Chemistry Reference Tables

Table A Standard Temperature and Pressure

Name	Value	Unit
Standard Pressure	101.3 kPa 1 atm	kilopascal atmosphere
Standard Temperature	273 K 0°C	kelvin degree Celsius

Temperature is a measure of the average kinetic energy. The higher the temperature, the higher the kinetic energy.

This table provides the values for standard temperature and pressure. These values may be needed when working with the combined gas law found on reference table T.

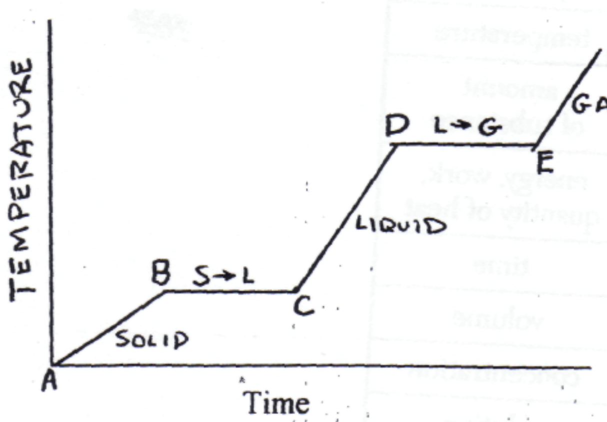
Combined Gas Law $$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Table B Physical Constants for Water

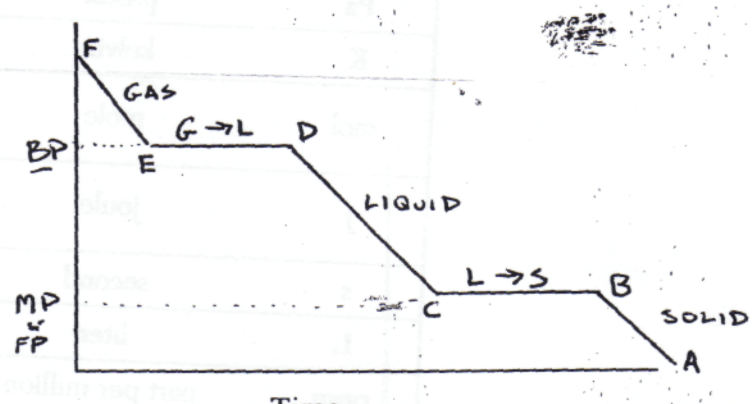
Heat of Fusion	333.6 J/g
Heat of Vaporization	2259 J/g
Specific Heat Capacity of H_2O (ℓ)	4.2 J/g•K

These are the values to be used in combination with a heating/cooling curve and the equations for "heat" listed on reference table T. A heating curve represents a substance being heated at a uniform rate while a cooling curve represents a substance cooled at a uniform rate.

Heating Curve



Cooling curve



Equations

Heat of fusion(line BC)	Specific Heat Capacity(line CD)	Heat of Vaporization(line DE)
$Q = m H_f$	$Q = m c \Delta T$	$Q = m H_v$

Table C **Selected Prefixes**

Factor	Prefix	Symbol
10^{3}	kilo-	k
10^{-1}	deci-	d
10^{-2}	centi-	c
10^{-3}	milli-	m
10^{-6}	micro-	μ
10^{-9}	nano-	n
10^{-12}	pico-	p

This table contains commonly used metric prefixes. For example;

1 kilogram = 1000 grams or 1×10^{3} grams

1 centimeter = 1/100 meter or 1×10^{-2} meter (1 meter = 100 centimeters)

1 milliliter = 1/1000 liter or 1×10^{-3} liter (1 liter = 1000 milliliters)

The atomic radius is measured in picometers (pm).

Table D **Selected Units**

Symbol	Name	Quantity
m	meter	length
kg	kilogram	mass
Pa	pascal	pressure
K	kelvin	temperature
mol	mole	amount of substance
J	joule	energy, work, quantity of heat
s	second	time
L	liter	volume
ppm	part per million	concentration
M	molarity	solution concentration

This table lists the name and symbol for various units of measurement that have been used throughout the year. When taking the Regents exam, use this chart to recognize any unit that is unfamiliar.

Table E **Selected Polyatomic Ions**

H_3O^+	hydronium	CrO_4^{2-}	chromate
Hg_2^{2+}	dimercury (I)	$Cr_2O_7^{2-}$	dichromate
NH_4^+	ammonium	MnO_4^-	permanganate
$C_2H_3O_2^-$ / CH_3COO^-	acetate	NO_2^-	nitrite
		NO_3^-	nitrate
CN^-	cyanide	O_2^{2-}	peroxide
CO_3^{2-}	carbonate	OH^-	hydroxide
HCO_3^-	hydrogen carbonate	PO_4^{3-}	phosphate
$C_2O_4^{2-}$	oxalate	SCN^-	thiocyanate
ClO^-	hypochlorite	SO_3^{2-}	sulfite
ClO_2^-	chlorite	SO_4^{2-}	sulfate
ClO_3^-	chlorate	HSO_4^-	hydrogen sulfate
ClO_4^-	perchlorate	$S_2O_3^{2-}$	thiosulfate

A polyatomic ion is a group of atoms that has a charge. A + sign after the formula of the ion indicates a +1 charge and a - sign after the formula of the ion indicates -1 charge. In a similar way 2- is the same as a -2 charge.

This chart will be used when writing the formulas or names for compounds containing polyatomic ions. See the examples below.

You can use either form of the acetate ion. Be careful. Some of the ions have very similar formulas and names such as nitrate and nitrite.

Examples:

Name	**Formula**	**Acid/Base ions**
Calcium hydroxide	$Ca(OH)_2$	*Acidic ion*
Potassium carbonate	K_2CO_3	Hydronium H_3O^+
Aluminum acetate	$Al(C_2H_3O_2)_3$	*Basic ion*
Sodium phosphate	Na_3PO_4	Hydroxide OH^-
Ammonium chloride	NH_4Cl	

Table F Solubility Guidelines

Ions That Form *Soluble* Compounds	Exceptions
Group 1 ions (Li^+, Na^+, etc.)	
ammonium (NH_4^+)	
nitrate (NO_3^-)	
acetate ($C_2H_3O_2^-$ or CH_3COO^-)	
hydrogen carbonate (HCO_3^-)	
chlorate (ClO_3^-)	
perchlorate (ClO_4^-)	
halides (Cl^-, Br^-, I^-)	when combined with Ag^+, Pb^{2+}, and Hg_2^{2+}
sulfates (SO_4^{2-})	when combined with Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, and Pb^{2+}

Ions That Form *Insoluble* Compounds	Exceptions
carbonate (CO_3^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
chromate (CrO_4^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
phosphate (PO_4^{3-})	when combined with Group 1 ions or ammonium (NH_4^+)
sulfide (S^{2-})	when combined with Group 1 ions or ammonium (NH_4^+)
hydroxide (OH^-)	when combined with Group 1 ions, Ca^{2+}, Ba^{2+}, or Sr^{2+}

This chart is used to determine the solubility of ionic compounds in water. When given an ionic compound, determine the two ions present and look them up on the tables. **The table on the left contains ions that form soluble compounds (those that do dissolve in water). Only the** halides and sulfates have exceptions, those that are insoluble (do not dissolve in water).

Examples:

NaCl is soluble but AgCl is insoluble

K_3PO_4 is soluble but $CaSO_4$ is insoluble

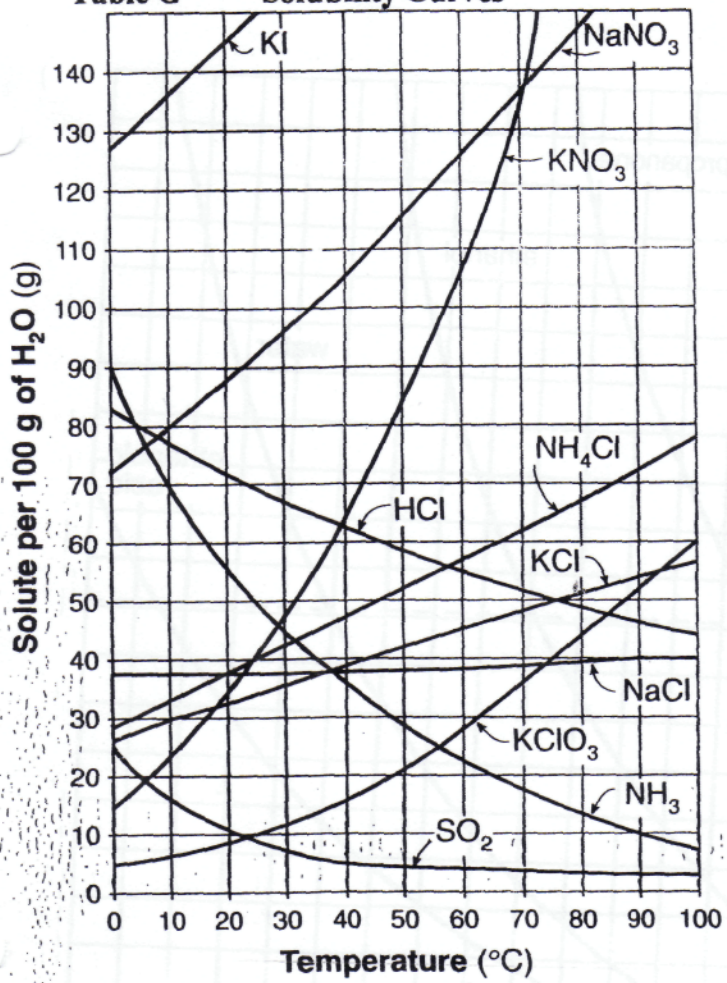
The table on the right contains ions that form insoluble compounds. The exceptions in this table are those compounds that are soluble in water.

This chart is also used to determine if a double replacement reaction will take place. If one, or both, of the products in a double replacement reaction are insoluble the reaction will take place.

Example: The following reaction will occur spontaneously.

Sodium bromide	+	Silver nitrate	→	Sodium nitrate	+	Silver bromide
NaBr(aq)	+	$AgNO_3$(aq)	→	$NaNO_3$(aq)	+	AgBr(s)
(soluble)		(soluble)		(soluble)		(insoluble)

Table G Solubility Curves



A solution is a homogeneous mixture. A solution consists of two parts, a solute and a solvent. The **solute** is the part of the solution that gets dissolved. The **solvent** is the part that does the dissolving.

The compounds listed on the graph represent the solutes. Water is the solvent.

An example of a solution is
NaCl(aq) (aq) = aqueous
This means water is the solvent.

Remember the phrase **"like dissolves like"**. This means that polar solutes dissolve best in polar solvents, nonpolar solutes in nonpolar solvents.

Thus NH3 dissolves well in water since they are both polar. Organic compounds are usually nonpolar and do not dissolve in water.

This table shows the solubility of various compounds in water. The number grams of solute that can dissolve in 100 g of water will depend on the temperature.

Example: How many grams of KCl can dissolve in 100 g of water at 60 ^{o}C?

Answer: Start at the bottom of the graph at 60 ^{o}C. Move up until you reach the curve for KCl. Then move across to determine the number of grams. The answer is 45g.

Some questions may involve more or less than 100 g of solvent. For these questions, read the graph as described above and then adjust the answer according to the amount of water.

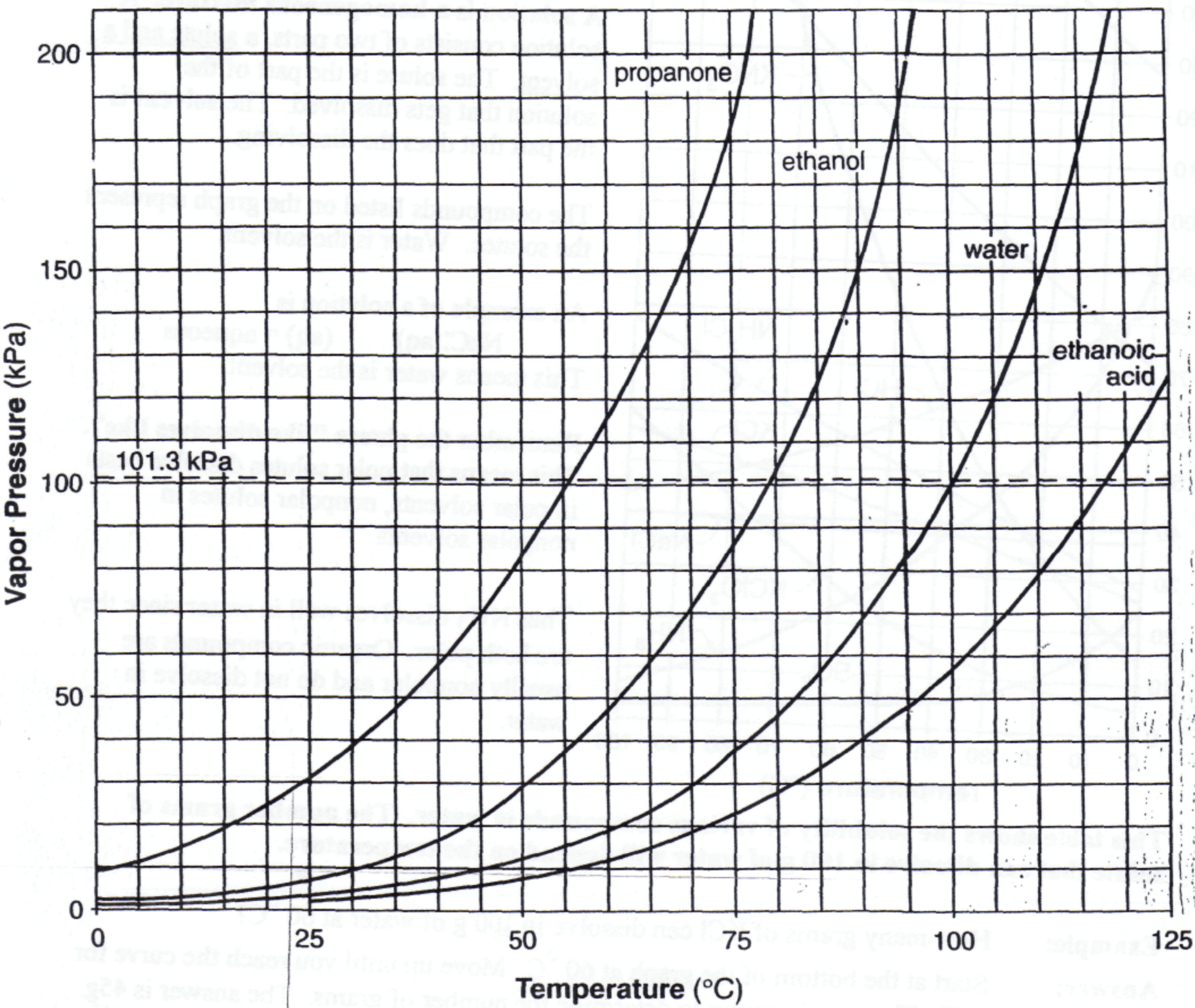
Example: How many grams of NH_3 can dissolve in 50 g of water at 10 ^{o}C ?

Answer: 70 grams dissolve in 100 g of water. Thus the amount that can dissolve in only 50 grams of water will be half that amount. The answer is 35. (If 200 grams of water were present you would double the amount to 140g.)

The type of solution may also be determined. Any point directly on the curve for a compound will be saturated, any point below a particular curve will be unsaturated, and above the curve is supersaturated.

Most curves show an increase in solubility with an increase in temperature. For solid substances, an increase in temperature results in an increase in solubility. For gases, an increase in temperature results in a decrease in solubility. SO_2, NH_3, and HCl are gases.

Table H **Vapor Pressure of Four Liquids**



This graph shows the vapor pressure of four different liquids as a function of temperature. The relationship between temperature and vapor pressure is direct. **The higher the temperature of a liquid, the higher the vapor pressure.**

The dashed line that cuts across the table represents standard pressure (101.3 kPa). **The point at which any of the four curves intersects the dashed line will represent the normal boiling point of the liquid.** For example the boiling point of propanone is 56 ^{o}C and the normal boiling point of water is 100 ^{o}C.

The relative force of attraction of these four liquids can also be determined. Propanone shows the greatest increase in vapor pressure as temperature increases. Propanone has the weakest force of attraction between molecules. Ethanoic acid shows the least increase in vapor pressure as temperature increase. Ethanoic acid has the strongest force of attraction

Summary **Propanone - weakest force of attraction and lowest normal boiling point**
Ethanoic acid - strongest force of attraction and highest normal boiling point

Table I **Heats of Reaction at 101.3 kPa and 298 K**

Reaction	ΔH (kJ)*
$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(\ell)$	–890.4
$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(\ell)$	–2219.2
$2C_8H_{18}(\ell) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(\ell)$	–10943
$2CH_3OH(\ell) + 3O_2(g) \longrightarrow 2CO_2(g) + 4H_2O(\ell)$	–1452
$C_2H_5OH(\ell) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(\ell)$	–1367
$C_6H_{12}O_6(s) + 6O_2(g) \longrightarrow 6CO_2(g) + 6H_2O(\ell)$	–2804
$2CO(g) + O_2(g) \longrightarrow 2CO_2(g)$	–566.0
$C(s) + O_2(g) \longrightarrow CO_2(g)$	–393.5
$4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$	–3351
$N_2(g) + O_2(g) \longrightarrow 2NO(g)$	+182.6
$N_2(g) + 2O_2(g) \longrightarrow 2NO_2(g)$	+66.4
$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$	–483.6
$2H_2(g) + O_2(g) \longrightarrow 2H_2O(\ell)$	–571.6
$N_2(g) + 3H_2(g) \longrightarrow 2NH_3(g)$	–91.8
$2C(s) + 3H_2(g) \longrightarrow C_2H_6(g)$	–84.0
$2C(s) + 2H_2(g) \longrightarrow C_2H_4(g)$	+52.4
$2C(s) + H_2(g) \longrightarrow C_2H_2(g)$	+227.4
$H_2(g) + I_2(g) \longrightarrow 2HI(g)$	+53.0
$KNO_3(s) \xrightarrow{H_2O} K^+(aq) + NO_3^-(aq)$	+34.89
$NaOH(s) \xrightarrow{H_2O} Na^+(aq) + OH^-(aq)$	–44.51
$NH_4Cl(s) \xrightarrow{H_2O} NH_4^+(aq) + Cl^-(aq)$	+14.78
$NH_4NO_3(s) \xrightarrow{H_2O} NH_4^+(aq) + NO_3^-(aq)$	+25.69
$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$	+3.88
$LiBr(s) \xrightarrow{H_2O} Li^+(aq) + Br^-(aq)$	–48.83
$H^+(aq) + OH^-(aq) \longrightarrow H_2O(\ell)$	–55.8

*Minus sign indicates an exothermic reaction.

This table list specific balanced equations and their heat of reaction (ΔH) values. The heat of reaction is the difference in potential energy between the products and the reactants in an equation. Negative numbers are for exothermic reactions, positive for endothermic. The top 6 equations represent complete combustion. The next 12 are synthesis (combination) reactions. The next six are dissociation (s to aq).

Example: How many kilojoules (kJ) are released when one mole of NH_3 is produced?

Answer: Since 2 moles of NH_3 are produced in the equation, divide -91.8 by two. **(-45.9)**

Example: How many kJ are absorbed when 3 moles of $C_2H_4(g)$ are produced?

Answer: One mole is produced in the equation, so multiply the ΔH value by 3. **(157.2)**

Table J Activity Series

Most likely to oxidize

Li is the strongest reducing agent on this chart.

(A reducing agent gets oxidized.)

Most li ly to reduce

I_2 is leas likely to reduce.

F_2 is the best oxidiz agent liste

(An oxidizi agent gets reduced)

Most ↓	Metals	Nonmetals	Most ↓
	Li	F_2	
	Rb	Cl_2	
	K	Br_2	
	Cs	I_2	
	Ba		
	Sr		
	Ca		
	Na		
	Mg		
	Al		
	Ti		
	Mn		
	Zn		
	Cr		
	Fe		
	Co		
	Ni		
	Sn		
	Pb		
	**H_2		
	Cu		
	Ag		
	Au		
Least			Least

Least likely to oxidize

**Activity Series based on hydrogen standard

This chart lists both metals and nonmetals in order of reactivity. The most reactive are on the top. Metals react by losing electrons and nonmetals react by gaining electrons.

This chart is used to predict whether a single replacement reaction will occur.

Compare metals to metals and nonmetals to nonmetals.

$$2\ Al(s) + 3\ CuCl_2(aq) \rightarrow 2\ AlCl_3(aq) + 3\ Cu(s)$$

The above reaction will occur because Al is listed higher than Cu on the table.

Any metal above H_2 on the table will react with hydrochloric acid to form hydrogen gas. (Cu, Ag, and Au do not react with hydrochloric acid and are considered non-spontaneous reactions.)

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

Table K Common Acids

Formula	Name
$HCl(aq)$	hydrochloric acid
$HNO_3(aq)$	nitric acid
$H_2SO_4(aq)$	sulfuric acid
$H_3PO_4(aq)$	phosphoric acid
$H_2CO_3(aq)$ or $CO_2(aq)$	carbonic acid
$CH_3COOH(aq)$ or $HC_2H_3O_2(aq)$	ethanoic acid (acetic acid)

The top three acids on this chart are strong acids. This means they ionize very well. In other words, they produce hydrogen ions (H^+) in solution. Strong acids are good electrolytes. The bottom three are weak acids. They produce few hydrogen ions and are considered weak electrolytes.

Arrhenius acid - produces hydrogen ions as as the only positive ion in solution

Bronsted - Lowry acid - proton (H^+) donor

Table L Common Bases

Formula	Name
$NaOH(aq)$	sodium hydroxide
$KOH(aq)$	potassium hydroxide
$Ca(OH)_2(aq)$	calcium hydroxide
$NH_3(aq)$	aqueous ammonia

The top three bases on this chart are strong bases. This means they are soluble in water and thus dissociate to produce hydroxide (OH^-) in solution. Strong bases are good electrolytes. Aqueous ammonia is a weak base. It produces few hydroxide ions and is considered a weak electrolyte

Arrhenius base - produces hydroxide ions the only negative ion in solution

Bronsted - Lowry base - proton (H^+) acceptor

Table M Common Acid-Base Indicators

An indicator is a substance that can determine the presence of acids and bases through specific color changes. This table lists 6 specific indicators.

Example: **methyl orange**
At pH < 3.2 the color is red
At pH > 4.4 the color is yellow
In the range of 3.2 to 4.4 the color will be a mix of red and yellow. (orange)

Indicator	Approximate pH Range for Color Change	Color Change
methyl orange	3.2–4.4	red to yellow
bromthymol blue	6.0–7.6	yellow to blue
phenolphthalein	8.2–10	colorless to pink
litmus	5.5–8.2	red to blue
bromcresol green	3.8–5.4	yellow to blue
thymol blue	8.0–9.6	yellow to blue

Remember:

Phenolphthalein is pink in a base.

Litmus paper is red in an acidic solution and blue in a basic, or alkaline, solution.

Table N Selected Radioisotopes

Nuclide	Half-Life	Decay Mode	Nuclide Name
^{198}Au	2.69 d	β^-	gold-198
^{14}C	5730 y	β^-	carbon-14
^{37}Ca	175 ms	β^+	calcium-37
^{60}Co	5.26 y	β^-	cobalt-60
^{137}Cs	30.23 y	β^-	cesium-137
^{53}Fe	8.51 min	β^+	iron-53
^{220}Fr	27.5 s	α	francium-220
^{3}H	12.26 y	β^-	hydrogen-3
^{131}I	8.07 d	β^-	iodine-131
^{37}K	1.23 s	β^+	potassium-37
^{42}K	12.4 h	β^-	potassium-42
^{85}Kr	10.76 y	β^-	krypton-85
^{16}N	7.2 s	β^-	nitrogen-16
^{19}Ne	17.2 s	β^+	neon-19
^{32}P	14.3 d	β^-	phosphorus-32
^{239}Pu	2.44×10^4 y	α	plutonium-239
^{226}Ra	1600 y	α	radium-226
^{222}Rn	3.82 d	α	radon-222
^{90}Sr	28.1 y	β^-	strontium-90
^{99}Tc	2.13×10^5 y	β^-	technetium-99
^{232}Th	1.4×10^{10} y	α	thorium-232
^{233}U	1.62×10^5 y	α	uranium-233
^{235}U	7.1×10^8 y	α	uranium-235
^{238}U	4.51×10^9 y	α	uranium-238

ms = milliseconds; s = seconds; min = minutes;
h = hours; d = days; y = years

Carbon - 14 undergoes a beta decay (β^-)

$$^{14}_{6}C \rightarrow {}^{14}_{7}N + {}^{0}_{-1}e$$

beta particle

Radium-226 undergoes an alpha decay (α)

$$^{226}_{88}Ra \rightarrow {}^{222}_{86}Rn + {}^{4}_{2}He$$

alpha particle

Iron - 53 undergoes positron emission (β^+)

$$^{53}_{26}Fe \rightarrow {}^{53}_{25}Mn + {}^{0}_{+1}e$$

positron

* All of the above are examples of natural transmutations.

The specific isotopes listed on this chart are radioactive. This means they will undergo a spontaneous decay of the nucleus. They decay mode symbol can be looked up on Table O. All elements with an atomic number greater than 83 will be radioactive.

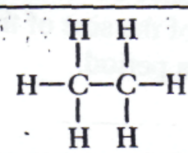
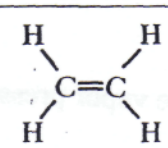
The half-life is the amount of time required for exactly one half of the nuclei in a radioactive sample to decay. The shorter the half-life the faster a substance decays. The half-life is not affected by temperature or pressure.

Table P Organic Prefixes

Prefix	Number of Carbon Atoms
meth-	1
eth-	2
prop-	3
but-	4
pent-	5
hex-	6
hept-	7
oct-	8
non-	9
dec-	10

This chart is used to determine the number of carbon atoms in an organic compound. It is used in combination with tables Q and R. See below.

Table Q Homologous Series of Hydrocarbons

Name	General Formula	Examples	
		Name	**Structural Formula**
alkanes	C_nH_{2n+2}	ethane	
alkenes	C_nH_{2n}	ethene	
alkynes	C_nH_{2n-2}	ethyne	H—C≡C—H

n = number of carbon atoms

There are three types of hydrocarbons listed on this table.
Alkanes - contain only single bonds and are saturated compounds
Alkenes - contain one carbon to carbon double bond and are unsaturated compounds
Alkynes - contain one carbon to carbon triple bond and are unsaturated compounds

Sample names and formulas for the first three members of each family:

Methane CH_4	Ethane C_2H_6	Propane C_3H_8
Ethene C_2H_4	Propene C_3H_6	Butene C_4H_8
Ethyne C_2H_2	Propyne C_3H_4	Butyne C_4H_6

A homologous series contain compound with similar structures and properties. Each member Differs from the next by a definite increment. For these hydrocarbons they differ by CH2.

Table S Properties of Selected Elements

This table lists information for selected elements. Some definitions of terms found on this chart are given below.

Ionization Energy (kJ/mole)

This is the energy required to remove the most loosely bound electron from an atom. The higher the ionization energy, the harder it is to remove an electron. Noble gases have the highest ionization energy within any period.

Electronegativity

This is a measure of the attraction an atom has for electrons in a bond. The higher the electronegativity, the stronger the attraction is for electrons. Fluorine has the highest electronegativity. The halogens have the highest electronegativity within any period. Noble gases have no values for electronegativity.

Atomic Radius (pm)

This is one half the distance between the adjacent nuclei of atoms in the solid phase. It is basically an indication of the size of the atom. Atomic radius increases going down a group and decreases going across a period.

Boiling Point (K)

Boiling occurs when the vapor pressure of the liquid is equal to the air pressure.

Melting Point (K)

This is the same temperature as the freezing point.

Density (g/cm^3)

Density is defined as mass/volume. Gases have low density since there is a much larger distance between molecules than in either the solid or liquid phase.

Table R Organic Functional Groups

Class of Compound	Functional Group	General Formula	Example
halide (halocarbon)	$-$F (fluoro-) $-$Cl (chloro-) $-$Br (bromo-) $-$I (iodo-)	$R-X$ (X represents any halogen)	$CH_3CHClCH_3$ 2-chloropropane
alcohol	$-OH$	$R-OH$	$CH_3CH_2CH_2OH$ 1-propanol
ether	$-O-$	$R-O-R'$	$CH_3OCH_2CH_3$ methyl ethyl ether
aldehyde	$-\overset{O}{\overset{\Vert}{C}}-H$	$R-\overset{O}{\overset{\Vert}{C}}-H$	$CH_3CH_2\overset{O}{\overset{\Vert}{C}}-H$ propanal
ketone	$-\overset{O}{\overset{\Vert}{C}}-$	$R-\overset{O}{\overset{\Vert}{C}}-R'$	$CH_3\overset{O}{\overset{\Vert}{C}}CH_2CH_2CH_3$ 2-pentanone
organic acid	$-\overset{O}{\overset{\Vert}{C}}-OH$	$R-\overset{O}{\overset{\Vert}{C}}-OH$	$CH_3CH_2\overset{O}{\overset{\Vert}{C}}-OH$ propanoic acid
ester	$-\overset{O}{\overset{\Vert}{C}}-O-$	$R-\overset{O}{\overset{\Vert}{C}}-O-R'$	$CH_3CH_2\overset{O}{\overset{\Vert}{C}}OCH_3$ methyl propanoate
amine	$-\overset{\vert}{N}-$	$R-\overset{R'}{\overset{\vert}{N}}-R''$	$CH_3CH_2CH_2NH_2$ 1-propanamine
amide	$-\overset{O}{\overset{\Vert}{C}}-\overset{\vert}{N}H$	$R-\overset{O}{\overset{\Vert}{C}}-\overset{R'}{\overset{\vert}{N}}H$	$CH_3CH_2\overset{O}{\overset{\Vert}{C}}-NH_2$ propanamide

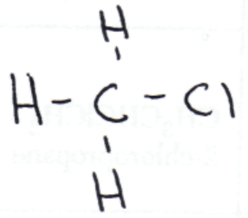
R represents a bonded atom or group of atoms.

A functional group is an atom or group of atoms that gives an organic compound specific properties. It can be used to identify the class of compound present in a structural or condensed formula.

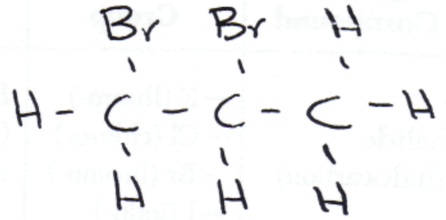
Sample Compounds - Names and structural formulas

Halide (halocarbon)

Chloromethane



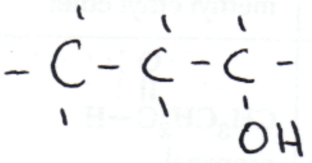
1,2-dibromopropane



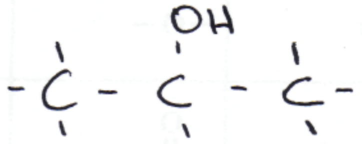
Alcohol

- OH

1-propanol



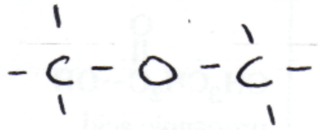
2-propanol (These two are isomers)



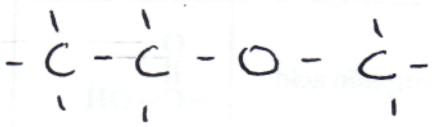
Ether

C-O-C

dimethyl ether



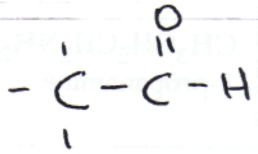
ethyl methyl ether



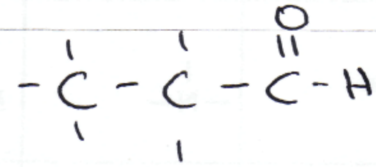
Aldehyde

-CHO

ethanal

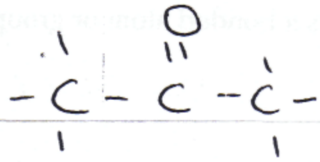


propanal

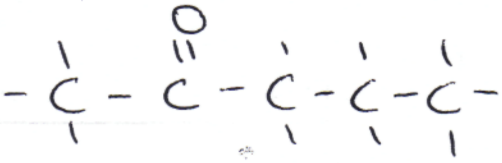


Ketone

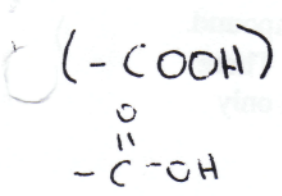
propanone

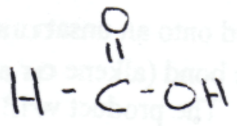


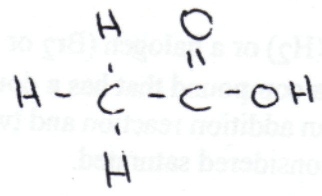
2-pentanone



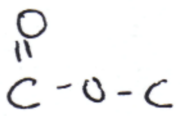
Organic Acid | **methanoic acid** | **ethanoic acid**

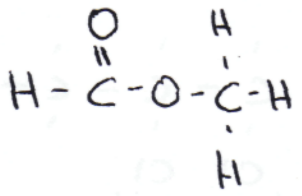


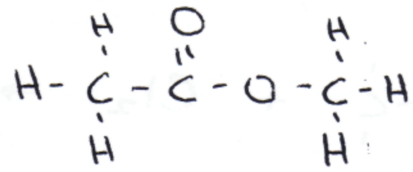




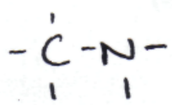
Ester | **methyl methanoate** | **methyl ethanoate**

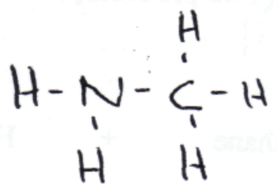


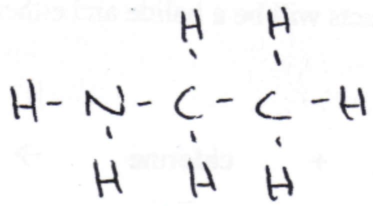




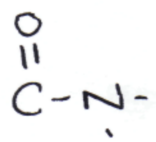
Amine | **methyl amine (1-methamine)** | **ethyl amine (1-ethanamine)**

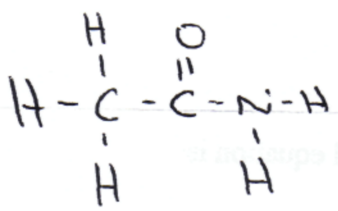


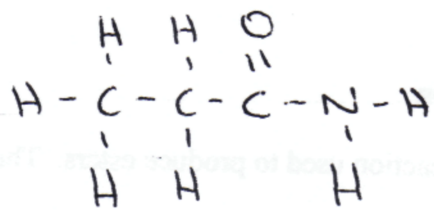




Amide | **ethanamide** | **propanamide**

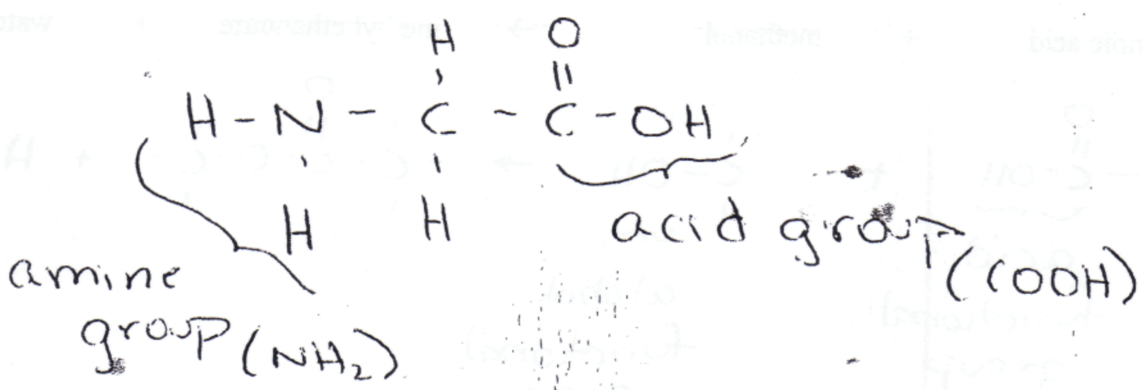






Amino Acid - **contain an acid functional group and an amine functional group**

The names are not required for amino acids.

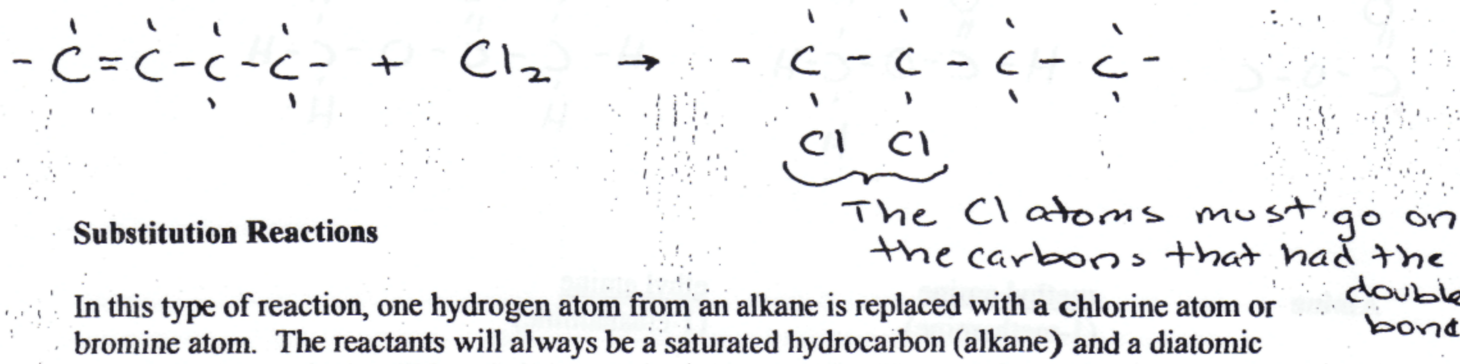


Addition reactions

In this reaction hydrogen (H_2) or a halogen (Br_2 or Cl_2) is added onto an unsaturated compound. Thus you must start with a compound that has a double or triple bond (alkene or alkyne). There is always one product in an addition reaction and two reactants. The product will contain only single bonds and thus is considered saturated.

Sample reaction

1-butene + chlorine → 1,2-dichlorobutane



Substitution Reactions

In this type of reaction, one hydrogen atom from an alkane is replaced with a chlorine atom or bromine atom. The reactants will always be a saturated hydrocarbon (alkane) and a diatomic element. The products will be a halide and either HCl or HBr. (Two products, two reactants)

Sample reaction

methane + chlorine → chloromethane + HCl

Esterification

This is the reaction used to produce esters. The general equation is

Organic acid + alcohol → ester + water

Let's look at a specific reaction.

Ethanoic acid + methanol → methyl ethanoate + water