Unit 0 - Review of Chemistry 11

Day 1:

☐ Significant Figures

Definition: The number of significant figures in a measurement is the number of digits that are accurately known plus on uncertain digit.

Rules:

1) All non-zero digits are significant. (1-9)
2) All zeros between non-zero digits are significant. 105, 305.3 0.104
3) Zeros that place the decimal are not significant. 0.345, 0.00045
4) Zeros that follow non-zero number in a decimal number are significant (0.00400) therefore zeros that follow non-zero numbers in a non-decimal number are not significant. (4200)

Exercise 1: State the number of significant figures in each of the following

a) 306
b) 30 600
c) 0.3060
d) 1.000

Addition/subtraction

1) Perform the operation with the numbers given as usual.
2) Round off the answer to the number of **decimal places** as contained in the **least accurate** number (fewest decimal places)

Exercise 2: Solve.

a) 4.60 + 3 =
b) 67.5 - 0.009 =
c) 200 - 87.3 =
d) 22.4420 + 56.981 =

Multiplication/division

1) Perform the operation with the numbers given as usual.
2) Round off the answer to the number of **significant digits** as contained in the **least accurate** number (fewest ~~decimal places~~ significant digits)

Exercise 3: Solve

a) 3.060 x 2.143 87 =
b) 3.14 x 36.3741 x 8.345
c) $\frac{101.3 \times 6.5384}{8.31 \times 276}$ =

Scientific Notation

Definition: A number expressed in scientific notation consists of a number between 1 and 10 plus a power of 10.

Rules:

1) Scientific notation need not to be used for numbers between 0.1 and 1000.
2) For number less than 0.1 or numbers greater than 1000, the numbers should be expressed as either"
   a) scientific notation plus base unit e.g. 5000 km = 5×10^3

   OR

   b) a number between 0.1 and 1000 plus the proper SI prefix 3000 g = 3 Kg

Exercise 4: Write the following numbers in scientific notation.

a) 175000
b) 0. 000 000 0945
c) 0.069×10^{-8}
d) $\dfrac{(4.0 \times 10^5) \times (6.0 \times 10^{10})}{(3.0 \times 10^{-2})}$

Base Units

Distance -
Mass -
Time -
Temperature -
Mole -

Atomic Number, Mass Number and Isotopes

Atomic Number: Te number of protons in the nucleus of an atom
Atomic Mass: The mass of an atom; the sum of the number of protons and the number of neutrons present in the nucleus.

Example: $^{23}_{11}\mathrm{Na}$ The smaller number is _____ and it is the _____
The larger number is _____ and it is the _____.

Number of Neutrons =

Exercise 5: What is the atomic number and atomic mass of the following:

Symbol	Atomic #	Atomic Mass	# of protons	# of electrons	# of neutrons
$^{16}_{8}\mathrm{O}$					
$^{40}_{19}\mathrm{K}$					
$^{239}_{92}\mathrm{U}$					

Isotope: atoms with different atomic mass which have the same atomic number.
The atoms of different isotopes can still be atoms of the same element.
They differ only in the number of neutrons in the nucleus.

Example: The atomic number of uranium is _____

U-235 The nucleus of U-235 contains _____ protons and _____ neutrons.

U-238 The nucleus of U-238 contains _____ protons and _____ neutrons.

Isotope	# neutrons	# protons	# of electrons
radium-226			
cesium-137			

Most elements as they occur naturally on earth are **mixtures** of several isotopes.
This is why the periodic table has the mass numbers are not nice whole numbers!

Example: Chlorine exists as a mixture of isotopes of approximate atomic mass 35g/mol (relative abundance 75.55%) and 37 g/mol (relative abundance 24.45%). It is safe to assume that no matter where you go the percentage abundance will be the same. What is the average atomic mass of Chlorine?

Atomic Mass = The sum of the percentages of each isotope of the element

=

Exercise 5: the composition of ordinary neon is: neon-20, 90.92%; neon-21, 0.26%; neon-22, 8.82%. Calculate the average atomic mass of neon.

Example: The atomic mass of the isotopes Li-6 and Li-7 in naturally occurring lithium are 6.0151214 and 7.0160030 respectively. The atomic mass of naturally occurring lithium given in the period c table is 6.941 g/mol. What are the value of percentage abundance for each isotope?

Exercise 6: Natural Rb consists solely of the isotopes Rb-85 and Rb-87. From the atomic mass of Rb(85.4678) calculate the relative percentage of these two isotopes.

Balancing Chemical Equations

When a chemical reaction occurs, it can be described by an equation. This shows the chemicals that react (called the _____) on the left-hand side, and the chemical that they produce (called the _____) on the right-hand side.

Steps for Balancing chemical equations

1. Write out the skeleton equation. Ensure that you have copied the entire chemical formulas correctly.
2. Begin by balancing the atoms that occurs in the largest number on either side of the equation. Leave hydrogen, oxygen, and any other elements until later.
3. Balance any polyatomic ions, such as sulphate, SO_4^{-2}, that occur on both sides of the chemical equation as an ion unit. That is, do not split a sulphate ion into 1 sulfur atom and 4 oxygen atoms. Balance this ion as one nit.
4. Balance any hydrogen or oxygen atoms that occur in a combined and uncombined state. For example, combined oxygen might be in the form of CO_2, while uncombined oxygen occurs as O_2
5. Finally, balance any other element that occurs in its uncombined state: for example, Na, Cl_2
6. Check your answer. Count the number of each type f atom on each side of the equation.

The concept of balancing equations

Take a look at this chemical word equation:

aluminum + oxygen → aluminum oxide

Write the skeleton equation:

Balance the chemical equations:

Example: Balance the following equation: $CO_2 + H_2 \rightarrow CH_4 + H_2O$

Introduction to Balancing Equations

Balance the following chemical equations.

1. $Hg + O_2 \longrightarrow HgO$
2. $H_2 + O_2 \longrightarrow H_2O$
3. $Mg + O_2 \longrightarrow MgO$
4. $O_2 + Cl_2 \longrightarrow OCl_2$
5. $Al + F_2 \longrightarrow AlF_3$
6. $KNO_3 \longrightarrow KNO_2 + O_2$
7. $Zn + HCl \longrightarrow ZnCl_2 + H_2$
8. $KClO_3 \longrightarrow KCl + O_2$
9. $KOH + HCl \longrightarrow KCl + H_2O$
10. $Na_2SO_4 + HCl \longrightarrow NaCl + H_2SO_4$
11. $Mg + HCl \longrightarrow MgCl_2 + H_2$
12. $NaOH + HCl \longrightarrow NaCl + H_2O$
13. $CaCO_3 + HCl \longrightarrow CaCl_2 + H_2CO_3$
14. $SiO_2 + HF \longrightarrow SiF_4 + H_2O$
15. $Pb(NO_3)_2 \longrightarrow PbO + NO_2 + O_2$
16. $Cu + H_2SO_4 \longrightarrow CuSO_4 + SO_2 + H_2O$
17. $PbS_2 + O_2 \longrightarrow Pb_3O_3 + SO_2$
18. $MnO_2 + HCl \longrightarrow MnCl_2 + Cl_2 + H_2O$
19. $Al + HCl \longrightarrow AlCl_3 + H_2$
20. $C_{25}H_{42} + O_2 \longrightarrow H_2O + CO_2$
21. Mercury + Oxygen gas $\longrightarrow$ Mercury (II) Oxide
22. Carbon + Oxygen gas $\longrightarrow$ Carbon monoxide
23. Iron + Oxygen gas $\longrightarrow$ Iron (III) Oxide
24. Copper (II) Oxide + Carbon $\longrightarrow$ Copper + Carbon Dioxide
25. Hydrochloric acid + calcium hydroxide $\longrightarrow$ Calcium chloride + water

Balancing Polyatomic Equations

1. $FeCl_3 + (NH_4)_2S \longrightarrow Fe_2S_3 + NH_4Cl$
2. $Al(OH)_3 + HNO_3 \longrightarrow Al(NO_3)_3 + H_2O$
3. $Ba(OH)_2 + Fe_2(SO_4)_3 \longrightarrow Fe(OH)_3 + BaSO_4$
4. $NH_4Cl + Na_3PO_4 \longrightarrow (NH_4)_3PO_4 + NaCl$
5. $Mg(OH)_2 + H_3PO_4 \longrightarrow Mg_3(PO_4)_2 + H_2O$
6. $Na_2CO_3 + H_3PO_4 \longrightarrow Na_3PO_4 + H_2O + CO_2$
7. $Pb(CH_3COO\text{-})_2 + (NH_4)_2S \longrightarrow PbS + CH_3COONH_4$
8. $NaI + MnO_2 + H_2SO_4 \longrightarrow Na_2SO_4 + MnSO_4 + H_2O + I_2$
9. $KMnO_4 + H_2SO_3 \longrightarrow K_2SO_4 + MnSO_4 + H_2SO_4 + H_2O$
10. $Al + H_2SO_4 \longrightarrow Al_2(SO_4)_3 + H_2O + SO_2$

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Day 2:

- Hydrated and Dehydration

Hydrate: A compounds that has a specific number of water molecules bonded to each formula unit. E.g. $MgSO_4 \cdot 7H_2O$

Anhydrous: A compounds that has no water molecules bonded to it. E.g. $CaSO_4$

Formula	Chemical Name
$CaSO_4 \cdot 2H_2O$	Calcium sulphate dehydrate
$LiCl_2 \cdot 4H_2O$	Lithium Chloride tetrahydrate
$Ba(OH)_2 \cdot 8H_2O$	
$Na_2CO_3 \cdot 10H_2O$	

Example: a hydrate of barium hydroxide, $Ba(OH)_2 \cdot x\ H_2O$ is used to make barium salts and to prepare certain organic compounds. Since it reacts with CO_2 from the air to yield barium carbonate, $BaCO_3$. It must be stored in tightly stopper bottles.

(a) a 50.0 g sample of the hydrate contains 27.2 g of $Ba(OH)_2$. Calculate the percent, by mass, of water in $Ba(OH)_2 \cdot x\ H_2O$.

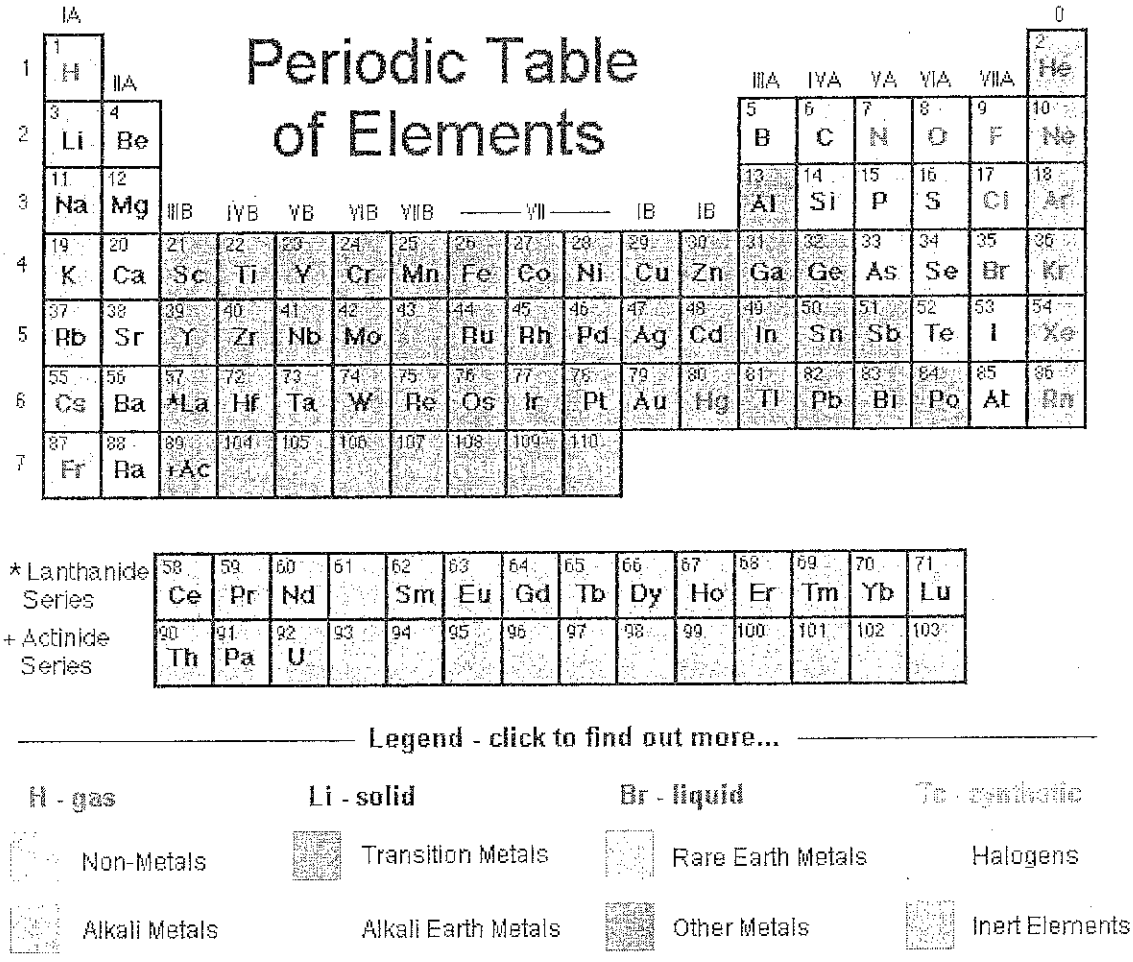
(b) Find the value of x in $(OH)_2 \cdot x\ H_2O$.

Exercise 8:

A 3.34 g sample of a hydrate has the formula $SrS_2O_3 \cdot x\ H_2O$ and contains 2.30 g of SrS_2O_3. Find the value of x.

The Modern Periodic Table

The most important difference between Mendeleev's table and today's table is that the modern table is organized by increasing atomic number, not increasing atomic weight.



Nomenclature – Name and Formulas of Compounds

❖ Ionic Compounds

1) Metal and non-metal where metals form **only one kind of ion**
   - Simplest compound (consists of a metal and a non-metal)
   - Written as: **Metal + non-metal + -ide**
   - Writing the name → NaCl
   - Writing the formula →

   magnesium oxide barium sulphide

2) Metal and non-metal where metals form **more than one kind of ion**.
   - Many transition metals can form more than one kind of ion. These are called **MULTIVALENT**.

Name of element	Symbol	Ionic Charge	Roman Numeral
Copper	Cu	1+, 2+	I, II
Iron	Fe	2+, 3+	II, III
Lead	Pb	2+, 4+	II, IV
Tin	Sn	2+, 4+	II, IV

   - written as: **Metal + roman numeral + non-metal + -ide**
   - Stock Naming System (Writing the Name) → CuCl
     $CuCl_2$
   - Writing the formula →

3) Metal and Polyatomic Ions

- Ions that consists of 2 or more elements are called **POLYATOMIC IONS.**

Name of polyatomic ion	Ion Formula	Ionic Charge
nitrate	NO_3^{-1}	1-
hydroxide	OH^-	1-
chlorate	ClO_3^-	1-
carbonate	CO_3^{-2}	2-
sulphate	SO_4^{-2}	2-
phosphate	PO_4^{-3}	3-

- These types of compounds are written similarly to Binary Ionic Compounds,
- Written as: **Metal + Polyatomic Ion**
- Writing the Name → $BaSO_4$
- Writing the Formula →

silver nitrate

calcium hydroxide

- In naming a compound containing an oxyanion, we start with the metal and end with the name of the oxyanion.

e.g. Cl and O form 4 polyatomic ion with the same charge.

ClO^- ⟶

ClO_2^- ⟶

ClO_3^- ⟶

ClO_4^- ⟶

❖ Molecular Compounds

- A binary compounds made up of 2 non-metals

Subscript	Prefix
1	Mono
2	Di
3	Tri
4	Tetra
5	Penta
6	Hexa
7	Hepta
8	Octa
9	Nona
10	Deca

- written as: **prefix + non-metal + prefix + non-metal + -ide**
- Writing the name → N_2O
- Writing the formula → diphosophorus trioxide
- Hydrogen is an exception to the rule
  e.g. H_2S
- if the first element has only 1 atom, we do not use "mono".
  e.g. CCl_4

❖ Naming Acids

1) Binary Acids
   - written as: Hydro (Stem) ic acid
   - Writing the name → $HCl_{(aq)}$
     $HF_{(aq)}$

2) Acids that are formed with polyatomic ions are names the same way as binary acids
   - Writing the name → $HCN_{(aq)}$

3) Acids that are formed from various combinations of oxyanions with hydrogen.

Anions	acids	Examples:
Ends in -ate	Suffix: -ic	Chlorate anion ClO_3^- Chloric Acid $HClO_3$
Ends in -ite	Suffix: -ous	Chlorite anion ClO_2^- Chlorous Acid $HClO_2$
Prefix hypo- & per-	Remains as part of the acid name	perchlorate anion ClO_4^- perchloric Acid $HClO_4$ hypochlorite anion ClO^- hypochlorous acid $HClO^-$

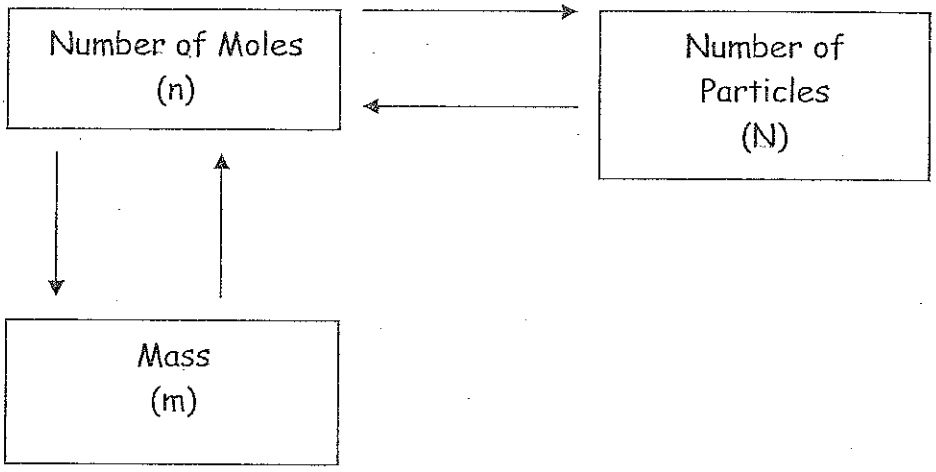
- Writing the name → HNO_2
- Writing the formula → phosphoric acid

❖ Naming Bases

- All aqueous solution of ionic hydroxides are bases
- Metal ions with one or more hydroxide anions.
  e.g. NaOH
  $Ba(OH)_2$

Unit 0 – Grade 11 Review

Conversions (Moles, Number of particles, Molar Mass and Mass)



$$n = \frac{m}{M} \quad \text{and} \quad n = \frac{N}{N_A}$$

Example 1: A sample contains 1.25 mol of nitrogen dioxide, NO_2.

(a) How many molecules are in the sample?

(b) How many atoms are in the sample?

Example 2: How many moles of CH_3COOH are in 23.6 g sample?

Example 3: What is the mass of 5.67×10^{24} molecules of $CoCl_2$?

Percentage Composition:

Definition: Relative mass of each element in the compound.

Example 1: A compound with a mass of 48.72 g is found to contain 32.69 g of Zinc and 16.03 g of Sulfur. What is the percentage composition of the compound?

Example 2: C_9H_8O is responsible for characteristic odour of cinnamon. Determine the % composition of the compound by calculating the mass % of C, H, and O.

Empirical & Molecular Formula:

Name of Compound	Molecular Formula	Empirical Formula	Lowest Ratio of elements
Hydrogen Peroxide	H_2O_2	HO	1:1
Benzene	C_6H_6	CH	1:1

Example 1: calculate the empirical formula of a compound that is 85.6% C and 14.4% hydrogen. The molar mass of the compound is determined to be 70 g/mol. What is the molecular formula?

Example 2: The percentage composition of a fuel is 81.7% carbon and 18.3% hydrogen. Find the empirical formula of the fuel. If the molar mass is 132 g/mol. What is the molecular formula?

Stoichiometry:

Example 1: Determine the number of moles of Oxygen that are needed to react with 0.58 mol of Vanadium to form divanadium penoxide.

Example 2: When carbon dioxide reacts with Lithium hydroxide, it produces lithium carbonate and water. If $1X\ 10^3$ g of CO_2 is used. What mass of LiOH must react to form the products?

The Limiting Reactant:

Example 1: Lithium nitride reacts with water to form ammonia and lithium hydroxide. If 4.87g of Lithium nitride reacts with 5.80 g of H_2O, find the limiting reactant. What mass of ammonia is produced.

Percentage Yield:

$$\text{Percentage Yield} = \frac{Actual Yield}{Theoretical Yield} \times 100\%$$

Example 1: $C_7H_8 + 2KMnO_4 \rightarrow KC_7H_5O_2 + 2MnO_2 + KOH + H_2O$

(a) 8.60 g of C7H8 is reacted with excess KMnO4. What is the theoretical yield, in grams, of $KC_7H_5O_2$.

(b) If the % yield is 70%, what mass of $KC_7H_5O_2$ can be expected?

(c) What mass of C_7H_8 is needed to produce 13.4 g of $KC_7H_5O_2$

Combustion:

Complete Combustion: Reaction of a compound/element with oxygen gas to form CO_2 and H_2O

Incomplete Combustion: Reaction of a compound/element with oxygen gas to form CO, CO_2 and H_2O.

Gas Laws

Boyle's Law: $P_iV_i = P_fV_f$ (@ constant temperature)

Example 1: A 50 cm^3 sample of N_2 is collected at 101.3 kPa. If the volume is reduce to 5 cm^3 and the temperature remains constant, what will the P_f of N_2 be?

Charle's Law: $\frac{V_i}{T_i} = \frac{V_f}{T_f}$ (@ constant pressure; Temp. must be in Kelvins)

Example 2: A balloon is filled with 2.50 L of dry Helium @ 23.5 °C. After the balloon is placed in the freezer, the volume is found to be 2.15 L. What was the temperature (in °C) inside the freezer if the pressure remained constant?

Gay-Lussac's Law: $\frac{P_i}{T_i} = \frac{P_f}{T_f}$ (@ constant temperature; Temp. must be in Kelvins)

Combined Gas Law: $\frac{P_iV_i}{T_i} = \frac{P_fV_f}{T_f}$

Ideal Gas Law: $PV = nRT$

T in Kelvins
V in Litres
n in moles
P in kPa
$R = 9.314 \frac{kPaL}{molK}$

Gas Stoichiometry:

Example 1: Sulfuric acid reacts with iron metal to produce gas and an iron (II) compound. What volume of gas is produced when excess sulfuric acid reacts with 409 g of iron @ 18.0°C and 100.3 kPa?

Example 2: A student reacts magnesium with excess dilute hydrochloric acid to produce hydrogen gas. She uses 0.15 g of magnesium metal. What volume of dry hydrogen does she collect over water @ 28°C and 101.8 kPa?

Assignment 0:

Identify the number of significant figures:

1) 3.0800 2) 0.00418 3) 7.09×10^{-5} 4) 91,600

5) 0.003005 6) 3.200×10^{9} 7) 250 8) 780,000,000

9) 0.0101 10) 0.00800 11) 3.461728 + 14.91 + 0.980001 + 5.2631

12) 23.1 + 4.77 + 125.39 + 3.581 13) 22.101 - 0.9307

14) 0.04216 - 0.0004134 15) 564 321 - 264 321

16) $\dfrac{(3.4617 \times 10^{7})}{(5.61 \times 10^{-4})}$ 17) $\dfrac{(4.7620 \times 10^{-15})}{[(3.8529 \times 10^{12})(2.813 \times 10^{-7})(9.50)]}$

Write the following numbers in scientific notation.

a) 0.0000943597
b) 8318.861
c) 1466
d) 0.0878292
e) 54170100000
f) 9505000
g) 0.0000405809
h) 0.0006418

Balance the chemical equations.

1. $HCl + MnO_2 \longrightarrow Cl_2 + MnCl_2 + H_2O$
2. $BiCl_3 + H_2S \longrightarrow Bi_2S_3 + HCl$
3. $AgNO_3 + H_2S \longrightarrow Ag_2S + HNO_3$
4. $NH_3 + Cl_2 \longrightarrow NH_4Cl + N_2$
5. $N_2O_5 + H_2O \longrightarrow HNO_3$
6. $P_4O_6 + H_2O \longrightarrow H_3PO_3$

7) $AlCl_3 + NH_4OH \longrightarrow Al(OH)_3 + NH_4Cl$

8. $Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$

9. $NaHCO_3 \longrightarrow Na_2CO_3 + CO_2 + H_2O$

10. $Ca_3(PO_4)_2 + H_2SO_4 \longrightarrow H_3PO_4 + CaSO_4$

11. $FeCl_3 + NH_4OH \longrightarrow Fe(OH)_3 + NH_4Cl$

12. $PbS + H_2O_2 \longrightarrow PbSO_4 + H_2O$

Predict the products of the reaction and then balance the chemical equation.

1. Potassium chromate + barium nitrate $\longrightarrow$? + ?

2. Copper metal + silver nitrate $\longrightarrow$? + ?

Write the skeleton equation for each word equation and then balance it.

1. Ammonia gas + oxygen gas $\longrightarrow$ nitrogen (II) oxide + water

2. Aluminum chloride + zinc sulphate $\longrightarrow$ Aluminum sulphate + zinc chloride

3. Write the balanced equation for the reaction of lead (II) nitrate with sodium iodide to form sodium nitrate and lead (II) iodide

Write the formula for the following compound:

1. ammonium sulfide
2. aluminum acetate
3. sodium acetate
4. sodium nitrate
5. calcium chloride dihydrate
6. zinc sulfite
7. barium chromate
8. silver bicarbonate
9. aluminum sulfate
10. cobalt chloride
11. potassium iodide
12. potassium nitrate
13. barium chloride dihydrate
14. lead(IV) chlorite
15. ferrous carbonate
16. mercury(II) nitrate
17. lead(II) phosphate
18. potassium hydroxide
19. lead(II) nitrite
20. zinc sulfite
21. potassium dichromate
22. sodium sulfite
23. magnesium carbonate
24. calcium fluoride
25. calcium bicarbonate
26. nickel nitrate
27. aluminum hydroxide
28. silver cyanide
29. silver phosphate
30. ammonium sulfite
31. sodium hypochlorite
32. zinc sulfate
33. ammonium chromate
35. tin(II) chloride
35. barium carbonate

Write the formula for the following compound:

1. HgF_2
2. KF
3. $NaOH$
4. $Be(ClO_4)_2$
5. $Mg_3(PO_4)_2$
6. Fe_2O_3
7. KCl
8. $CaSO_4$
9. $SnCrO_4$
10. $Hg(OH)_2$
11. $(NH_4)Cr_2O_7$
12. $KMnO_4$
13. K_2O
14. Al_2O_3
15. CaH_2
16. $KClO_4$
17. $SbCl_3$
18. $CuCO_3$
19. $AuCl_3$
20. ZnO
21. Al_2S_3
22. $Ba(OH)_2$
23. NH_4Cl
24. CuS
25. NH_4NO_3
26. MgI_2
27. $Mg_3(PO_4)_2$
28. $CaCO_3$
29. $CoCl_3$
30. $CuSO_3$
31. $Ba_3(PO_4)_2$
32. $NaCN$
33. $FeCl_3$
34. BrO_3

1. The element Silver has two stable isotopes. The lighter has a mass of 106.90509 and the heavier one has a mass of 108.90470. If the average mass of ~~bromine~~ Silver atoms is found in nature to be 107.77 amu, what is the relative abundance of each isotope?

2. A 2.78 g sample of hydrate iron(II) sulphate, $FeSO_4 \cdot xH_2O$ was heated to remove all the water of hydration. The mass of the anhydrous iron (II) sulphate was 1.52 g. Calculate the number of water molecules associated with each formula unit of $FeSO_4$.

3. Convert each quantity to a mass in grams.
(a) 14.8 mol of $BaCrO_4$
(b) 5.58×10^{20} molecules of C_3H_8
(c) 4.27×10^{21} atoms of He

4. How many atoms of H are in a mixture of 3.49×10^{23} molecules of H_2O and 78.1 g of CH_3OH?

5. One molecule of C_2H_5OH requires three molecules of O_2 for complete combustion. What mass of O_2 is required to react completely with 92.0g of C_2H_5OH?

6. Citric acid is composed of carbon, hydrogen and oxygen. When a 0.5000 g sample of citric acid was subjected to carbon-hydrogen combustion analysis, 0.6871 g of carbon dioxide and 0.1874 g of water were produced. Using a mass spectrometer, the molar mass of citric acid was determined to be 192 g/mol.
a. What are the parentages of carbon, hydrogen, and oxygen in citric acid?
b. What is the empirical formula of citric acid?
c. What is the molecular formula of citric acid?

7. A compound has the formula X_2O_5, where X is an unknown element. The compound is 44.0% oxygen by mass. What is the identity of element X?

8. 20.8g of calcium phosphate, 13.3 g of silicon dioxide, and 3.90 g of carbon react according to the following equation:
$2Ca_3(PO_4)_2 + 6SiO_2 + 10\,C \rightarrow P_4 + 6CaSiO_3 + 10CO$
Determine the mass of calcium silicate that is produced.

9. Silicon dioxide reacts with hydrofluoric acid to produce silicon tetrafluoride and water vapour.
(a) 12.2 g of silicon dioxide is reacted with 4.3 g HF. What mass of water is produced.
(b) If the actual yield of water is 4.50 g, what is the percentage yield of the reaction?
(c) Assuming the yield obtained in part (b), what mass of SiF_4 is formed?

10. Ammonia gas, NH_3, is used in the production of fertilizer. At 55°C, sample of ammonia gas is found to exert a pressure of 7.5 atm. What pressure will the gas exert if its volume is reduced one fifth of its original volume at 55°c?

11. Neon gas is widely used as the luminous gas in signs. A sample of neon has a volume of 5.5 L of 750 torr at 10.0°C. If the gas is expanded to a volume of 7.5 L at a pressure of 400 torr, what will its final temperature be in °C?

12. (a) at constant pressure, the temperature, in Kelvin, is doubled. What effect will this have on a gas? Explain.
(b) at constant pressure, the temperature, in degrees Celsius, is doubled. How is this different from the situation in part (a)? How will the effects on a gas be different? Explain.

13. The chemical equation below describes what happens when a match is struck against a rough surface to product light and heat.
$P_4S_{3(s)} + O_2 \rightarrow P_4O_{10} + SO_2$
(a) Balance this chemical equation.
(b) If 5.3 L of oxygen gas were consumed, how many litres of sulphur dioxide would be produced?

14. Nitrogen monoxide, NO, is one of the gases that is responsible for smog. It is produced in various ways, one of which is during the combustion of ammonia.
$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$
If 25.0 L of ammonia reacts with 27.5 L of oxygen at STP, what mass of nitrogen monoxide will be produced?