Db dubnium	106 (269) Sg seaborgium	107 (270) Bh bohrium	108 (269) Hs hassium	109 (277) Mt meitnerium	110 (281) Ds darmstadtium	111 (282) Rg roentgenium	112 (285) Cn copernicium	113 (286) Nh nihonium	114 (290) Fl flerovium	115 (290) Mc moscovium	116 (293) Lv livermorium	117 (297) Ts tennessine	118 (294) Og oganesson

57 138.91 3+ 1.1 La lanthanum	58 140.12 3+ 1.1 Ce cerium	59 140.91 3+ 1.1 Pr praseodymium	60 144.24 3+ 1.1 Nd neodymium	61 (145) 3+ — Pm promethium	62 150.36 3+, 2+ 1.2 Sm samarium	63 151.96 3+, 2+ — Eu europium	64 157.25 3+ 1.2 Gd gadolinium	65 158.93 3+ — Tb terbium	66 162.50 3+ 1.2 Dy dysprosium	67 164.93 3+ 1.2 Ho holmium	68 167.26 3+ 1.2 Er erbium	69 168.93 3+ 1.3 Tm thulium	70 173.04 3+, 2+ — Yb ytterbium	71 174.97 2+ 1.0 Lu lutetium
89 (227) 3+ 1.1 Ac actinium	90 232.04 4+ 1.3 Th thorium	91 231.04 5+, 4+ 1.5 Pa protactinium	92 238.03 6+, 4+ 1.7 U uranium	93 (237) 5+ 1.3 Np neptunium	94 (244) 4+, 6+ 1.3 Pu plutonium	95 (243) 3+, 4+ — Am americium	96 (247) 3+ — Cm curium	97 (247) 3+, 4+ — Bk berkelium	98 (251) 3+ — Cf californium	99 (252) 3+ — Es einsteinium	100 (257) 3+ — Fm fermium	101 (258) 2+, 3+ — Md mendelevium	102 (259) 2+, 3+ — No nobelium	103 (262) 3+ — Lr lawrencium

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