Revision Seminar

Chemistry 11

Tutorials and Worksheets

List of Tutorials and Worksheets

Tutorial 1- Bonding

Bonding is how substances are held together.

Information about the bonding can be obtained from –

- **Melting point** – the higher the melting point, the stronger the bonding is.
- **Electrical conductivity** – conductors have charged particles that are able to move.

Keep in mind –

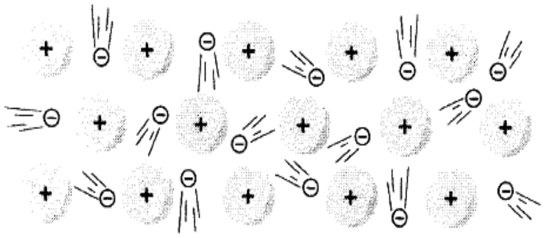
Bonding is the electrostatic attraction between something positive and something negative.

There are four types of substances with distinctive physical properties, which give us clues as to the structure and bonding present.

Type of Substance	Melting Point	Electrical Conductivity
Metallic		
Ionic		
Covalent molecular		
Covalent network		

Metals

Metal atoms lose their valence electrons to become positive ions. The valence electrons are delocalised and mobile.

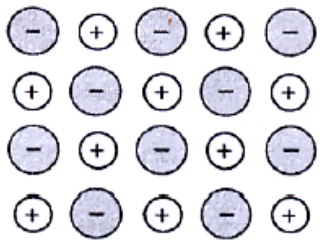


*The **metal bond** is the electrostatic attraction between the positive metal ions and the valence electrons.*

Property	Explanation
Conduct electricity	
Conduct heat	
Ductile and malleable	
High melting point	
Shiny	

Ionic

Metal and non-metal combine.
Metal loses electrons to become positive ions.
Non-metal gains electrons to become negative ions.



*The **ionic bond** is the electrostatic attraction between positive metal ions and negative non-metal ions.*

Property	Explanation
Non-conductor when solid	
Conductor when liquid	
Brittle	
High melting point	

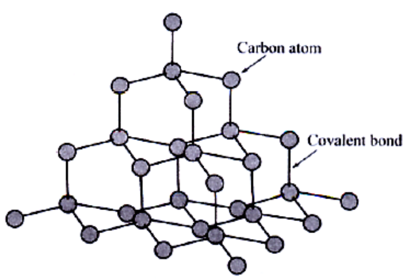
<u>Covalent</u>

A covalent bond forms when non-metal atoms come together. Electrons are shared so that each atom has a stable electron configuration.

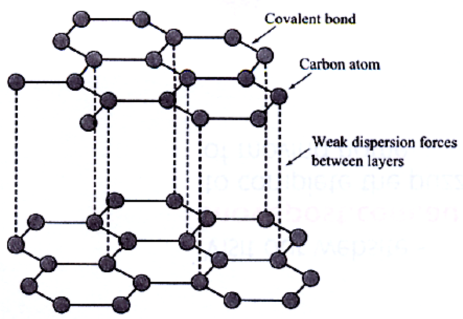
*A **covalent bond** is the electrostatic attraction between the negative shared electrons and the positive nuclei of the two adjacent atoms.*

If covalent bonds extend throughout the substance, it is called a **<u>covalent network</u>**.

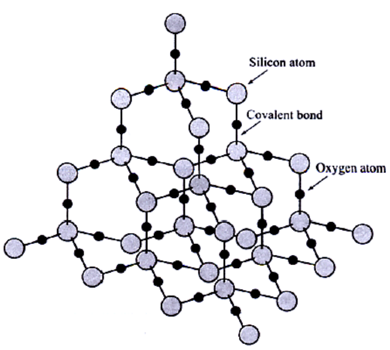
Diamond



Graphite

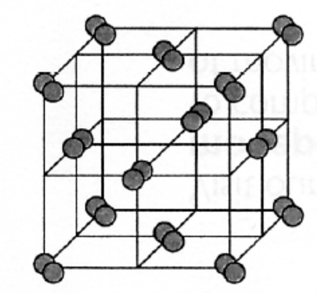


Silicon dioxide (sand)



The very strong covalent bonds explain the properties of covalent networks, which include very high melting point, very hard, unreactive and insoluble in water.

When the covalent bonding is restricted to just a few atoms, a **covalent molecule** forms. The bonds in a molecule are very strong, but molecules attract to each other only weakly.



Property	Explanation
Very low melting point	
Soft/waxy	
Non-conductor of electricity	

To decide which group a substance belongs to –

Metals – all have metallic bonding.

Non-metals – all have covalent bonding, with only carbon, silicon and a few of their compounds forming covalent networks. The rest are found as small molecules.

Metal and non-metal combined – all have ionic bonding.

Tutorial 2 – Formulas

Chemistry 11

ELEMENTS

Metals

Found as a continuous network – so cannot identify groups of atoms.

*The formula for a **metal** is its symbol only*

Non-metals

Carbon and silicon are found as continuous networks. Most other elements are found as small molecules, so the number of atoms must be shown. These include the diatomic (2 atom) molecules – H_2, O_2, $C\ell_2$, N_2, F_2, Br_2, and the other two main molecules of P_4 and S_8. The inert gases are found as single atoms.

The formula for a non-metal depends on its structure –

***Continuous networks** – symbol only.*

***Small molecules** – the formula must indicate the number of atoms present.*

***Inert gases** – symbol only.*

NON-METAL COMPOUNDS

Non-metal compounds can also be found as continuous networks or small molecules.

*The formula of a **continuous network** represents the simplest ratio of atoms present.*

*The formula of a **molecule** represents the actual number of atoms present.*

The rules for naming are the same for both types of compound.

The first element is given its element name.

The second element is modified to end in **"-ide"** (eg oxide, sulfide, and chloride)

The number of each element is indicated by a prefix, although it is rare for mono (1) to be used in naming most compounds.

Mono –1	Hex - 6
Di – 2	Hept - 7
Tri – 3	Oct - 8
Tetra – 4	Non – 9
Pent – 5	Dec – 10

For example –

Carbon dioxide – First element is C

Second element is O (from oxide)

Two O are present (di)

=> CO_2

PF_3 - First element is phosphorus

Second element is fluorine => fluoride

Three F are present => tri

=> Phosphorus trifluoride

IONIC COMPOUNDS

When a metal and a non-metal combine to form an ionic compound, a continuous network results – so no individual molecules can be identified.

*The formula of an **ionic compound** is the simplest ratio of ions present.*

Each ion has a stated charge or valence. This can be determined for simple (single atom) ions, directly from the periodic table, using electron configurations.

Eg Potassium has one valence electron in its outer shell => loses that electron to become more stable => +1 charge.

Most of the transition metals are +2, except for silver which is +1 and chromium which is +3. Several transition metals take more than one charge, but this is indicated in the name –

Eg Fe^{2+} is iron II.

Polyatomic ions are molecules that have a charge. Except for ammonium (NH_4^+) and mercury I (Hg_2^{2+}) these ions are negative.

To write a formula, simply look at the charge on each ion. The total amount of negative must equal the total amount of positive charge present.

Eg magnesium chloride => magnesium is Mg^{2+} and chloride is $C\ell^-$

To have equal amounts of positive and negative charge, two chlorides will be needed.

=> $MgC\ell_2$

Zinc hydroxide => zinc is Zn^{2+}, hydroxide is OH^-

Two hydroxides will be needed to balance the charge. Brackets must be placed around the OH^- group to indicate that two of the entire group is present.

=> $Zn(OH)_2$

In summary –
Must have both positive and negative ions.
There must be enough of each ion so that the total charge is zero.
Brackets are only used if there is more than one polyatomic ion present.
Roman numerals indicate the charge on ions that can have more than one charge.

Tutorial 3- Representing Bonding

11 Chemistry

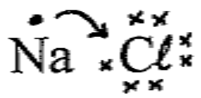
The group number gives the number of valence electrons in an atom.

An electron dot diagram shows the valence electrons for a single atom, a molecule or a compound.

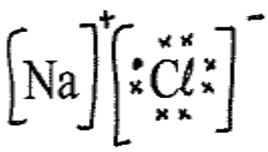
Simple Ionic Compounds

Metal atoms lose electrons to gain stable structure. Non-metal atoms gain electrons to gain stable structure.

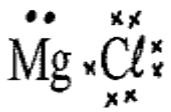
Eg $NaC\ell$



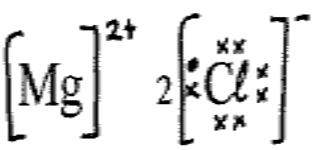
The extra electron from the sodium goes to the chlorine.



Eg $MgC\ell_2$



Magnesium has two electrons to lose, chlorine gains only one, and so two chlorine atoms are needed.

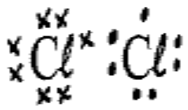


Molecules

Molecules are made from non-metals, all of which require extra valence electrons.

A **covalent** bond forms where the electrons are **shared**.

Eg $C\ell_2$



Eight electrons are obtained by sharing pairs of electrons. This is a **stable octet**.

The number of bonds likely to form can be identified from the number of electrons required –

Carbon	4 valence electrons	forms 4 bonds
Oxygen	6 valence electrons	forms 2 bonds
Hydrogen	1 valence electron	forms 1 bond
Nitrogen	5 valence electrons	forms 3 bonds

Valence Bond Diagrams

A line can replace each shared pair of electrons. Non-bonding electrons can be omitted.

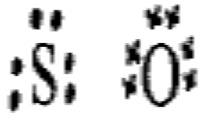
Eg $C\ell - C\ell$

$O = O$

$H - O - H$

Dative Bonds

Some formulas do not fit the expected number of atoms and bonds. Consider combining sulfur and oxygen.

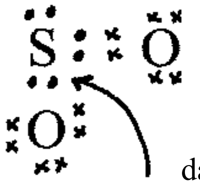


Might expect S = O.

But this is not found – we have SO_2 and SO_3 instead.

> In some cases, oxygen (in particular) will attach to a lone pair of electrons in another atom. This is a **dative** or **co-ordinate** bond, where two atoms share a pair of electrons, but both electrons have come from one atom only.

Eg SO_2



dative bond

This is represented as

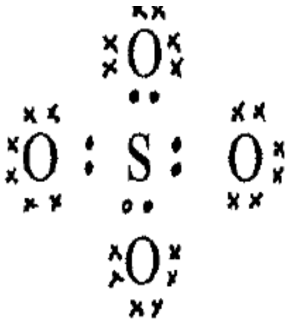
Charged Molecules

> Polyatomic ions are simply covalent molecules with a charge. This charge must be considered when drawing the electron dot diagram.

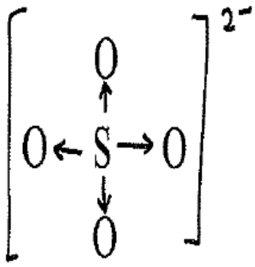
Eg SO_4^{2-}

Add in two electrons for the negative two charge

S has 8 electrons, so the O are placed in as dative bonds



Put in brackets and charge



These rules can be used for all electron dot diagrams –

1.	Decide if it is ionic – must use brackets and charge.
2.	If a molecule is present, select the central atom.
3.	Draw in the valence electrons of the central atom.
4.	Put in electrons due to charge if it is a molecular ion.
5.	Use single, double or triple bonds until the central atom has an octet.
6.	Remaining atoms are added as dative bonds
7.	Put in brackets and charge if it is an ion.
8.	Check that all atoms (except H) have 8 electrons.

Note – if an acid is given, it must be drawn as a molecule (not ions unless in water). Any hydrogen present will be mostly attached to the oxygen atoms.

EXAMPLES

Draw an electron dot diagram, and valence bond diagram where appropriate for the following –

1. carbon dioxide
2. calcium chloride
3. nitrogen gas
4. sodium sulfide
5. carbonate ion
6. phosphate ion
7. hydrogen chloride
8. sulfuric acid
9. oxygen gas
10. aluminium oxide
11. nitrate ion
12. hydroxide ion

Tutorial 4 – Writing and Balancing Equations

Chemistry 11

An equation represents what is happening in a chemical reaction.

Reacting substances are called **Reactants** (or sometimes reagents).

Produced substances are called **Products**.

Reactants → Products

The Law of Conservation of Mass states –

Matter cannot be created or destroyed in a chemical reaction

Hence the number of atoms in a chemical reaction must be constant for both reactants and products.

Each substance in a reaction needs to have its state identified –

(s) = solid	(l) = liquid	(g) = gas	(aq) = in water (aqueous)

If words are given instead of formulas, make sure you write the correct formulas for all reactants and products – it is terribly difficult to balance an equation with the wrong formulas!

For example –

Hydrogen gas and oxygen gas react to form water.

H_2 + O_2 → H_2O

2 x H atoms 2 x O atoms 2 x H atoms and 1 x O atom

Oxygen is not balanced, so an extra O is needed on the products side. So we need two H_2O.

H_2 + O_2 → $2H_2O$

2 x H atoms 2 x O atoms 4 x H atoms and 2 x O atoms

Now the hydrogen is not balanced, so two extra H's are needed on the reactants side. So we need 2 H_2.

$2H_2$ + O_2 → $2H_2O$

4 x H atoms 2 x O atoms 4 x H atoms and 2 x O atoms

All atoms are balanced, so now just put in state symbols –

$2H_2(g)$ + $O_2(g)$ → $2H_2O(l)$

Practice examples

1. $H_2 + O_2 \longrightarrow H_2O$
2. $H_2 + Cl_2 \longrightarrow HCl$
3. $CO_2 + C \longrightarrow CO$
4. $N_2 + O_2 \longrightarrow NO$
5. $N_2 + H_2 \longrightarrow NH_3$
6. $Na + O_2 \longrightarrow Na_2O$
7. $CO + O_2 \longrightarrow CO_2$

8. $Fe + Cl_2 \longrightarrow FeCl_3$

9. $Cu + O_2 \longrightarrow CuO$

10. $Zn + O_2 \longrightarrow ZnO$

11. $SO_2 + O_2 \longrightarrow SO_3$

12. $Mg + N_2 \longrightarrow Mg_3N_2$

13. $P + O_2 \longrightarrow P_2O_3$

14. $C + Fe_2O_3 \longrightarrow Fe + CO$

15. $Fe_2O_3 + CO \longrightarrow Fe + CO_2$

16. $P + O_2 \longrightarrow P_2O_5$

17. $C_6H_6 + O_2 \longrightarrow CO_2 + H_2O$

18. $NO + O_2 + H_2O \longrightarrow HNO_3$

19. $PbS + O_2 \longrightarrow PbO + SO_2$

20. $C_2H_6O + O_2 \longrightarrow CO_2 + H_2O$

21. $H_2S + O_2 \longrightarrow H_2O + S$

22. $Fe + O_2 \longrightarrow Fe_3O_4$

23. $Na_2O_2 + H_2O \longrightarrow NaOH + O_2$

24. $MnO_2 + HCl \longrightarrow MnCl_2 + H_2O + Cl_2$

25. $FeS + O_2 \longrightarrow Fe_2O_3 + SO_2$

26. $C_8H_{18} + O_2 \longrightarrow H_2O + CO_2$

27. $S + Fe_2O_3 \longrightarrow Fe + SO_2$

28. $ZnS + O_2 \longrightarrow ZnO + SO_2$

29. $CaCN_2 + H_2O \longrightarrow CaCO_3 + NH_3$

30. $NH_3 + O_2 \longrightarrow N_2 + H_2O$

31. $NO_2 + H_2O \longrightarrow HNO_3 + NO$

32. $CS_2 + O_2 \longrightarrow CO_2 + SO_2$

33. $Sb + Cl_2 \longrightarrow SbCl_3$

34. $C_{10}H_{16} + Cl_2 \longrightarrow C + HCl$

35. $CH_3OH + O_2 \longrightarrow H_2O + CO_2$

36. $CS_2 + Cl_2 \longrightarrow CCl_4 + S$

37. $C_2H_2 + O_2 \longrightarrow H_2O + CO_2$

38. $Pb(NO_3)_2 \longrightarrow PbO + NO_2 + O_2$

39. $Cu + HNO_3 \longrightarrow Cu(NO_3)_2 + NO_2 + H_2O$

40. $KMnO_4 + HCl \longrightarrow MnCl_2 + KCl + H_2O + Cl_2$

41. potassium hydrogen carbonate + nitric acid → potassium nitrate + water + carbon dioxide

42. aluminium carbonate → aluminium oxide + carbon dioxide

43. iron III carbonate + hydrochloric acid → iron III chloride + water + carbon dioxide

44. copper II sulfate + iron → iron II sulfate + copper

45. ammonium hydroxide + sulfuric acid → ammonium sulfate + water

46. chromium hydroxide + sulfuric acid → chromium sulfate + water

47. iron III oxide + carbon monoxide gas → iron + carbon dioxide

48. ammonium hydrogen carbonate → ammonium carbonate + water + carbon dioxide

49. potassium + water → potassium hydroxide + hydrogen gas

50. sodium hydroxide + sulfuric acid → sodium sulfate + water

51. magnesium + oxygen gas → magnesium oxide

52. sodium + water → sodium hydroxide + hydrogen gas

53. aluminium carbonate + hydrochloric acid → aluminium chloride + water + carbon dioxide

54. zinc oxide + nitric acid → zinc nitrate + water

55. ammonium carbonate + nitric acid → ammonium nitrate + carbon dioxide + water

Tutorial 5 Chemical Calculations

Chemistry 11

Relative Atomic Mass *(A_r) is the mass of one atom of an element compared to 1/12 of the mass of one atom of carbon-12.*

It is the average of all isotopes for a given element. for example –

Boron has approximately 20% B-10 and 80% B-11.

$$A_r(B) = \frac{20}{100} \times 10 + \frac{80}{100} \times 11$$

$$= 10.8$$ (no units are required as this is a relative value)

Relative Molecular Mass *(M_r) is the mass of one molecule of a substance compared to 1/12 of the mass of one atom of carbon-12.*

It is found by adding up all the A_r for the formula of the molecule, for example –

Ammonia is NH_3

$M_r(NH_3)$ = $A_r(N)$ + $3 \times A_r(H)$

= 14.01 + 3(1.008)

= 17.034 (no units remember)

When a substance is not found as discrete molecules, eg ionics or covalent networks, the **relative formula mass** is found. The symbol is still M_r and it is calculated in exactly the same manner, for example –

Sodium bromide is NaBr

$M_r(NaBr)$ = $A_r(Na)$ + $A_r(Br)$

= 22.99 + 79.90

= 102.89

*A **mole** is a number used to count atoms. It is also called **Avogadro's Number** (N_A) and is equal to 6.022×10^{23}*

Atoms are so small that a very large number is needed to count them. When we compare one substance with another we are always looking at how many atoms are involved, so the mole is a convenient number to use (like counting eggs in dozens).

1 mole of H_2 has 6.022×10^{23} molecules of H_2

2 mole of H_2 has $2 \times 6.022 \times 10^{23}$ molecules of H_2

1 mole of H_2 has $2 \times 6.022 \times 10^{23}$ atoms of H atoms (2 H atoms in each molecule)

Or – 1 mole of H_2 has 2 mole of H atoms (this is called a **mole ratio**)

Molar mass (M) is the mass of the relative atomic, molecular or formula mass of a substance in grams.

This is the mass of one mole, and is found by finding A_r or M_r and giving the unit of "g" . For example –

The molar mass for silver nitrate ($AgNO_3$) is

$$M(AgNO_3) = 107.9 + 14.01 + 3(16.00)$$
$$= 169.91 \text{ g}$$

Once the molar mass is known, it is easy to find out number of mole or mass using the following formula –

$$n = \frac{m}{M}$$

For example –

How many mole is 8.05 g of sodium hydroxide?

$$n = \frac{m}{M}$$

M(NaOH) = 39.998 g

$$n(NaOH) = \frac{8.05}{39.998}$$
$$= 0.201 \text{ mol (3 sf)}$$

What is the mass of 2.50 mole of potassium iodide?

$M(KI) = 166.0$ g

$m = M \times n$

$= 166.0 \times 2.50$

$= 415$ g (3 sf)

Percentage composition *is the percentage by mass of each element in a compound.*

This can be calculated by comparing the mass of each element with the total mass of the compound, for one mole. For example –

What is the percentage composition of iron III oxide?

Fe_2O_3

Fe	2 x 55.85	111.70	%Fe = (111.7/149.7) x 100 = 74.62%
O	3 x 16.00	48.00	%O = (48/149.7) x 100 = 25.38%

Total = 149.70 g

<u>Questions</u>

1. How many atoms of oxygen in 2 mole of -
   a. oxygen gas
   b. lead II sulfate
   c. potassium oxide
2. What is the mass of 4.50 mole of
   a. hydrogen gas
   b. sulfur trioxide
   c. sodium carbonate - 10 - water
3. How many mole is 100.0 g of -
   a. carbon dioxide
   b. magnesium sulfide
   c. copper II carbonate
4. What is the percentage composition of -
   a. sodium chloride
   b. nitrogen dioxide
   c. ammonium nitrate

Tutorial 6 – Stoichiometry

Chemistry 11

Stoichiometry is the relationship between atoms in a formula or substances in a chemical reaction.

<u>In a formula</u>

The formula represents the number of atoms present in a molecule (made of non-metals only) or the ratio of atoms in a network (mostly ionic substances – metal and non-metal combined).

H_2O is a molecule, NaCl is a network.

It doesn't matter if the formula is for a molecule or network – we can still identify the ratio of atoms involved.

In H_2O, the 2 after the hydrogen indicates that 2 atoms of hydrogen are present. If no number is given (as for the oxygen) then 1 atom is present.

So for H_2O, the stoichiometric ratio is –

2 hydrogen : 1 oxygen

In Na_2CO_3, there are 2 atoms of sodium, 1 atom of carbon, and 3 atoms of oxygen. And the stoichiometric ratio is –

2 sodium : 1 carbon : 3 oxygen

In equations

A correctly balanced equation represents the number of particles of each substance involved. The numbers in front of each substance indicate how many of each particle is reacted or produced.

For example –

$$2H_2 \quad + \quad O_2 \quad \rightarrow \quad 2H_2O$$

In this reaction, 2 molecules of hydrogen gas react with 1 molecule of oxygen gas, to give 2 molecules of water.

Stoichiometric ratio is –

2 hydrogen : 1 oxygen : 2 water

Using this ratio

Once the stoichiometric ratio is established for any substance or reaction, this can be used to determine quantities.

If one molecule of water has 2 hydrogen atoms, then 2 molecules would have 4 atoms of hydrogen. 10 molecules would have 20 hydrogen atoms, and so on.

This allows us to move up to a much larger number that is needed in chemistry, due to the very small size of atoms. This number is **the mole**, a number used to count atoms, and equal to 6.023×10^{23}.

So 6.023×10^{23} molecules of water would contain $2 \times 6.023 \times 10^{23}$ atoms of hydrogen.

Or more simply – 1 mole of water contains 2 mole of hydrogen atoms.

The stoichiometric ratio can hence be represented as a ratio of mole as well as ratios of atoms, molecules etc.

In the water reaction –

$$2H_2 \quad + \quad O_2 \quad \rightarrow \quad 2H_2O$$

2 mole of hydrogen molecules reacts with 1 mole of oxygen molecules to give 2 mole of water molecules.

Examples

1. What is the ratio between the various atoms in $CuSO_4$?
2. What is the mole ratio of magnesium atoms to oxygen molecules in –

   $2Mg + O_2 \rightarrow 2MgO$
3. How many mole of magnesium is needed to produce 20 mole of magnesium oxide in the equation above?
4. How many mole of magnesium oxide is produced from 6 mole of oxygen molecules?
5. How many atoms of oxygen are in 24 molecules of carbon dioxide (CO_2)?
6. How many mole of silver nitrate ($AgNO_3$) contains 4 mole of nitrogen atoms?
7. How many mole of hydrogen ions is needed to react with 3 mole of zinc metal in the equation below?

   $Zn + 2H^+ \rightarrow Zn^{2+} + H_2$
8. How many mole of hydrogen gas are produced from 5 mole of hydrogen ions in the equation above?
9. Methane (CH_4) burns in oxygen gas to form carbon dioxide and water. Write a balanced equation to represent this.
10. How many mole of carbon dioxide are produced from 1.8 mole of methane?
11. How many mole of oxygen are needed to produce 3 mole of water?
12. The formula for sodium hydrogen carbonate decahydrate is $NaHCO_3.10H_2O$. How many mole of water is present in each mole of sodium hydrogen carbonate decahydrate?
13. How many oxygen atoms are present in 10 formula units (not molecules because it is ionic) of sodium hydrogen carbonate decahydrate?
14. Nitrogen gas reacts with hydrogen gas to form ammonia (NH_3). Write a balanced equation.
15. How many mole of ammonia are produced for each mole of hydrogen gas used?

Tutorial 7 – Limiting Reagent

Chemistry 11

In chemical reactions we can determine the amount of each reactant required for a particular result. However, in practice it is rare for exact stoichiometric quantities to be used. Usually there is one reactant in excess compared to the other reactant.

*The reactant of which there is too much is said to be in **excess**.*

*The reactant that reacts completely is called the **limiting reagent**.*

When reaction problems are presented with masses of two reactants, it is necessary to determine which one is in excess and which is the limiting reagent.

*To do this, the **number of mole** of each reacting quantity must be calculated, then compared to the **ratio of mole required** by the reaction.*

Example

2.50 g of magnesium reacts with 3.85 g of hydrochloric acid to form hydrogen gas as shown below –

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

1. What is the limiting reagent?

n(Mg) = 2.50/24.30 = 0.1029 mol   M(Mg) = 24.30 g

n(HCl) = 3.85/36.458 = 0.1056 mol   M(HCl) = 36.458 g

According to the equation $\frac{n(Mg)}{n(HCl)} = \frac{1}{2}$

For this example $\frac{n(Mg)}{n(HCl)} = \frac{0.1029}{0.1056}$

The number of mole of Mg is too large for the number of mole of HCl - only need ½ x 0.1056 mole = 0.0528 mol. Therefore – Mg is in excess, and HCl is the limiting reagent.

2. What is the mass of excess reactant remaining?

You know how much the reaction starts with = 0.10239 mol

You worked out how much would be needed = 0.0528 mol

Mole remaining = 0.10239 – 0.0528 = 0.0501 mol

n(Mg) = 0.0501 mol

m(Mg) = 0.0501 x 24.30 M(Mg) = 24.30 g

= 1.22 g

3. What mass of magnesium chloride would be produced in the reaction?

You know the number of mole of limiting reagent = 0.1056 mol

This must determine the number of mole of product, as this is totally used in the reaction.

$n(MgCl_2)$ = ½ n(HCl) from the equation

= ½ x 0.1056

= 0.0528 mol

$m(MgCl_2)$ = 0.0528 x 95.20 $M(MgCl_2)$ = 95.20 g

= 5.03 g

Try these –

Sulfuric acid solution containing 20.0 g of pure H2SO4 was added to 6.08 g of magnesium to produce hydrogen gas –

$Mg + H_2SO_4 + 7H_2O \rightarrow H_2 + MgSO_4.7H_2O$

Identify the limiting reagent.

What mass of hydrogen is produced?

What mass of magnesium sulfate crystals ($MgSO_4.7H_2O$) would be obtained if the solution was evaporated to dryness?

1.600 g of sodium hydroxide is added to a solution containing 1.472 g of H_2SO_4.

$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

Identify the limiting reagent.

Calculate the mass of sodium sulfate produced in the reaction.

Calculate the mass of unused reactant in the reaction.

Tutorial 8 – Significant Figures

Chemistry 11

Significant figures reflect how accurately a measurement has been taken. The number of significant figures reflects all the measurements known to be true, plus the first digit that has been estimated (error present).

Most calculations expect you to answer to **3 significant figures**, though at times it will be appropriate to have 4 or possibly even 2, depending on the measurements that are taken.

Recognising significant figures

All non-zero digits are significant.

Zeros at the beginning of a number are not significant.

Zeros in between and after non-zeros are significant.

When not sure of the number of significant figures, simply convert the number into standard notation, and this should help.

For example

3.41	3 sf's	
2.008	4 sf's	
0.00065	2 sf's	6.5×10^{-4}
0.500	3 sf's	
600	1, 2 or 3 sf's – depends upon accuracy of measurement – this will usually be indicated.	
600.0	4 sf's	

Using significant figures in calculations

Your calculator is limited by the input you make and how it is set up to deal with measurements. On normal settings you will often get an answer that has too few or too many significant figures. But you have to remember that you are dealing with measurements which have error associated with these. The calculator will not take this into account, so you have to be careful when reading the calculator display.

Adding/subtracting

Take a 65 kg student and add a 100 g bag of rice.

What value would you expect on your calculator? ________________________

What value would you expect if you stand on bathroom scales? ______________

As the error is in the last significant figure and you cannot have two errors in a measurement – so you must stop recording numbers at the first error!

When measurements are added or subtracted, the final answer cannot be more accurate than the least accurate of the measurements taken originally.

Multiplying/dividing

When you multiply or divide you can end up with a lot of numbers. But not all of these will be significant. You cannot improve the accuracy of your measurements simply by multiplying or dividing.

E.g. we measure the length and breadth of the room as 6.7 m (using a rule with no graduations) by 10.124 m (using a metre rule).

On your calculator the area is _______________________

But – one of the measurements was only 2 significant figures – we cannot calculate the area of the room more accurately that that – so we have to round to -

____

When measurements are multiplied or divided, the resulting answer can only have as many significant figures as the original measurement with the least number of significant figures.

Tutorial 9 – Volume Calculations

Chemistry 11

Molar Volume

This is the volume occupied by one mole of any gas for a given temperature and pressure.

At **STP** – Standard Temperature (0^oC) and Pressure (100.0 kPa) – molar volume is 22.71 **L**

This means that if the volume at STP is known, then the number of mole can be calculated using the following formula –

$$\mathbf{n = \frac{V}{22.71}}$$

Examples

1. How many mole of oxygen gas is 56.0 L at STP?

   n = 56.0/22.71
   = 2.47 mol

2. What is the mass of nitrogen gas in 42.5 L at STP?

   n = 42.5/22.71
   = 1.87 mol
   m = 1.87 x 28.02
   = 52.4 g

3. What volume of hydrogen gas is produced at STP by 6.89 g of magnesium reacting in excess hydrochloric acid solution?

$$Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$$

n(Mg) = 6.89/24.31
= 0.283 mol

$n(H_2)$ = n(Mg)
= 0.283 mol

$V(H_2)$ = 0.283 x 22.71
= 6.44 L

Gay-Lussac's Law

Equal volumes of gases at the same temperature and pressure occupy the same volume.

Reacting volumes will be in the same ratio as the reacting mole for a chemical reaction, at constant temperature and pressure.

Use Gay-Lussac's Law when **only** volumes are given.

The Law only applies to **gases** – so be careful not to use it anywhere else.

Example

Carbon monoxide reacts with oxygen to produce carbon dioxide as follows –

$$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$

What volume of oxygen is needed to react with 24.0 L of carbon monoxide?

V(CO) = 24.0 L

Ratio is 2:1 therefore need ½ as much oxygen gas as carbon monoxide.

$V(O_2)$ = ½ V(CO)
= ½ x 24.0
= 12.0 L

Tutorial 10 - Solutions

Chemistry 11

Definitions –

Solution – a homogeneous mixture where one substance is dissolved in another

Solute – the substance that is dissolved

Solvent – the substance that does the dissolving

Saturated – a solution that contains the maximum amount of solute dissolved

Unsaturated – A solution that has less than the maximum amount dissolved and can dissolve more solute.

Supersaturated – a solution with more than the maximum amount dissolved. This is made by saturating a solution at high temperature, then cooling carefully in a smooth and clean container.

Solubility – the amount of solute needed per 100g of solution to make a saturated solution.

Solubility depends upon two major factors –
Temperature
Type of substance

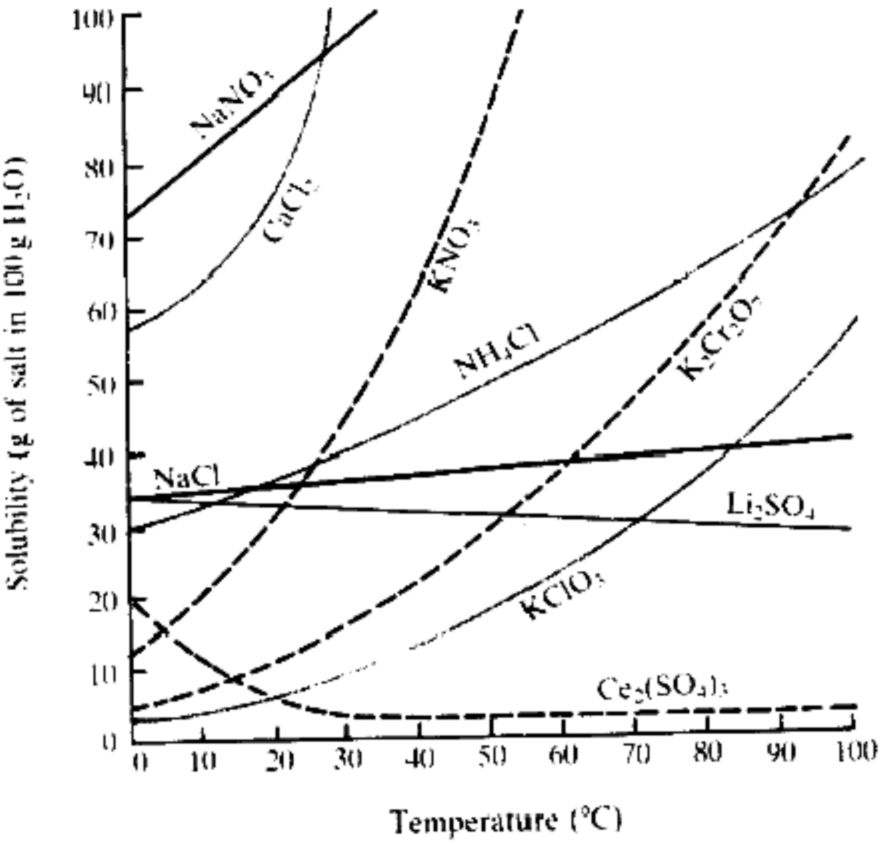
A substance will dissolve if there is sufficient energy to break the bonds between particles of solute and between particles of solvent, and if there is attraction between solvent and solute so that new bonds form.

Increasing the temperature, increases the amount of energy available to break the bonds within the solute, hence more can dissolve.

In general solids increase solubility with increased temperature.

For gases, the increase in temperature leads to more particles having the energy to leave the solution as gases, and the solubility decreases.

The solubility curve below show the relationship between temperature and solubility –



The range of solubility of various salts can be observed in the graphs above. Some substances show a dramatic increase in solubility with increased temperature – eg ______________________, some scarcely change at all, eg ____________________ and some actually decrease solubility (not easily explained) eg ___________________.

These solubility curves can be used for simple calculations –

1. What mass of solute will dissolve in 200 g of solution of potassium nitrate at 40 °C?

2. What mass of crystals will form if 100 g of saturated ammonium chloride solution is cooled from 80 °C to 10 °C?

3. At what temperature is the solubility of sodium nitrate and calcium chloride the same?

Solution Calculations

Solutions are made by dissolving a solute (usually a solid) in a solvent (usually water). The concentration of solutions can be determined in a number of ways –

Mole per litre

$$\text{Concentration} = \frac{\text{mole of solute (mol)}}{\text{volume of solution (L)}}$$

This is the standard measurement of concentration.

Grams per litre

$$\text{Concentration} = \frac{\text{mass of solute (g)}}{\text{volume of solution (L)}}$$

Often used for making up solutions – as a mass of solute must be weighed.

Examples

1. What mass of solute is needed to make 250 mL of 5.00 g L^{-1} sodium chloride?

2. 67.8 g of copper II nitrate is dissolved to make 0.500 L of solution. What is the concentration in g L^{-1} and mol L^{-1}?

Other Calculations

Using the formula –

$c = n/V$

A number of mole can be calculated if the volume and concentration for a solution are known. This can be extended into reaction calculations as shown in the following example –

What volume of 2.00 mol L^{-1} hydrochloric acid is needed to react with 24.5 mL of 0.400 mol L^{-1} barium hydroxide solution?

Reaction is –

$$2HC\ell + Ba(OH)_2 \rightarrow 2H_2O + BaC\ell_2$$

$n(Ba(OH)_2)$ = 0.400 x 0.0245
= 0.00980 mol

$n(HC\ell)$ = 2 x $n(Ba(OH)_2)$
= 2 x 0.00980
= 0.0196 mol

$V(HC\ell)$ = 0.0196/2.00
= 0.00980 L
= 9.80 mL

Try this example

What mass of barium chloride would be produced if 25.0 mL of 0.0598 molL^{-1} HCl is reacted with 20.0 mL of 0.0248 molL^{-1} barium hydroxide solution? (Note – you will need to find the limiting reagent first).

Tutorial 11 – Writing Ionic Equations

Chemistry 11

We use chemical equations to represent what is happening in a chemical reaction. It is logical then, to only include those chemical species that are actually involved. Anything else should be completely ignored.

Precipitation Reactions

When two solutions are combined, sometimes there is a combination of ions that is insoluble. This combination will form a solid that is observed as a precipitate (falling solid).

To write an equation for a precipitation reaction –

- Solubility rules are used to decide if a solid forms
- The formula for the solid is written on the product side of the equation
- The correct number and type of atoms used to make this solid are written on the reactant side.

Eg – silver nitrate solution is combined with sodium chloride solution

Ions are Ag^+, NO_3^-, Na^+, and Cl-

AgCl is insoluble

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

Other Reactions Involving Solutions

There are many other reactions that do not involve precipitation, but which occur in solution, and hence need to have ions considered.

For this type of reaction, it is important to write an equation showing all reactants and products, then eliminate those species that are not involved.

Use *"Livings Law"*

Only strong acids and ionics in solution are written as ions.

The steps to writing an equation are as follows

- Write all reactants and products with correct chemical formulas and state symbols.
- Use Livings Law to decide which species need to be ions.
- Rewrite the equation with ions where appropriate.
- Cancel out ions that do not change from reactant to product.
- Rewrite showing only the species that are involved in the reaction.
- Make sure the equation is balanced for both atoms and charge.

Example – solid calcium carbonate is reacted with hydrochloric acid

$CaCO_3(s)$ +	$HCl(aq)$ →	$CaCl_2(aq)$ +	$H_2O(l)$ +	$CO_2(g)$
Solid	Strong acid => ions	Ionic solution => ions	Liquid	Gas

$CaCO_3(s) + H^+(aq) + Cl^-(aq) \rightarrow Ca^{2+}(aq) + Cl^-(aq) + H_2O(l) + CO_2(g)$

$CaCO_3(s) + H^+(aq) + \cancel{Cl^-}(aq) \rightarrow Ca^{2+}(aq) + \cancel{Cl^-}(aq) + H_2O(l) + CO_2(g)$

The $Cl^-(aq)$ is common to both reactants and products – so is not involved. This is called a spectator ion.

$CaCO_3(s) + H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$

$CaCO_3(s) + \mathbf{2}H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$

Balancing atoms and charge is very important. Usually charge is balanced once the atoms are balanced, but this is not always the case – so watch out!