Chemistry 2202- Unit 1

Solubility of an Ionic Compound

Dissociation

< refers to the separation of ions that occurs when an **ionic compound** dissolves in water.

$$Al_2(SO_4)_{3(s)} \rightarrow 2\ Al^{3+}_{(aq)} + 3\ SO_4^{2-}{}_{(aq)}$$

< the solubility of an ionic compound will determine whether or not it will completely dissociate into ions in solution:

i) high solubility (aq) will completely dissociate into ions;
ii) low solubility (s) will not completely dissociate.

SOLUBILITY OF IONIC COMPOUNDS AT SATP - GENERALIZATIONS

Anion	Cl^- Br^- I^-	S^{2-}	OH^-	SO_4^{2-}	CO_3^{2-} PO_4^{3-} SO_3^{2-}	CH_3COO^-	NO_3^-
High Solubility (aq)	most	Group 1 NH_4^+ Group 2	Group 1 NH_4^+ Sr^{2+} Ba^{2+} Tl^+	most	Group 1 NH_4^+	most	all
Low Solubility (s)	Ag^+ Pb^{2+} Tl^+ Hg_2^{2+} (Hg^+) Cu^+	most	most	Ag^+ Pb^{2+} Ca^{2+} Ba^{2+} Sr^{2+} Ra^{2+}	most	Ag^+	none

All Group 1 compounds, including acids and ammonium compounds, are assumed to have high solubility in water.

Q. Predict which of the following ionic compounds will dissociate and write a dissociation equation.
$CaSO_4$, NH_4OH, Na_2SO_4, $MnCO_3$, $Mg(NO_3)_2$

A: Check solubility using table.

ATOMIC STRUCTURE

Atoms are composed of three basic subatomic particles.

Electron
- negatively charged particle
-almost no mass; 1/1836 the mass of a proton
-found outside the nucleus

Proton
- positively charged particle
-each proton has a mass of 1.67×10^{-24} g

Neutron
-found in the nucleus
- no electric charge (neutral)
-has a mass almost equal to that of a proton
-found in the nucleus

Nucleus
- A small structure located at the centre of the atom
-it is made up of protons and neutrons
-it has all the positive charge and the majority of the mass
-it takes up only 10^{-15} of the volume of the atom

Atomic Number and Mass Number (see MHR Text p. 43-46)

***Atoms of an element always have the same number of protons.

Atomic # = # of protons

*** In a neutral atom: **# of protons = # electrons**

*** The mass of an atom is made up of protons and neutrons. The mass number is equal to:

Mass # = # of protons + # of neutrons

Ex1 How many protons, electrons, and neutrons are in the following atoms?

Element	Atomic #	Mass #	#p	#e	#n
Carbon (C)		12			
Carbon (C)		13			
Magnesium (Mg)		24			
Magnesium (Mg)		25			
Magnesium (Mg)		26			

Ex2: Complete the following chart for the following atoms

Atomic Number	Mass Number	Number of protons	Number of electrons	number of neutrons	Element Symbol
8				9	
		11		10	
	40		20		
	88	38			

Isotopes: are atoms of the same element that have different numbers of neutrons, thus different mass numbers, hence different atomic masses.

Atomic mass - is the mass of an atom (grams).

Distinguish Between Isotopes of an Element

$^{A}_{z}X$ where X is the chemical symbol for the element
Z is the atomic #

A is the mass #

Ex: Hydrogen has 3 isotopes:

Hydrogen - 1 has 1 proton, 0 neutrons
Hydrogen - 2 has 1 proton, 1 neutron
Hydrogen - 3 has 1 proton, 2 neutrons

Symbols: $^{1}_{1}H$ $^{2}_{1}H$ $^{3}_{1}H$

Complete the following chart for the following isotope.

Isotope Symbol	$^{40}_{?=}K$	$^{?=}_{?=}N^{3-}$		
Name				**phosphoride**
Mass Number			**80**	**32**
Atomic Number			**35**	
Number of protons				
Number of electrons			**36**	
Number of neutrons		**16**		

Isotopes and Atomic Mass

-Rather than express the mass of an element as a very small number of grams attempts were made to compile tables of atomic masses based on a standard mass.

-For a very long time hydrogen was assigned a value of 1.00 and the masses of other elements were taken relative to that mass. Later the masses of elements were compared to a value of 16.000 for the oxygen atom.

-All isotopes to not occur in the same abundance in nature.

-In 1961, an international conference of chemists and physicists met to decide upon a single isotope as the basis for a new universal atomic mass table. The conference selected the most common isotope of carbon, carbon - 12, as the new standard.

-The **unified atomic mass unit** (symbol μ) is a new unit for measuring the small masses of the atoms. One unified atomic mass unit has a mass of 1.661×10^{-27} kg.

-One atom of carbon - 12 has a mass of 12μ.

-On the periodic table, the atomic mass of an element represents the average relative mass of all the naturally occurring isotopes of that element.

Determine Average Atomic Mass of an Element

Need to know the atomic masses and the abundance of its isotopes.

Avg. Atomic Mass = (atomic mass of Isotope A) (fraction of A) + (atomic mass of Isotope B) (fraction of B) +...

Ex 1: Calculate the average atomic mass of magnesium, given the relative atomic masses and abundances of its 3 naturally occurring isotopes:

Isotope	**Atomic mass (μ)**	**Abundance (%)**
Mg – 24	23.98	78.60
Mg – 25	24.99	10.11
Mg – 26	25.98	11.29

Ex 2: Calculate the average atomic mass of silver.

Isotope	**Atomic mass (μ)**	**Abundance (%)**
Ag - 107	106.9	51.8
Ag- 109	108.9	48.2

Ex 3: Calculate the average atomic mass of calcium.

Isotope	**Atomic mass (μ)**	**Abundance (%)**
Ca –20	39.00	98.6
Ca – 21	40.56	1.00
Ca - 22	39.99	1.40

The Mole (see MHR Text P. 47-48)

Reading a Balanced Chemical Equation

Q. Write a balanced chemical equation for the formation of hydrazine, N_2H_4, from its elements.

$N_2 + 2\,H_2 \rightarrow N_2H_4$

1 molecule of N_2 + 2 molecules of H_2 → 1 molecule of N_2H_4

6.022×10^{23} particles + $2(6.022 \times 10^{23}$ particles) → 6.022×10^{23} particles

1 mole + 2 mole → 1 mole

Q. Why do scientist use such a large number?

Since atoms, molecules and ions are extremely small particles, it is easier to use a large number when you are talking about their amounts (mass)

Ex1: 1 atom of Oxygen has a mass of 2.65×10^{-23}g (extremely small mass)

6.02×10^{23} atoms (1 mole) has a mass of 16.00 g (manageable mass)

Ex2: 1 molecule of sugar ($C_{12}H_{22}O_{11}$) has a mass of 5.69×10^{-22} g (extremely small mass)

6.02×10^{23} molecules (1 mole) has a mass of 342.34 g (manageable mass)

The Mole

-is 6.02×10^{23} particles (atoms, ions or molecules). Often referred to as Avagadro's number after the scientist who determined it.

- is the number of atoms in exactly 12 g of the most common isotope of carbon - carbon-12.

Types of particles:

atoms in elements
molecules in molecular compounds
formula units in ionic compounds

Formula of the Mole

$$\mathbf{n = \frac{N}{N_A}}$$

where n = # of moles
N = # particles
N_A = Avagadro's Number = 6.022×10^{23} particles/mol

Converting moles to number of particles or opposite

Ex1: How many atoms of Fe are in 0.50 mol of Fe?

Ex2: How many moles of CH_4 molecules are there in 9.0×10^{23} molecules of CH_4?

Ex3: Consider a 0.575 mol sample of potassium carbonate, K_2CO_3.

a) How many formula units are in the sample?

b) How many potassium ions are in the sample?

c) How many carbonate ions are in the sample?

d) How many oxygen atoms are in the sample?

e) How many atoms are in the sample?

Ex 4. A sample contains 1.25 moles of nitrogen dioxide, N_2O?

a) How many molecules are in the sample?

b) How many atoms are in the sample?

Ex 5: How many moles of are present in a sample of carbon dioxide, CO_2, made up of 5.83×10^{24} molecules ?

Molar Mass (see MHR Text p. 55-57)

How do chemists determine the number of moles in a sample?
It is impossible to count the atoms, molecules, or ions in a sample.

Chemists define Avagadro's number as the number of carbon atoms in exactly 12 grams of carbon-12. Accurately, there are 6.022045×10^{23} atoms of C-12 in 12 g.
or
1 mole of C-12 has a mass of 12 g.

Chemists then determine the mass of 1 mole of each element relative to carbon 12. These values are placed on the periodic table and are called the <u>average atomic mass</u> or

molar mass

- is the mass of one mole of particles of a substance.
- units are grams per mole, **g/mol**, $g \cdot mol^{-1}$

NOTE: 1) **Molar masses of elements are found on the periodic table.**

Ex: Molar mass of Na is 22.99 g/mol
Molar mass of Au is 196.97 g/mol

2) Molar mass of a compound is found by determining

- the number of each element contained in the formula
- multiply that number by the element's molar mass
- add all molar masses together

Ex: a) Find molar mass of Na_2CO_3

b) Find molar mass of $Mg(NO_3)_2$

c) Find molar mass of $Al_2(Cr_2O_7)_3$

d) Find the molar mass of $CrCl_3 \cdot 6\ H_2O$

Moles to Mass Calculations (see MHR Text p. 58-59)

1.0 mol of oxygen atoms = 16.00 g (molar mass)

0.50 mol of oxygen atoms = __?__ g

2.0 mol of oxygen atoms = __?

Formula: n = m where m = mass

M n = # of moles
M = molar mass

How do you determine mass given number of moles? **Rearrange the mole equation**

mass = moles x molar mass
= 0.50 mol x 16.00 g/mol
= 8.0 g

mass = moles x molar mass
= 2.0 mol x 16.00 g/mol
= 32 g

Ex1. What is the mass of a 2.40 mol of aluminium nitrate?

Ex 2. What is the mass of 0.58 mol of iron (III) oxide?

Mass to Moles Calculations (see MHR Text p. 59-60)

Ex 1. Find the number of moles in 14.1 g of xenon hexafluoride.

Ex 2. Calculate the number of moles of 0.59 kg of lead (II) phosphide. Express answer in scientific notation.

Particles to Mass (see MHR Text p. 62-63)

Ex 1: What mass does 3.01×10^{23} atoms of potassium have?

Ex 2: What mass does 8.96×10^{25} molecules of dinitrogen tetraoxide have?

Mass to Particles (see MHR Text p. 63-64)

Ex 1: How many atoms (particles) of calcium are in 5.00 g?

Ex 2: How many formula units does 10.0 g of magnesium hydroxide contain?

Percent Composition (see MHR Text p. 83-85)

Percent Composition is the percentage of the total mass; each type of atom present in compound contributes.

Ex1: Determine the percentage composition of sucrose, which has the formula $C_{12}H_{22}O_{11}$.

Ex2: Determine the percentage composition of $C_3H_8O_2$.

Ex 3: Determine the percentage composition of nitrogen in N_2O_5.

Determining Empirical Formulas (see MHR Text p. 87-91)

Determining the chemical formula from experimentally determined % composition data is more important.

Empirical Formula is the lowest whole number ratio of atoms (or ions) in a compound.
Molecular Formula is the actual number of atoms of each element in one molecule of the compound.
Ex 1: A compound consists of 40.1 % sulfur and 59.9% oxygen. What is its empirical formula?

Ex 2: A compound consists of 81.7 % carbon and 18.3 % hydrogen. What is its empirical formula?

Ex 3: A compound consists of 43.7 % phosphorus and 56.4 % oxygen. What is its empirical formula?

Determining Molecular Formula (see MHR Text p. 95-98)

Ex1: A compound contains 40.0 % carbon, 6.7% hydrogen and 53.3% oxygen by mass. The molar mass of the compound is 90.1 g/mol. Determine the molecular formula of the compound.

Ex 2: A compound contains 75.0 % carbon, 5.05% hydrogen and 20.0 % oxygen by mass. The molar mass of the compound is 240.28 g/mol. Determine the molecular formula of the compound.

Experimental Determination of Empirical Formula (MHR Text p. 101 - 103)

Ex1: After a 4.626 g sample of silver oxide is heated, 4.306 g of silver metal remains. What is the empirical formula of the compound?

Ex2: 25.8 g of hydrated sodium carbonate was heated to remove all the water. If 9.55 g of the anhydrous salt remains determine the formula of the hydrate.

Ex3: When 18.8 g of a chloride of antimony is heated, 7.66 g of antimony metal is recovered. What is the empirical formula of the compound?

Reading Balanced Chemical Equations (see MHR Text p. 114 - 118)

$$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(l)}$$

Read: 2 molecules of hydrogen gas reacts with 1 molecule of oxygen gas to produce 2 molecules of water

or 2 moles of hydrogen gas reacts with 1 mole of oxygen gas to produce 2 moles of water

NOT 2 grams of hydrogen gas reacts with 1 gram of oxygen gas to produce 2 grams of water

***** cannot compare masses, only moles**

STOICHIOMETRY is the prediction of how much of one substance will react with or be produced in a chemical reaction relative to the amount of another substance in the reaction.

NOTE:

1) The coefficients of a B.C.E. tell you the ratio (in moles) of reactants to products in a chemical reaction.

$$2\ H_{2(g)}\ +\ O_{2(g)}\ \rightarrow\ 2\ H_2O_{(l)}$$

ie: 2 mole of H_2 reacts with 1 mole of O_2
2 mole of H_2 produces 2 mole of H_2O
1 mole of O_2 produces 2 mole of H_2O

2) **Mole Ratio** is the ratio of one substance to another substance in a B.C.E.

$$\frac{2\text{ mole }H_2}{1\text{ mol }O_2} \qquad \frac{2\text{ mole }H_2}{2\text{ mole }H_2O} \qquad \frac{1\text{ mole }O_2}{2\text{ mole }H_2O}$$

Gravimetric Stoichiometry is the prediction of the number of moles or the mass of one substance given the moles or mass of another in the reaction.

Mole → Mole

Ex #1: In 1781, Henry Cavendish proved that water was the only product from the combustion of hydrogen. How many moles of hydrogen are required to react with exactly 4.13 mol of oxygen?

Ex #2: Ammonia is produced from the reaction of nitrogen and hydrogen. How many moles of ammonia are produced from 8.50 mol of hydrogen?

Ex #3: What moles of water are produced when 4.60 mol of oxygen is reacted with sufficient hydrogen?

Mole → Mass

Ex #4: What mass of water is produced when 8.60 mol of oxygen is reacted with sufficient hydrogen?

Ex #5: 5.00 mol of hydrogen will produce what mass of ammonia?

Mass→ Mole

Ex #6: How many moles of water are produced, from the reaction of 8.00 g of hydrogen?

Ex 7: 3.00g of nitrogen will produce how many moles of ammonia?

Ex #8: 7.50 g of H_2 will require how many moles of nitrogen to react completely?

Mass to Mass Stoichiometery

Ex 1: What mass of carbon dioxide is produced when 99.5 g of propane (C_3H_8) is burned?

Ex 2: Iron reacts slowly with hydrochloric acid to produce Iron(II) chloride and hydrogen gas. What mass of HCl is required to react with 3.56 g of Iron.

Ex 3: Copper metal reacts with silver nitrate to produce a precipitate of silver. Calculate the mass of silver produced, if 5.00 g of copper reacted with excess silver nitrate.

Theoretical Yield is the amount of product (usually in grams) that is formed based on the stoichiometry of the chemical reaction.

Some problems will say find the theoretical yield of the product instead of find the mass.

Theoretically, if 99.5 g of propane is burned with sufficient oxygen, 298 g of carbon dioxide should be produced. However, that may not be the case when the experiment is carried out? Some mass of carbon dioxide could have been lost due to experimental errors.

For most chemical reactions you do not get the theoretical yield. There are a number of reasons:

1) the rxn is slow and may not have been left long enough for complete rxn.
2) impurities present
3) rxn sets up an equilibrium (more next year)

Actual Yield is the amount of product that is actually produced in a chemical reaction when performed in the lab.

Percent Yield is the ratio of the actual yield to the theoretical yield expressed as a percent.

$$\% \text{ yield} = \frac{\text{actual yield (A.Y.)}}{\text{theoretical yield (T.Y.)}} \times 100\%$$

Ex1: Copper metal and silver nitrate react to produce silver metal. What is the percent yield of this reaction if 4.00 g of copper reacts completely and 15.20 g of silver is collected in the lab?

Ex2: Zinc and lead (II) nitrate react to produce lead metal. What is the percent yield of this reaction if 8.00 g of zinc reacts completely and 26.6 g of lead is collected in the lab?

LIMITING REAGENTS (see MHR Text p. 128 - 131)

Analogy: 1 ham + 2 bread $\rightarrow$ 1 sandwich

How many sandwiches can be made from 6 slices of ham and 10 slices of bread?

limiting reagent (reactant) is the reactant that is completely consumed in a chemical reaction. (It determines the amount of product produced.) **excess reagent (reactant)** is the reactant that is present in **more** than the required amount for complete reaction.

Ex1: What mass of magnesium oxide is produced from the reaction of 15.0 g of magnesium and 12.0 g of oxygen.

$2\ Mg_{(s)}$	$+\ O_{2(g)}$	$\rightarrow$	$2\ MgO_{(s)}$
m= 15.0 g	m= 12.0 g		m = ?

Which reactant do you use to find mass of product produced? The one that is used up first -the limiting reagent. It will produce the least amount of product

b) Which reactant is in excess? What mass of the excess reactant remains after the reaction?

Ex2: a) What is the theoretical mass of white precipitate produced when 7.98 g of sodium phosphate reacts with 9.55 g of calcium nitrate?

b) Based on the data was collected in the lab, what is the percent yield of the precipitate?
mass of filter paper + ppt = 5.99 g
mass of filter paper = 0.92 g

c) What mass of excess reagent remains at the end of the experiment?

Ex 3: a) A mass of 4.00 g of lead (II) nitrate was reacted with 3.50 g potassium iodide. Determine the theoretical mass of precipitate that forms. (Determine the limiting reagent).

b) The reaction flask was allowed to stand overnight and the precipitate separated from the solution by filtration, washed several times and allowed to dry overnight.

mass of filter paper	= 0.85 g
mass of filter paper + ppt	= 5.18 g

What is the actual mass of lead (II) iodide produced?

c) Determine the % yield of the precipitate.

d) From the mass of potassium iodide used, calculate the mass of lead (II) nitrate that reacted.

e) What mass of lead (II) nitrate was in excess (ie: did not react)?

Matter as Solutions (see MHR Text p. 237-240)
Solution is a homogeneous mixture; a mixture of uniform composition prepared by dissolving a solute in a solvent.
Solute: a substance, either solid, liquid, or gas, that dissolves in a solvent to produce a solution.
Solvent: a substance, either solid, liquid or gas, in which a solute dissolves to produce a solution; commonly a liquid.

Examples of Different Solution Types
1) **Solids in Liquids:** Ex: salt in water solute: salt solvent: water

Aqueous Solutions (aq) are those solutions where water is the solvent.
Soluble: term used to describe a solid that dissolves in a liquid.
Insoluble: term used to describe a solid which does not dissolve in a given liquid.

2)**Liquids in Liquids:** Ex 1: ethanol in water (alcohol) Ex 2: acetic acid in water (vinegar)

Miscible: is a term referring to two liquids that will dissolve in one another. (Liquid of smaller amount is referred to as the solute).

Immiscible: is a term referring to two liquids that will not dissolve when mixed. Ex: oil in water

3) **Gases in Liquids**
Ex: Carbonated drinks contain dissolved CO_2. Solute: CO_2 Solvent: flavoured liquid

4)**Liquids in Gases:** Ex: water vapour in air (Moist air)

5)**Gases in Gases:** Ex: air around us Solute: oxygen, etc... Solvent: nitrogen

6) **Liquids in Solids**
Ex: tooth fillings liquid mercury with solid silver, tin or copper
amalgam is a solution of a liquid in a solid

7)**Solids in Solids**

Ex: Brass	85% Cu	15% Zn	
Stainless Steel	74% Fe	18% Cr	8% Ni
Sterling Silver	92.5% Ag	7.5% Cu	

Alloys: a solution of two or more metals that have been melted together then cooled back to the solid state.

Three Types of Solutions
1) **Saturated Solution:** a solution that contains the maximum amount of solute that can be dissolved in a given amount of solvent at a given temperature.
2) **Unsaturated Solution**: a solution capable of dissolving more solute in a given amount of solvent at a given temperature.
3) **Supersaturated Solution**: a solution containing more solute than can normally be dissolved in a given amount of solvent at a given temperature.

Q. Given a solution of sodium sulfate decahydrate, how can you tell if it is saturated, unsaturated or supersaturated?
Add a crystal of the salt. If the solution is unsaturated, the crystal will dissolve. If the solution is saturated, the crystal will not dissolve. If it is supersaturated, the solution will crystallize.

<u>Two Classes of Solutes</u>

1) Non-electrolytes

-are those compounds that **do not conduct** an electric current in aqueous solution or molten state.

-generally include **molecular compounds** because they are nonionic. When they dissolve, they separate into individual neutral molecules.

Ex: $C_{12}H_{22}O_{11(s)} \rightarrow C_{12}H_{22}O_{11(aq)}$

2) <u>Electrolytes</u>

-are those compounds that **conduct** an electric current in aqueous solution or molten state.

-generally include **ionic compounds** and **acids,** which form ions when dissolved in water.

Strong Electrolytes: are those in which a large portion of the solute exists as ions.

1) ionic compounds with high solubility

$$NaCl_{(s)} \rightarrow \underset{100\%}{Na^+_{(aq)}} + Cl^-_{(aq)}$$

2) a strong acid - there are 6 which form 100% ions in solution.

$HClO_4$, H_2SO_4, HNO_3, HCl, HBr, HI

Weak Electrolytes: are those where only a small portion of the solute exists as ions.

1) ionic compounds of low solubility (chart)

$$\underset{<50\%}{BaSO_{4(s)}} \rightarrow Ba^{2+}_{(aq)} + SO^{2-}_{(aq)}$$

2) a weak acid - any acid that is not one of the 6 strong.

$$\underset{1.3\%}{CH_3COOH_{(l)}} \rightarrow H^+_{(aq)} + CH_3COO^-_{(aq)}$$

<u>Concentrated vs Dilute Solutions</u>

Concentrated Solution: is one which has a large amount of solute dissolved in a given amount of solvent.

Dilute Solution: is one which has a small amount of solute dissolved in a given amount of solvent.

Qualitatively

These are relative terms used when comparing the concentrations of two similar solutions.

Ex: Sol^n A 5 g of NaCl in 100 mL of H_2O

Sol^n B 10 g of NaCl in 100 mL of H_2O

** Sol^n B is more <u>concentrated</u> than Sol^{n}A.

** Sol^n A is more <u>dilute</u> than Sol^n B.

Quantitatively

Concentrated Sol^n is a term used often to describe acid & base solutions of known concentration.

Ex:	Concentrated HCl	12 mol/L
	Concentrated H_2SO_4	18 mol·L^{-1}
	Concentrated HNO_3	16 mol/L

**** Less than 3 mol/L is considered a dilute acid solution**

Expressing Concentration (see MHR Text p. 255-265)

There are many ways of expressing concentration.

Percent by mass (m/v)

– often used for a solid dissolved in liquid sol^n.

– is the number of grams of solute per 100 mL of sol^n.

– $$\% (m/v) = \frac{\text{mass of solute (g)}}{\text{Total solution volume (mL)}} \times 100\%$$

Ex1: A solution contains 8.0 g of sodium chloride in 50. mL of solution. What is the percent by mass of solution?

Ex 2: How many grams of sodium chloride would you need to prepare 2.0 L of 2.0 % (m/v) sodium chloride solution?

Percent by volume (v/v)

- often used for a liquid dissolved in liquid sol^n.
- is the volume of solute in mL per 100 mL of sol^n.
- $$\% \text{ (v/v)} = \frac{\text{volume of solute (mL)}}{\text{solution volume (mL)}} \times 100\%$$

Ex 3: If 40.0 mL of rubbing alcohol is diluted to 200.0 mL of solution. What is the % by volume?

Ex 4 : Rubbing alcohol is sold as a 70.0% (v/v) solution of isopropyl alcohol in water. What volume of isopropyl alcohol is used to make 500.0 mL of rubbing alcohol?

Parts per Million (ppm) & Parts per Billion (ppb)

- are used to express very small quantities of substances
- is used to describe the ***mass*** of the solute compared to the ***mass*** of the solvent
- $$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$$

$$\text{ppb} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^9$$

Ex: Water containing more than 50 ppb lead is unfit to drink. A certain sample of water is found to contain 8.5×10^{-5} g of lead in 1.0×10^{3} g of water. Is it safe to drink?

Molar Concentration (see MHR Text p. 266-268)

Molar Concentration (or molarity) in SI it is defined as the number of moles of solute dissolved in a litre of solution.

Formula: concentration = $\dfrac{\text{amount of solute in moles}}{\text{volume of sol}^{n}\text{ in litres}}$

$C = \dfrac{n}{v}$ Units: mol/L or $mol \cdot L^{-1}$ or M

Problems:

1. Calculate the concentration of a solution made by dissolving 0.900 mol of sodium chloride in 500.0 mL of water.

2. Calculate the concentration of an antacid solution made by dissolving 15.0g of sodium bicarbonate (baking soda) in enough water to make 250.0 mL of solution.

3a) Potassium Permanganate is a powerful oxidizing agent. How many moles of potassium permanganate are in a 2.00 L solution that has a concentration of 0.0025 mol/L?

b) What mass of potassium permanganate was dissolved?

4. What mass of washing soda, sodium carbonate decahydrate, is necessary to make 400.0 mL of a 0.0500 mol/L solution?

5. Vinegar is a dilute solution of acetic acid. What volume of 0.800 mol/L vinegar solution contains 1.60 mol of acetic acid?

6. Sodium hydroxide, commonly known as caustic soda, has many uses in the labratory and in industry. What volume of 0.600 mol/L NaOH can be prepared from 4.8 g of solute?

PREPARATION OF A STANDARD SOLUTION (from a solid reagent)

Standard Solution: is a solution of precisely known concentration made using precision equipment to measure mass of solute and volume of solution.

Precision mass - electronic balance (analytical)
Precision volume - volumetric flask

Steps for the Preparation of A Standard Solution

1. Calculate the mass of solute necessary to make the volume and concentration of the solution needed.
2. Obtain the necessary mass of solute in a clean, dry beaker, and weigh on the analytical balance.
3. Transfer the solute to the volumetric flask using a funnel.
4. Using an analytical balance and weighing by difference, calculate the exact mass of solute transferred to the volumetric flask.
5. Dissolve the solid in distilled water, using less than one-half of the final solution volume.
6. Add distilled water, using a dropping pipet for the final few milliliters while using a meniscus finder to set the bottom of the meniscus on the calibration line.
7. Stopper the flask and mix the solution slowly by inverting the flask several times.

DILUTION (MHR Text p. 272-273 & 276)

Dilute Solution: is one that has a small amount of solute dissolved in a given amount of solvent.

Dilution: is the process of decreasing the concentration of a solution by adding more solvent.

Note: 1) $c_i > c_f$ where i initial (before dilution)
2) $v_i < v_f$ f final (after dilution)

Dilution Calculation: When diluting a solution the amount of solute does not change.

ie: $n_i = n_f$ but $n = c\,v$

thus $c_i v_i = c_f v_f$ ** used for dilution problems **only**.

Problem #1: A student requires 100.0 mL of a 6.0 M solution of HCl to do an experiment. What volume of concentrated HCl (12 mol/L) is required to make this solution?

Problem #2: Concentrated acetic acid is 99.5% pure and has a concentration of 17.4 mol/L. What is the concentration of vinegar solution if 200.0 mL of concentrated acetic acid is diluted to fill a 4.00 L bottle?

Hints for Solving Dilution Problems

1) Given two volumes in the problem, the smaller volume will be the initial one, the larger volume the final. $v_i < v_f$
2) Given two concentrations in the problem, the larger concentration will be the initial one, the smaller concentration the final. $c_i > c_f$
3) A CONCENTRATED solution or volume of concentrated indicates values for c_i and v_i.
4) Concentration and Volume values following phrases such as:
   1) diluted to,
   2) prepared a solution of,
   3) required to make, indicate values for c_f and v_f.

<u>Steps for the Preparation of A diluted</u> .

These are the steps involved in diluting small quantities of concentrated solutions.
1. Use a 10.0 mL graduated pipette to deliver a specific volume of concentrated solution into a clean 500.0 mL volumetric flask.
2. Add distilled water using a wash bottle to bring the volume of in the flask up to the fill line on the neck of the flask.
3. Stopper, invert the volumetric flask/container several times, and label.

Concentration of Ions in Solution (see MHR Text p. 299-300)

Concentration of Ions

Most ionic compounds actually exist as individual ions in solution. Often it is the concentration of an individual ion in solution that we are concerned with and not the concentration of the compound.

Ex 1: What is the concentration of the lead ion and the nitrate ion in a solution of 3.0 mol/L lead (II) nitrate?

Step 1: Write a dissociation equation. **** Dissociation equation must be balanced.**

Step 2: Use coefficients from the equation to determine conc^n of each ion.

If lead (II) nitrate were dumped into our water supply.

$$\underset{}{Pb(NO_3)_2\,(s)} \rightarrow \underset{\text{harmful}}{Pb^{2+}(aq)} + \underset{\text{not harmful}}{2\ NO_3^-\,(aq)}$$

Ex 2: What is the concentration of each ion in a solution of 0.34 mol/L iron (III) sulfide?

Ex 3: What is the concentration of each ion in solution if 5.30 g of magnesium hydroxide was dissolved to form 500.0 mL of solution?

Ex 4: Given a solution of aluminium sulfide where the concentration of aluminium ion is 0.025 mol/L

i) what is the conc^n of the solute?
ii) what is the conc^n of the sulfide ion?

**Solution Stoichiometry (**MHR Text p. 303 - 307)

Solution Stoichiometry: is a method of predicting the concentration, the volume, or the mass of a substance in a chemical reaction, given the concentration and volume of another substance.

Ex1: A student uses 32.5 mL of 0.150 M NaOH solution to neutralize 50.0 mL of HNO_3. What is the molar concentration of HNO_3?

Ex 2: Calculate the volume of a 0.0250 M hydrochloric acid solution that can be neutralized by the reaction with 10.0 mL of a 0.0500 M calcium hydroxide solution.

Ex 3: What is the theoretical yield (in grams) of a precipitate produced when 80.0 mL of a 2.00 mol/L solution of silver nitrate reacts completely with excess calcium chloride solution?

Ex 4: 25.0 mL of 1.0 mol/L barium hydroxide reacts with 15.0 mL of 0.075 mol/L sulfuric acid. What mass of precipitate is produced?

<u>Gaseous State of Matter</u> (see MHR Text p. 66-70)

Gas

- is a substance that fills and assumes the shape of its container, diffuses rapidly, and mixes readily with other gases. (**Diffuse**-move spontaneously through any available space.)
- is highly compressible - decrease in volume when pressure is applied.
- is affected by temperature change - as temperature increases the volume and/or the pressure increases.

Amadeo Avagadro (Italian): proposed that equal volumes of gases at the same temperature and pressure contain equal number of molecules.

Avagadro's Hypothesis

At a constant temperature and pressure, the volume occupied by a gas depends directly on the number of gas particles (or number of moles).

Molar Volume

- is the volume that one mole of a gas occupies at specified temperature and pressure. It has been determined to be 22.4 L/mol at STP
- is useful since it is more convenient to measure the volume of a trapped gas than to measure its mass.

Standard Temperature and Pressure (STP)

Since gases measurements are temperature and pressure dependent, there are standard conditions set.

Standard Temperature is $0.00^{o}C$
Standard Pressure is 101.3 kPa

Q. How to calculate volumes of gases not at STP?
Answer. Use Ideal Gas Law (PV = nRT) - not in this course

Calculating Moles and Volume of a Gas

At STP 1 mole of gas = 22.4 L
2 moles = _____ ?
0.5 moles = _____ ?

Thus: $v_{gas} = n \times MV$ where v - volume of gas
n - moles of gas
V_m - molar volume

At STP 22.4 L = 1 mole
11.2 L = _____ ?
44.8 L = ______ ?

Thus: $n_{gas} = \frac{v}{V_m}$

Ex 1: What amount (in moles) of oxygen is available for a combustion reaction if its volume is 5.6 L at STP?

Ex 2: What volume is occupied by 0.024 mol of CO_2 at STP?

Ex 3: What mass of oxygen is available for a combustion reaction if its volume is 11.2 L at STP?

Ex 4: What volume is occupied by 88.0 g of carbon dioxide gas at STP?

Gas Stoichiometry (see MHR Text p. 119-123)

Gas Stoichiometry: is a method of using mole ratios to predict the amount (volume, moles, mass) of a gas used in a chemical reaction.

Ex 1:Water can be decomposed to produce hydrogen and oxygen gas. What volume of oxygen is produced if 10.0 L of hydrogen was produced at STP?

Ex 2: If 8.9 kg of propane gas (C_3H_8) is burned what is the total volume of gas that is produced at STP?

Ex 3: When 55.0 mL of 0.524 M solution of potassium carbonate reacts with 27.4 mL of a 0.724 M solution of phosphoric acid, what volume of carbon dioxide gas is produced at STP.
$3\ K_2CO_{3\ (aq)} + 2\ H_3PO_{4\ (aq)} \rightarrow K_3PO_{4\ (aq)} + 3CO_{2\ (g)} + 3H_2O_{\ (g)}$

b) If the percent yield of carbon dioxide is 96.0% , then what is the actual yield?