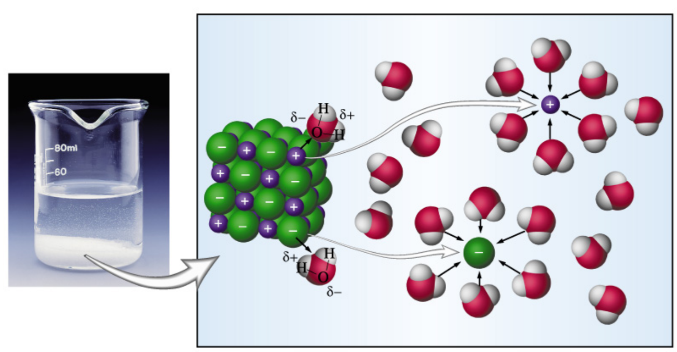
Solutions

A solution is a homogeneous mixture — it looks like one substance.
An aqueous solution will be a clear mixture with only one visible phase.
Be careful with the definitions of clear and colourless. Clear can be seen through. Colourless has no colour (pure water). A solution can be blue and clear.
Give examples of liquid solutions and liquids that are not solutions.

A solution consists of a solute that is dissolved in a solvent. In sugar water, sugar is the solute, and water is the solvent.
For two miscible liquids (then mix to form a solution) the liquid in the greater amount is the solvent.
For a solute to dissolve in a solvent, the solute must be broken into molecular or ionic sized pieces.

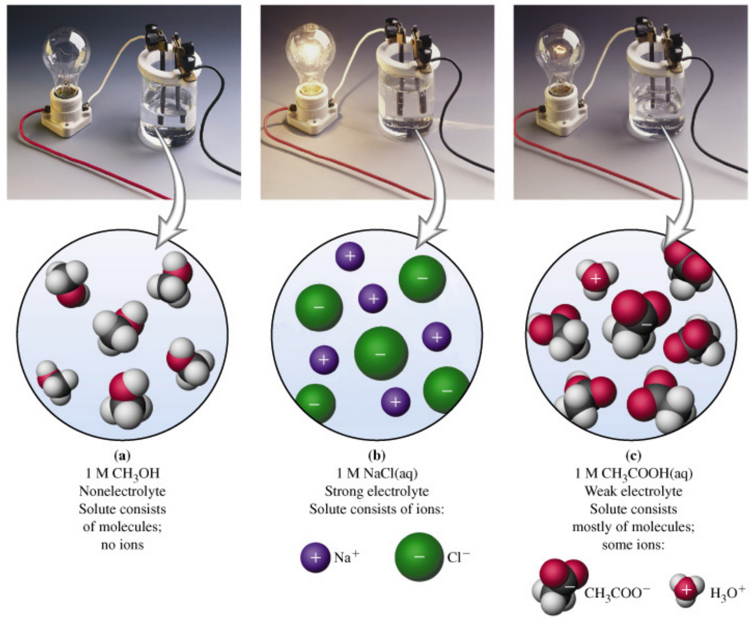
In water solutions, the subscript *(aq)* is put after the solute. This indicates that the molecules (or ions) of the solute have been pulled apart and are mixed with the water.
For a solute to be soluble in water, It must have an attraction to the water molecules that is stronger then the attraction the solute has for itself. (Between one molecule and another.)

Water molecules are polar molecules; they have a partially positive and a partially negative end.
Other molecules or ionic compounds that have polar parts, can dissolve in water.
Sugar and alcohols have O—H bonds that are polar and can dissolve in water. NaCl is an ionic compound that can dissolve in water.



Compounds that dissolve in water form weak bonds with the water molecules.
The attraction of water molecules forms a "hydration sphere."

The nature of the solution can be determined by the conductivity of the solution.
A solution conducts electricity through the movement of ions within the solution.
An ionic solution will conduct electricity readily. It is a strong electrolyte.



A covalent solution will not conduct electricity. It is a non-electrolyte
Fewer ions in solution means lower conductivity. A solution that conducts a little electricity is a weak electrolyte.

Example 1:
Predict the conductivity of the following:
Pure water

$NaCl_{(aq)}$

$C_{12}H_{24}O_{12(aq)}$ sucrose

$CH_3COOH_{(aq)}$ acetic acid

$Al(OH)_{(aq)}$

Diet coke

There are many ways to measure the concentration of a solution, but the most common is called Molarity.
Molarity equals the moles of solute per liter of solution
The symbol for molarity is:

$$M = \frac{\text{moles}}{L} = \text{kmol}\bullet\text{m}^{-3}$$

[NaCl] means the concentration of sodium chloride in moles per liter.

To make 1.0 L of a 1.0 M solution of NaCl:
Weight 1 mol (58.44 g)
Add less than 1 L of water and mix the salt until it dissolves.
Add enough water to make 1.0 L of solution.
Why all these steps?

A Graduated cylinder should be used to measure volumes of water precisely.
The best container to measure accurate, large volumes is a volumetric flask.
No reactions should be conducted in these pieces of glassware, because they are designed for accurate and precise measurements.

Ex 2:
Explain how you would make 250.0 mL of a 0.500 M solution of glucose ($C_6H_{12}O_6$)

Ex 3:
285 g of aluminium chloride is dissolved in enough water to make 2.50 L of solution. What is the concentration of this solution?

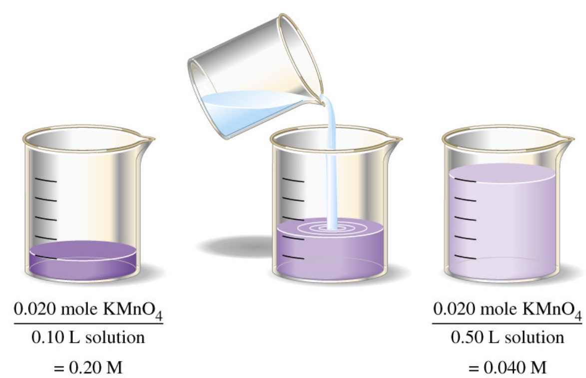
What is the concentration of chloride ions in this solution?

Ex 4:
What volume of a 0.250 M NaCl solution would be required to have 12.5 g of sodium chloride?

Follow Up Problems 3.14, 4.1
Problems 3.92, 94

2 — 2 Dilution

A solution can be diluted to decrease its concentration. The new solution has the same number of moles, but a different volume.



$\text{moles}_A = \text{moles}_B$

$M_A \text{vol}_A = M_B \text{vol}_B$

Ex 1:
250 mL of a 0.200 M solution is diluted to a new volume of 750 mL. What is the new concentration?

Ex 2:
How much of a 12 M HCl solution is required to make 2.00 L of a 0.500 M HCl solution?

Precipitation Reactions

A reaction of soluble ions that produces an insoluble product is called a precipitation reaction.

$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

$AgCl_{(s)}$ is the precipitate.

It is important to note that $NaCl_{(aq)}$ really means $Na^{+}_{(aq)}$ ions and $Cl^{-}_{(aq)}$ ions.

There are three ways to write these reactions:
Molecular equation assumes that ionic substances are molecules in solution:

$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$

This is false assumption, but you are expected to know that these compounds are ions in solution.

Complete ionic equation treats ionic solutions as separated ions and includes all the ions in the reaction:

$Ag^{+}{}_{(aq)} + NO_3^{-}{}_{(aq)} + Na^{+}{}_{(aq)} + Cl^{-}{}_{(aq)} \rightarrow AgCl_{(s)} + Na^{+}{}_{(aq)} + NO_3^{-}{}_{(aq)}$

3) The net ionic reaction only include the ions that change:

$Ag^{+}{}_{(aq)} + Cl^{-}{}_{(aq)} \rightarrow AgCl_{(s)}$

An ion that is not involved in the overall reaction (it is omitted in the net equation) is called a spectator ion.

A precipitation reaction can be predicted based on some simple rules.
Please see:
Table 4.1, p. 143

Ex 3:
Write the molecular, complete and net ionic equation for these mixtures:

Calcium chloride solution and sodium carbonate solution.

Barium iodide solution and copper sulfate.

Follow Up Problems 3.15, 4.3
Problems 3.96, 4.24, 28, 29, 31, 33

2 — 3 Acid Base Reaction Stoichiometry

An acid base reaction produces water and a salt.

Most of our acid base chemistry is aqueous, and most of our definitions tend to emphasize this.

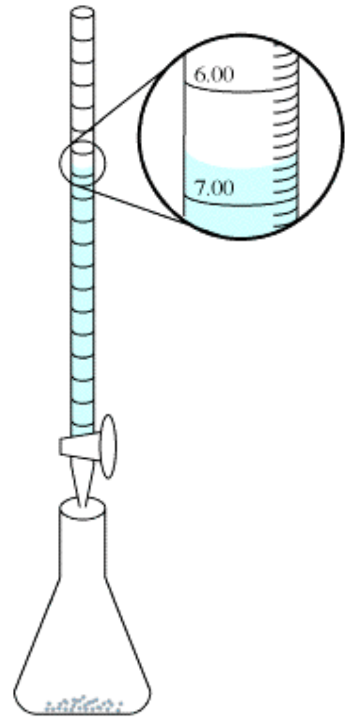
An acid is a chemical that can donate H^+ ions, and will produce $H^+_{(aq)}$ ions in a water solution.

A base will accept H^+ ions. In a water solution it will produce $OH^-_{(aq)}$ ions.

The H^+ ion from the acid and the OH^- ions from the base will produce water.
The anion from the acid and the cation from the base will produce an ionic compound: a salt.

For acid base reaction we usually right the molecular equation.

Acid base reaction also are a common reaction for a titration.



Setup for a titration of a solution against a solid.

A titration is a quantitative volumetric analysis technique.
A solution of a known concentration is used to determine the moles in an unknown solution.
Stoichiometry is used to determine the unknown.

Either the base or the acid is placed in the burette. The other is placed in the flask. The substance in the burette is added to the flask until the moles of acid equals the moles of base.

Burette: Used to add a variable amount of solution precisely.

Pipette: Used to add a specific amount of solution very precisely.

Indicator: Informs when the reaction is complete, often with a colour change.

Ex 1:
25.00 mL of an HCl solution is titrated with a0.100 M NaOH solution (the NaOH is in the burette). 32.62 mL of the NaOH solution was used to react with the HCl. What is the concentration of the HCl solution?

Ex 2:
100.0 mL of 0.0452 M aluminium hydroxide solution is to react with a 0.123M Sulfuric acid (H_2SO_4) solution. What volume of the acid solution will be needed?

Ex 3:
To determine the molar mass of a mono-protic acid, 0.0850 g of the acid is titrated with 24.80 mL of a 0.0355 M potassium hydroxide solution. What is the molar mass of the acid?

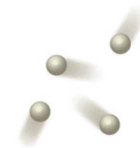
Ex 4: (This is like the upcoming lab)
A solution of oxalic acid ($C_2H_2O_4 \bullet 2H_2O$)is made with 0.625 g of oxalic acid in enough water to make 250.00 mL of solution.

25.00 mL of the oxalic acid solution (diprotic acid) is titrated with 27.50 mL of an unknown NaOH solution. What is the concentration of the NaOH solution?

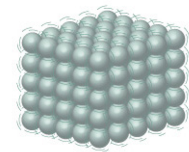
Follow Up Problems 4.5
Problems 3.100, 4.49, 50, 52, 113
Follow Up Problems 3.17

2 — 4 Solids Liquids and Gasses

From the Kinetic Molecular Theory, we know something of the arrangement of particles in each phase.
What do these diagrams indicate about each phase?


Gas


Liquid


Solid

The gas phase is the can alter its volume significantly. This makes density and molar calculations more complicated.

Gasses have been found to act very generically. There behavior has been described by the ideal gas law:

$PV = n\mathrm{R}T$

P = pressure (kPa)
V = volume (L)
n = moles
T = temperature (K)
R = ideal gas constant, 8.314 J mol^{-1} K^{-1}

This law is a reasonable model for most gasses.
Gasses are usually described at Standard Temperature and Pressure (STP): 273K, 101.3 kPa.

Ex 1:
What volume does a gas occupy at STP?

Ex 2:
What is the density of a sample of nitrogen gas if one mole is held at 455 kPa and 725K?

Because the gas law applies to all gasses, reaction molar ratios can be expressed as volume relationships if temperature and pressure are kept constant.

If 1 mole of nitrogen combine with 3 moles of hydrogen to produce 2 moles of ammonia.

Then, 10 L of nitrogen will combine with 30 L of hydrogen to produce 20 L of ammonia if the conditions are constant.

Another relationship derived from the ideal gas law is for a fixed amount of gas.
A fixed amount of gas will have *n* and R a constant.

$$\frac{PV}{T} = k$$

If the *P*, *V* or *T* change, they still must generate a constant.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Ex 3:
A sample of gas at STP occupies 15.0 L. The pressure is changed to 450 kPa and the temperature is lowered to 125K. What is the new volume?

Ex 4:
A 1.00 g sample of Calcium carbonate is decomposed to form Calcium oxide and carbon dioxide gas. The reaction is contained in a 10.0L container at 350K. The volume of the solid is negligible. What is the pressure of the gas collected?

Follow Up Problems 5.2, 3, 4, 5, 7, 11, 12

2 — 5 Energy in Reactions

Reactions evolve or consume energy.
Energy is described from the point of view of the reaction.
For a reaction, there is the system (reactant and products) and the surroundings (everything else)

If a reaction gives off energy the energy of the system decreases.
If the reaction consumes energy the energy of the system increases.

The concept of enthalpy (H) is used to describe the internal energy of a system.
If energy leaves the system, the reaction is exothermic and $\Delta H < 0$.
If energy enters the system, the reaction is endothermic and $\Delta H > 0$.
Ex 1:
Give examples of endothermic and exothermic reactions.

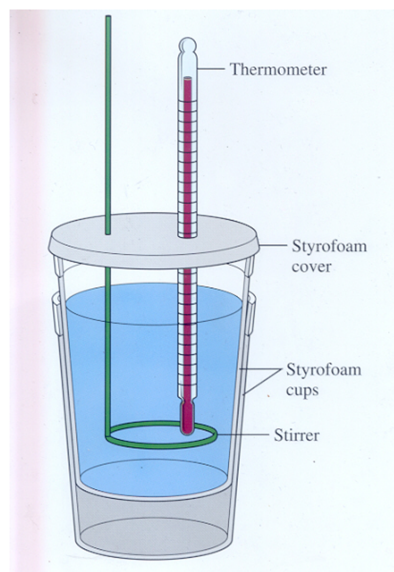
In a chemical reaction, the energy depends on the breaking and forming of bonds.
A chemical reaction requires the breaking of the old bonds in the reactants, and forming of new bonds in the products.

It requires energy (endothermic) to break bonds. Forming bonds is an exothermic process.

The overall energy of the reaction depends on both of these processes.

Ex 2:
Why are melting and boiling endothermic and condensation and freezing exothermic?

Experimentally we can measure the enthalpy of a reaction through calorimetry.
A reaction is carried out so that the energy of the reaction is measured.
The simplest way to do this is with a coffee cup calorimeter.



If the system releases energy, the water in the calorimeter absorbs the energy.
If the system consumes energy, the water is the calorimeter loses the energy.
The energy that water gains or loses can be measured by:

$$Q = mc\Delta T$$

m = mass of water; c = specific heat capacity of water
ΔT = change in temperature

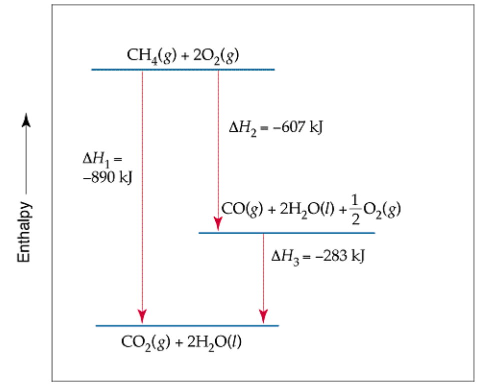
Specific heat capacity is the amount of energy required to change the temperature of a given amount of a substance.
For water c = 4.18 J g^{-1} K^{-1}
We assume the only the water in the calorimeter absorbs energy (perfect calorimeter)
The enthalpy of a reaction is usually calculated per mole of reactant.

Ex 3:
2.50 g of sugar ($C_{12}H_{24}O_{12}$) is combusted with excess oxygen in a calorimeter that contains 835 g of water. The water warms from 18.5 °C to 29.7 °C. What is the enthalpy per mole for the combustion of sugar?

Ex 4:
5.00 g of ammonium nitrate is dissolved into 0.400 L of water. The enthalpy of solution for ammonium nitrate is +176kJ/mol. What temperature change would you expect for the water in this calorimeter. (ρ_{water} = 1.00 g/mL)

Follow Up Problems 6.2, 3
Problems 6.19, 29, 33, 46, 48, 49, 52, 58

2 — 6 Hess's Law and Enthalpy of Formation



Enthalpies of reaction can be calculated theoretically. One method uses a principle called Hess's Law.
Enthalpy is a state function. Its value depends on the initial and final states, not how the process occurs.
If two or more reactions accomplish the same thing as one reaction, the overall enthalpies will be the same.

Multiple chemical reactions can be combined algebraically to find an overall reaction.

$$N_{2(g)} + 2O_{2(g)} \rightarrow 2NO_{2(g)} \qquad \Delta H = 68kJ$$

or

$$N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)} \qquad \Delta H = 180kJ$$
$$2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)} \qquad \Delta H = -112kJ$$
$$\overline{N_{2(g)} + 2O_{2(g)} \rightarrow 2NO_{2(g)} \qquad \Delta H = 68kJ}$$

To use Hess's Law:

- Reverse the sign of ΔH if the reactants and products are switched
- If a reaction is multiplied by an integer, the ΔH must also be multiplied by that integer. (The ΔH value is for the written reaction)

Ex 1:
Given

A + 2B → 2C   ΔH = 25 kJ
B + E → C + D   ΔH = -15 kJ

Calculate ΔH for: A + 2D → 2E

A second theoretical method to calculate enthalpy is through Enthalpies of formation.

Standard Enthalpy of Formation (ΔH°_{f}) is the enthalpy of the reaction that produces a substance from its elements in their standard standard states. (The "°" indicates the standard states)

Ex 2:
Write the reaction for the enthalpy of formation for the following and find their values in the appendix of your text or your IB Data booklet.

a) $H_2O_{(l)}$

b) $CaCO_{3(s)}$

c) $CH_3COOH_{(l)}$ (ethanoic acid)

d) $N_{2(g)}$

The enthalpy of a reaction can be calculated from the enthalpies of formation of the products and the reactants.

$$\Delta H^{\circ}_{rxn} = \sum n \Delta H^{\circ}_{f}(\text{products}) - \sum n \Delta H^{\circ}_{f}(\text{reactants})$$

Ex 3:
Calculate the enthalpy for the following reactions:
$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$

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

$C_2H_{4(g)} + H_{2(g)} \rightarrow C_2H_{6(g)}$

Follow Up Problems 6.7, 8, 9
Problems 6.62, 63, 65, 70, 72, 73, 75, 77