

ISSN 1327-6743

Prepared by the
Science Teachers' Association
of Western Australia (Inc.)



Chemistry

1996 TEE Solutions*



Production, distribution and sales:

Science Teachers' Association
of Western Australia (Inc.)
PO Box 1099
Osborne Park WA 6916

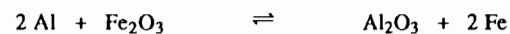
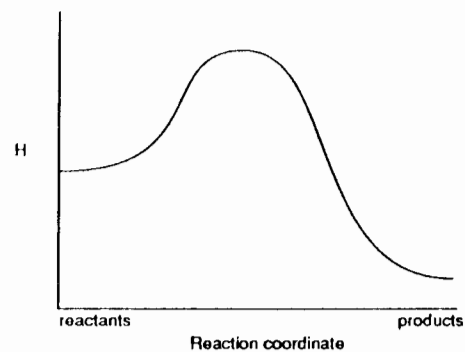
***Question papers and solutions can be
obtained from:***

Secondary Education Authority
27 Walters Drive
Osborne Park WA 6017

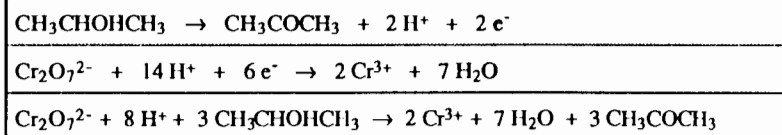
** These solutions are not a marking key. They are a guide to possible answers at a depth that might be expected of Year 12 students. It is unlikely that all possible answers to the questions are covered in these solutions.*

© 1997 The Science Teachers' Association of Western Australia. STAWA appreciates the support of the Secondary Education Authority, which provided the marker's guides used as a basis for these solutions.

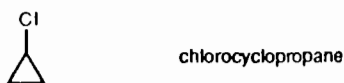
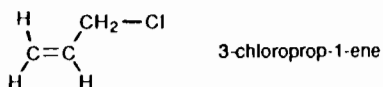
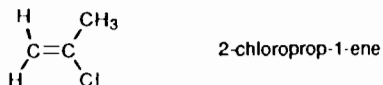
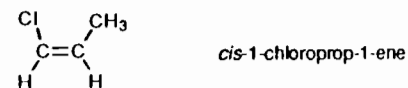
6.



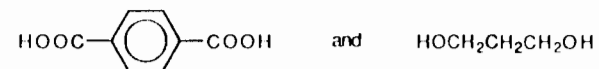
7.



8.



9.



10.

'They have been protonated' or
 $\text{CH}_3(\text{CH}_2)_3\text{NH}_2 + \text{H}^+ \rightarrow \text{CH}_3(\text{CH}_2)_3\text{NH}_3^+$

'Ions do not readily evaporate' or
 'The charge holds the species in the solution'.

'The species have been deprotonated' or
 $\text{CH}_3(\text{CH}_2)_3\text{NH}_3^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{CH}_3(\text{CH}_2)_3\text{NH}_2$

'All the volatile amine must be removed before the stink is lost, but only a little needs to be present for the stink to be observed'.

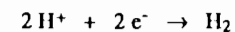
11.

→	solution becomes more blue
←	white precipitate forms <u>and</u> solution becomes more red
→	solution becomes more red <u>and</u> solution becomes more blue

12.

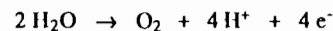
There is a larger surface area, and so there are more collisions between species.
The molecules are closer together, and so there are more collisions between species.
The catalyst lowers the activation energy, and so collisions between species are more effective.

13.



By adding a lot of NH_3 to give a visible blue.

Use a Pt electrode.



SOME EXPLANATIONS concerning the Part 1 and Part 2 answers

Part 1.

6. The C=O group is polar, and with only two methyl groups the propanone is miscible with water in all proportions; although the -OH group in 1-hexanol is highly polar, the hexyl group is non-polar and with 6 C atoms it is of sufficient size to prevent the 1-hexanol from dissolving to any significant extent in water.

20. All four statements are true, but only I and II follow from the experiment described.

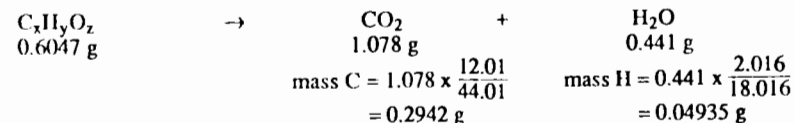
28. When any ore of a moderately reactive metal is roasted in air, the metal will end up as metal oxide.

Part 2.

1. (d) Hydrogen gas from a cylinder is colourless and odourless, but hydrogen gas when first produced by chemical reaction has quite a pungent stink—It is not known whether this is because it contains individual H atoms or excited H₂ molecules, but it certainly is much more reactive than cylinder hydrogen and reactive enough to affect the inside of the nose. When it is left to stand it loses its stink. Both 'pungent' and 'odourless' were accepted here, even though when you do this experiment in the laboratory you will certainly detect the odour of the hydrogen.
2. (b) For a bond to be non-polar, not only must the bond be between like atoms, but the environment of those like atoms must be the same; thus CH₃CH₂OH and CH₃CH₂Cl have polar C-C bonds. In CH₃(CH₂)₉OH, however, a C-C bond at the other end of the chain and remote from the OH group will be virtually non-polar.
- 12.(c) An equally good answer is that a catalyst provides an alternative reaction path. A complete answer would include both means by which a catalyst may speed a reaction.
- 13.(c) An electrolysis reaction is slow, and Cu²⁺ is not highly coloured. Hence it takes considerable time before either the blue of the copper ion or the eating away of the copper electrode becomes visible. In order to produce the deep blue (and hence more easily visible) Cu(NH₃)₄²⁺ you have to add a lot of NH₃—you have to add enough to react with all the H⁺ of the sulfuric acid before the complex ion will form.

PART 3

1. (a)



	C	H	O
mass (g)	0.2942	0.04935	0.6047 - (0.2942 + 0.04935) = 0.2612
+ At Wt	12.01	1.008	16.00
= (mol)	0.02450	0.04896	0.01632
+ smallest no (0.01632)	1.501 (= 1½)	3.000	1
x 2	3.002	6.000	2
Round	3	6	2



(b) 57 mL of 'A' at 373 K and 1.00 atm has a mass of 0.1367 g

$$24470 \text{ mL} \dots\dots\dots 0.1367 \times \frac{24470}{57} \text{ g}$$

$$\dots\dots\dots 298 \text{ K} \dots\dots\dots 0.1367 \times \frac{24470}{57} \times \frac{373}{298}$$

$$= 73 \text{ g}$$

ALTERNATIVELY

$$PV = nRT$$

$$\therefore n = \frac{PV}{RT} = \frac{1.00 \times 0.057}{0.08206 \times 373}$$

$$\text{Hence } \frac{1.00 \times 0.057}{0.08206 \times 373} \text{ mol} = 0.1367 \text{ g}$$

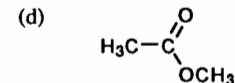
$$\therefore 1 \text{ mol} = 0.1367 + \frac{1.00 \times 0.057}{0.08206 \times 373} = 0.1367 \times \frac{0.08206 \times 373}{1.00 \times 0.057}$$

$$= 73 \text{ g}$$

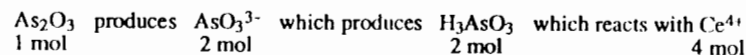
Note: the volume of gas is given in the question to 2 significant figures because in an experiment it would be impossible to measure any more accurately; thus in this instance 2 significant figures is the appropriate accuracy for the answer (although 3 significant figures was accepted).

(c) MWt of C₃H₆O₂ = (3 × 12.01 + 6 × 1.008 + 2 × 16) = 74.08 ≈ 73

∴ Molecular formula = empirical formula = C₃H₆O₂



2. (a) It follows from the equations that



$$\begin{array}{l} \text{(b)} \quad 0.2476 \text{ g As}_2\text{O}_3 \equiv 46.34 \text{ mL ? M Ce}^{4+} \\ \text{no of mol As}_2\text{O}_3 = \frac{0.2476}{197.84} \end{array} \quad \left| \quad \begin{array}{l} \text{no of mol Ce}^{4+} = \frac{0.2476}{197.84} \times 4 \end{array} \right.$$

$$46.34 \text{ mL ? M Ce}^{4+} \text{ contains } \frac{0.2476}{197.84} \times 4 \text{ mol Ce}^{4+}$$

$$\therefore 1000 \text{ mL ? M Ce}^{4+} \text{ contains } \frac{0.2476}{197.84} \times 4 \times \frac{1000}{46.34} \text{ mol Ce}^{4+}$$

$$= 0.1080$$

$$\therefore \text{Concentration of Ce}^{4+} = 0.1080 \text{ mol L}^{-1}$$

$$\therefore \text{Concentration of Ce(SO}_4)_2 = 0.1080 \text{ mol L}^{-1}$$

Note: all the factors that go to make up the answer of 0.1080 M are either exact values or accurate to 4 or more significant figures; thus 4 significant figures is the appropriate answer for the answer (although 3 significant figures was accepted).

3. (a) $20.00 \text{ mL } 0.04343 \text{ M OH}^- \equiv 20 \text{ mL ? M H}^+ \text{ (from HSO}_4^-)$

\therefore She wants 0.04343 M H^+

$$\therefore \text{In } 500.0 \text{ mL she wants } 0.04343 \times \frac{500.0}{1000} \text{ mol}$$

$$= 0.04343 \times \frac{500.0}{1000} \times 120.06 \text{ g NaHSO}_4$$

\therefore If the cleaner is 410 g/kg , she needs more, that is

$$= 0.04343 \times \frac{500.0}{1000} \times 120.06 \times \frac{1000}{410} \text{ g toilet cleaner}$$

$$= 6.359 \text{ g}$$

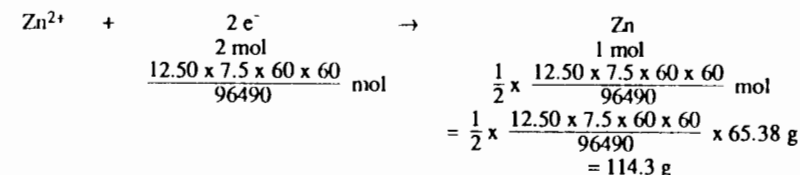
(b) The toilet cleaner has more NaHSO_4 than claimed.

$$\text{Hence its content is } 410 \times \frac{20.00}{18.00} = 456 \text{ g/kg}$$

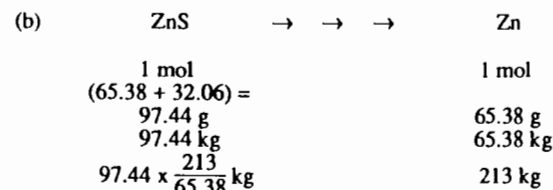
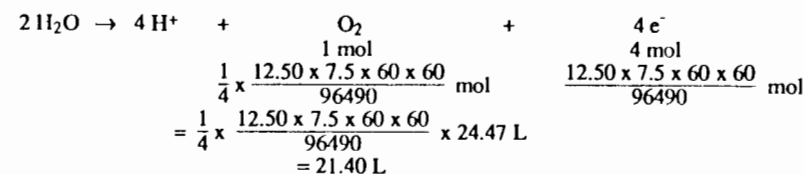
4. (a)

$$\text{(i) time (t) = } 7.5 \times 60 \times 60 \text{ seconds}$$

$$\begin{aligned} Q &= i \times t \\ &= 12.50 \times 7.5 \times 60 \times 60 \text{ C} \\ &= \frac{12.50 \times 7.5 \times 60 \times 60}{96490} \text{ F (= } 3.498 \text{ F or mol of e}^-) \end{aligned}$$



(ii)



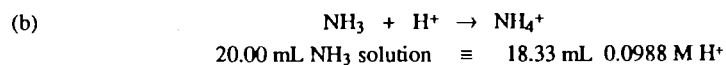
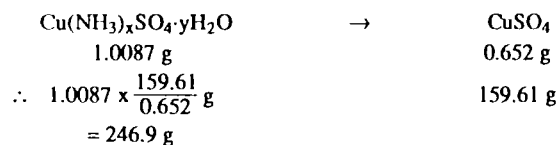
But because of the incomplete extraction the amount of ZnS in the ore must have been more, that is

$$97.44 \times \frac{213}{65.38} \times \frac{100}{89} \times \frac{100}{92} \text{ kg.}$$

$$\text{Hence the percentage of ZnS in the ore is } \frac{97.44 \times \frac{213}{65.38} \times \frac{100}{89} \times \frac{100}{92}}{503} \times 100 = 77 \%$$

Note: the 2 efficiencies are given in the question to 2 significant figures because they result from practical inadequacies in the experiment and it would be impossible to determine them with greater precision; thus in this instance 2 significant figures is the appropriate accuracy for the answer (although 3 significant figures was accepted).

5. (a) MWt of $\text{CuSO}_4 = 63.55 + 32.06 + 4 \times 16.00 = 159.61$)



	no of mol $\text{H}^+ = \frac{18.33}{1000} \times 0.0988$	
no of mol $\text{NH}_3 = \frac{18.33}{1000} \times 0.0988$		
20.00 mL NH_3 solution contains $\frac{18.33}{1000} \times 0.0988$		mol NH_3
250.00 mL NH_3 solution contains $\frac{18.33}{1000} \times 0.0988 \times \frac{250.0}{20.00}$		mol NH_3
	$= 0.02264 \text{ mol}$	

(c) Number of mol of complex in the 1.4009 g = $\frac{1.4009}{246.9}$
 Hence $x = \frac{0.02264}{\frac{1.4009}{246.9}} = 3.99 \approx 4$

(d) MWt of $\text{Cu}(\text{NH}_3)_4\text{SO}_4 = 159.61 + 4 \times (14.01 + 3 \times 1.008) = 227.75$

\therefore contribution to MWt of $y\text{H}_2\text{O} = 246.9 - 227.75 = 19.15$

[Note: because this is a small difference between two large numbers the percentage error is considerable]

$y = \frac{19.15}{18.02} = 1.06 \approx 1$

\therefore the formula of the complex is $\text{Cu}(\text{NH}_3)_4\text{SO}_4 \cdot \text{H}_2\text{O}$

PART 4)

This section of the paper is designed to give you the opportunity to demonstrate what you know rather than what you don't know. Hence there is no one model answer. Several students could write quite different essays and yet all could score full marks. Clear setting out and logical order are important. Clear and concise English expression is important. A brief introduction and conclusion ideally should be included.

In general, then, for full marks we are looking for two pages of

- * good chemistry,
- * on the topic given,
- * written in reasonable English,
- * with reasoning shown (for example, an inter-relating of evidence and theory),
- * with a beginning, a middle, and an end, and
- * either no errors or errors which are trivial.

Diagrams, graphs, drawings, schematic outlines, and so on, should be included if this is an appropriate method for getting your message across.

1. Although this topic appears 'open-ended', your essay almost certainly has to be based on a consideration of intermolecular (including inter-ionic) forces. Probably your essay should include the following points:
 - * substances dissolve if solute/solvent interactions > solute/solute interactions + solvent/solvent interactions
 - * van der Waals Forces, consisting of hydrogen-bonding (the strongest) and the two weaker forces dipole/dipole interaction the dispersion force (also called London dispersion force or London force)
 - * polar substances dissolving in each other
 - * non-polar substances dissolving in each other
 - * non-polar and polar substances not dissolving in each other
 [Note: like does *not* attract like; polar substances attract each other, but the immiscibility of non-polar and polar substances is a result of the fact that the non-polar molecules cannot overcome the stronger attractive forces between the polar molecules; it is not that there is some special and different attraction of non-polar molecules for each other.]
 - * ionic substances. Here the attractive forces between ions are very strong so that some are insoluble in water (where the ion/ion attractions are stronger than the ion/water attractions) while some are soluble (where the ion/water attractions are stronger). But because the ion/ion forces are so strong, all are insoluble in non-polar solvents.
 Other matters which you might mention include:
 - ‡ the strength of the inter-atomic forces in infinite covalent substances is so great that they are insoluble in all solvents
 - ‡ covalent substances with very large molecules (for example, cellulose or PVC) are insoluble in all solvents
 - ‡ water is very highly polar and is sometimes called the universal solvent
 - ‡ soaps and detergents help non-polar substances to 'dissolve' in polar water
 - ‡ substances 'dissolve' because of chemical reaction, for example, carbonates dissolve in H^+ and many metals in H^+
 - ‡ metal alloys are solid solutions
 - ‡ substances whose molecules contain both non-polar and polar groups (for example, 1-propanol is completely soluble in both water and cyclohexane)
 - ‡ the 'like dissolves like' rule. Note: this is *not* a 'like attracts like' rule
 The topic states 'Why substances dissolve'. A good essay must include reasoning. A statement in 2 pages of the solubility rules (for ionic compounds) is worth nothing. A statement in 2 pages that like attracts like, along with examples is worth only half marks if there is no argument in terms of inter-molecular forces.

2. In this essay you must write about both *proton-transfer* and *electron-transfer*, and you must give examples involving *predicting* reactions. There is however some flexibility of approach. The essay could be constructed around

- * a discussion of the principles of acid/base reactions followed by a discussion of the principles of redox reactions, or
- * the principles of acid/base reactions contrasted with the principles of redox reactions.

Probably you should make the following points:

- * acid/base reactions are proton-transfers
- * strong acids are powerful proton-donors while weak acids are less powerful proton-donors
- * strong bases are powerful proton-acceptors while weak bases are less powerful proton-acceptors
- * strong acids and strong bases react together (for example H^+ and OH^-)
- * weak acids and strong bases react together (for example CH_3COOH and OH^-)
- * strong acids and weak bases react together (for example H^+ and NH_3)
- * weak acids and weak bases may react together (for example CH_3COOH and NH_3)
- * weak acids and weak bases will not react together if one, other, or both are sufficiently weak (for example NH_4^+ and Cl^-)
- * redox reactions are electron-transfers
- * strong reducers (or strong reducing agents) are powerful electron-donors
- * strong oxidisers (or strong oxidising agents) are powerful electron-acceptors
- * strong reducers and strong oxidisers react together
- * whether weaker reducers and weaker oxidisers react together depends on their relative strengths and this can be deduced from an E° Table.

Possible other points include:

- ‡ conjugate acid/base pairs
- ‡ strong bases have weak conjugate acids and strong acids have weak conjugate bases
- ‡ for the reaction $acid_1 + base_2 \rightleftharpoons base_1 + acid_2$, the stronger the $acid_1$ and the stronger the $base_2$ the larger the value of K for the reaction
- ‡ some oxidisers can be reduced to differing extents (for example Fe^{3+} to Fe^{2+} or Fe)
- ‡ some reducers can be oxidised to differing extents (for example Fe to Fe^{2+} or Fe^{3+})

3. Although the brevity of this topic may imply an open-ended essay, you will probably know the standard properties of transition metals that are included in the syllabus.

The following points are probably necessary:

- * position of the transition elements in the Periodic Table
- * with one or two electrons in the outermost shell the transition elements are metals
- * incomplete d sub-shells in the transition metals
- * variable oxidation states for most transition metals in their compounds
- * colour in most compounds of transition metals
- * transition metals ions form complexes

[An outstanding essay would go further than the above glib syllabus statements (with examples), and point out the following:

- * at the beginning of the transition metals, with Sc and its sub-group (and to a lesser extent Ti and its sub-group), the metals are sufficiently close to inert-gas electron configurations for them to lose 3 (or 4) electrons and show an invariable oxidation state of III (or IV) with a colourless ion. The metals to the left have less non-metallic character and so are less prone to complex-ion formation.
- * at the end of the transition metals, with Zn and its sub-group (and to a lesser extent Cu and its sub-group), the d sub-shell has been completed and the

stability of the 18-set (or filled d sub-shell) leads again to an invariable oxidation state of II for Zn and Cd (or I for Ag) with a colourless ion. The metals to the right have more non-metallic character and so have similar tendency to complex-ion formation as the transition metals near them.

Possible other points include:

- ‡ how the d block comes about
- ‡ the similar energies of the 3d and 4s subshells (or nd and (n+1)s subshells)
- ‡ transition metals are generally not as strong reducers as the metals of the main groups
- ‡ the simple 1+, 2+, and 3+ ions are typical of metals, while covalent ions such as CrO_4^{2-} and MnO_4^- are typical of non-metals
- ‡ the general periodic table trend of more metallic in the lower left and more non-metallic in the upper right does not apply to the transition metals; an activity series of these metals can be read from the E° Table as Mn Zn Cr Fe Cd Co Ni Cu Hg Ag Pt Au, and this does not follow any identifiable trend.

ACTUAL STUDENT ESSAYS

The essays reproduced below were written by students in the 1996 TEE. All the essays gained full marks. Some (but not all) minor errors of spelling, punctuation, and grammar have been corrected—we recognise that under examination conditions students are likely to make trivial errors, and you are not penalised for this. With the elimination of slip-ups, the aim has been to produce what the student meant to write.

Essay 1 (Why substances dissolve)

Substances dissolve as a result of bond formation between solute and solvent particles. This process will only occur if the solute-solvent attractions are stronger than the solute-solute and solvent-solvent attractions. Thus, similar intermolecular forces between solutes and solvents will result in a solution and hence the phrase "Like dissolves Like".

This concept can be more easily understood if one examines the various intramolecular and intermolecular forces existing in structures. Intramolecular forces can be grouped as metallic, ionic, covalent network and covalent molecular.

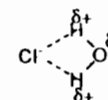
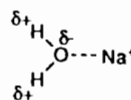
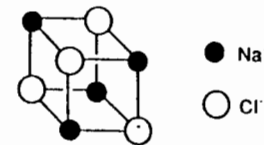
Metallic bonding

Metallic bonding occurs as a result of the electrostatic attractions between delocalised electrons and positive metal nuclei. Thus metals will not dissolve in water or any other common solvents as the bonding in these substances are not metallic. Instead, metals will dissolve in other metals and consequently alloys such as brass are possible.

Ionic bonding

Ionic bonds are formed between oppositely charged ions. For example, NaCl exists as a crystalline structure as a result of the electrostatic attractions between the positive Na^+ and negative Cl^- .

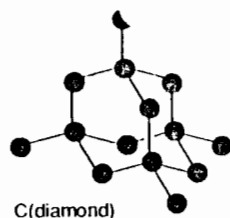
Most ionic substances will dissolve in polar substances such as water because the negative end of the water molecule will "attack" the positive sodium ion. Similarly, attractive forces occur between the positive hydrogen end of the water molecule and the negative chloride ion.



However with some ionic substances such as silver chloride the attractions between the ions are so strong that the substance is insoluble. Solubilities of ionic substances are summarised in tables of Solubility Rules.

Covalent Network

Covalent bonds form when two elements of similar electronegativity share electrons. A covalent network substance such as diamond or silicon dioxide consists of atoms individually bonded covalently to each of its neighbours. This results in very strong bonding occurring throughout the structure and so covalent network substances do not dissolve.

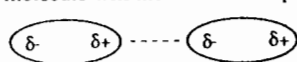


Covalent Molecular

These structures consist of groups of covalent molecules bonded together with weak intermolecular or van der Waals forces.

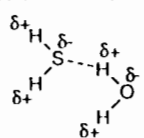
A further classification can be made as to the types of intermolecular forces which exist in covalent molecular substances. These forces are: dispersion forces, dipole-dipole attractions and hydrogen bonding.

Firstly, dispersion forces. At any moment in time, the electron cloud surrounding a molecule, may not be symmetrically distributed. This will result in a slight dipole and this charged molecule will then induce a dipole in its neighbours.



This attraction is known as a dispersion force and occurs in non polar molecules such as I_2 , C_6H_{14} and so on. Consequently, such molecules will not dissolve in polar

substances but rather in non-polar organic solvents such as kerosene. These solvents are also held together by dispersion forces and so the solute-solvent attractions will overcome the solute-solute/solvent-solvent bonds.

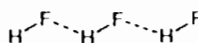


Secondly, dipole-dipole attractions. These occur when polar molecules such as H_2S and HBr bond. Because individual atoms have a small charge they will form dipole-dipole bonds.

Therefore polar molecules will dissolve in polar solvents such as H_2O due to the presence of similar intermolecular forces.

Lastly, hydrogen bonding. Hydrogen bonds form between molecules which have hydrogen covalently bonded to either F, O or N. Since hydrogen has a very low electronegativity and F, O and N have high electronegativity the bond formed is very polar. Hydrogen bonds can be seen as an extreme form of dipole-dipole attraction. In any case, bonds form between the positively charged hydrogen end and non-bonding electron pairs in the F, O or N atom of another molecule.

Examples of such molecules are HF , NH_3 and H_2O . In light of this, it is not surprising that covalent-molecular substances in which hydrogen bonding exists are highly soluble in substances such as water and liquid ammonia.



Thus the answer to the question 'why do substances dissolve?' is 'because of similar intermolecular forces'. A solute will dissolve in a solvent that has similar bonding, as the solute-solvent forces overcome the respective solute and solvent forces. Conversely, solutes will not dissolve in solvents that have dissimilar bonding.

Essay 1 (Why substances dissolve)

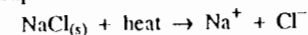
Dissolving

When a solid is added to a liquid, it may interact chemically with the liquid to change the state of the solid, and a *solution* may be formed. The liquid is referred to as a *solvent*, and the solid as a *solute*. Liquids may also dissolve in liquids, a concept referred to as miscibility.

Interaction of solute and solvent

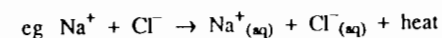
When a solute is added to a solvent, three main interactions must take place in order to form a solution:

- (1) The bonds in the solute must be overcome. This is an endothermic process, as it requires energy to overcome the chemical bonds. For example, with an ionic solid, the ionic lattice must be disrupted in order to allow the ions to dissociate



- (2) The forces between the molecules of the solvent must be overcome. For example, in an aqueous solution the hydrogen bonds and other van der Waal's forces must be overcome between the water molecules. This process is also endothermic, requiring energy to disrupt these bonds.

- (3) The solute and solvent molecules interact to form a solution.



This process is exothermic, as heat is given off when the interaction occurs.

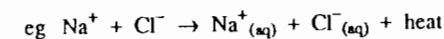
Bonding of Solute and Solvent

Examining the bonding of solute and solvent is critical to understanding why solutions form. Bonding in the solute and solvent determine the magnitude of the energy changes involved. For example, overcoming the very polar ionic lattice takes up a large amount of energy



Overcoming the forces between polar molecules is also relatively difficult, and thus ΔH for such a reaction is also large and positive. Overcoming relatively weak dispersion forces in non-polar solutes requires little energy, and thus ΔH is small and positive.

The interaction between a polar solute and a polar solvent is large, and thus ΔH for step 3 is large and negative in this case



However, the interaction between either a polar solvent and non-polar solute, a polar solute and a non-polar solvent, or a non-polar solvent and solute are small, and therefore ΔH is small and negative.

Adding the three ΔH 's determines overall energy change. If this is small, dissolution is likely to occur. If it is large and positive, then dissolution is unlikely.

eg Case 1 - polar solvent, polar solute

$$\Delta H_1 = \text{large +ve}$$

$$\Delta H_2 = \text{large +ve}$$

$$\Delta H_3 = \text{large -ve}$$

$\therefore \Delta H \text{ solution} = \text{small}$, and thus the solute is likely to dissolve

Case 2 - polar solvent, non-polar solute

$$\Delta H_1 = \text{small +ve}$$

$$\Delta H_2 = \text{large +ve}$$

$$\Delta H_3 = \text{small -ve}$$

$\therefore \Delta H \text{ solution} = \text{large and +ve}$, and thus the solute is not likely to dissolve.

Essay 2 (Proton-transfer and electron transfer)

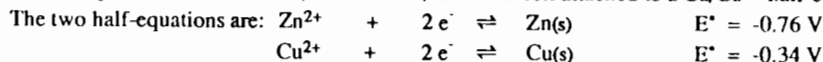
Chemists use concepts such as electron-transfer and proton-transfer to determine whether or not a reaction will occur spontaneously.

Electrochemistry is the study of electron transfer in "redox" reactions. In such a reaction, one substance must donate an electron to another substance. The electron donating substance is termed the reducing agent or reductant, while the chemical accepting the electron is the oxidising agent or oxidant. If an electrochemical cell is set up between two substances

(that is, the oxidation half-equation is separated from the reduction half-equation so that the electron transfer must occur through an external conducting wire), then the tendency of a current of electrons to flow through the external wire will produce a voltage.

In order to predict whether a redox reaction will take place, we must first determine what the voltage of that particular cell will be. A table of standard reduction potentials has been established to illustrate the likelihood of a substance being reduced by the standard half-cell (H^+/H_2) under standard conditions. If a half-equation has a positive standard reduction potential (E°) then it is likely that the reaction will take place when connected to a (H^+/H_2) half-cell. To determine what reactions will occur, the possible reactions are written down. The half-equation with the most positive E° is most likely to happen and therefore, if the reaction is to take place, the reverse of the other reaction will occur. This reaction is therefore "flipped" and the sign of its E° is changed. The two E° 's are now added and if the final E° is greater than 0.0 volt, the reaction is likely to occur spontaneously.

For example, in the Daniel Cell, there is a Zn/Zn^{2+} half-cell attached to a Cu/Cu^{2+} half-cell.



It is likely that $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu(s)}$ will occur as its E° is greater than 0 V. Therefore, the half-equation in the Zn/Zn^{2+} couple must be: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2e^-$ $E^\circ = +0.76\text{V}$.

The combined E° of these half-cells is 1.10 V. As this is greater than 0 V it is likely that the reaction will occur.

This concept can also be used in electrolytic cells in which a current is used to produce a chemical change. All of the possible half-equations are written and the oxidation half-equations with the most positive oxidation potentials, and the reduction half-equations with the most positive reduction potentials are the reactions most likely to occur if sufficient voltage is produced.

The transfer of H^+ is another helpful concept. This involves the understanding of whether a substance is most likely to act as an acid or a base, and hence is used in reactions involving acids and bases.

Strong acids such as HCl , H_2SO_4 and HNO_3 , completely ionise in water to produce H^+ . Therefore, if it is to react with another substance, the reaction must involve the acid donating a proton. Weak acids, however are less easily ionised. HSO_4^- only slightly ionises to form SO_4^{2-} and H^+ and therefore it is possible that a reaction could take place in which HSO_4^- didn't produce H^+ . Similarly, as SO_4^{2-} is not so readily produced, it is a stronger base than HSO_4^- and can, at times, react to form HSO_4^- , for example, $\text{SO}_4^{2-} + \text{H}_2\text{O} \rightleftharpoons \text{HSO}_4^- + \text{OH}^-$. It is because of this reaction, which only occurs to a small extent, that a solution of sodium sulfate is slightly basic.

Weak bases such as NH_3 don't readily ionise - $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$, occurs only to a small extent, and hence NH_3 will react with fewer acids than the strong base OH^- , and when it does react it reacts to a smaller extent.

Essay 2 (Proton-transfer and electron transfer)

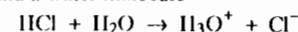
Chemists will always want to find out whether the reactions that they are going to do will occur or not. The reactions can therefore take place if the forward reaction will favour the products and the reverse reaction will not favour the reactants to any significant extent.

The two reactions, forward and reverse, are the ones which determine whether the reaction will go to completion or not. If the reaction goes to completion or the amounts of products are much larger than the amounts of reactants, then the reaction is said to have taken place.

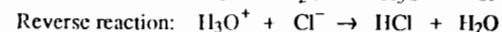
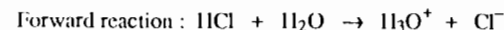
The two reactions are considered to compete against each other. It depends on the strength (relative to each other) of both directions of reaction.

Consider the acid-base reaction. According to Bronsted-Lowry theory, an acid is a substance that donates a proton and a base is a substance that accepts a proton. A base which accepts a proton becomes a potential proton-donor and hence an acid (the conjugate acid of the original base) and an acid which donates a proton becomes a potential proton-acceptor and hence a base (the conjugate base of the original acid). For example there can be a reaction between a

hydrogen chloride molecule and a water molecule

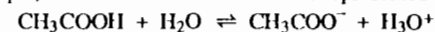


If the reaction takes place then the hydrogen chloride molecule will ionise to give an acidic solution due to the oxonium ion. If a chemist wants to analyse the extent of the reaction, we must consider two reactions



There is a competition between acids to donate protons and between bases to accept protons. That which donates its proton is the stronger acid, and that which accepts the proton is the stronger base. Now we must identify the two acids and the two bases: the acids are HCl and H_3O^+ , and the bases are Cl^- and H_2O . If we compare the strength of the acids or the strength of the bases we will come up with the same result. Considering the acids, HCl is a stronger acid than H_3O^+ and therefore HCl will donate its proton more readily and therefore the reaction favours the forward direction. Considering the bases, H_2O is a stronger base than Cl^- and therefore H_2O will accept the proton more readily and so the reaction favours the forward direction. Note that the stronger acid HCl has the weaker conjugate base Cl^- , and the weaker acid H_3O^+ has the stronger conjugate base H_2O .

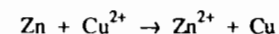
Consider another example, the ionisation of ethanoic acid represented by



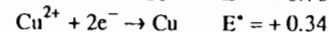
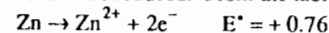
The two acids here are CH_3COOH and H_3O^+ . From chemical data H_3O^+ is a stronger acid than CH_3COOH and therefore the reverse reaction is favoured and the reaction does not occur to a large extent. I have used the double arrow to show that an equilibrium is set up.

In Redox reactions the same principles apply, but in this case electrons are transferred instead of protons. Something that is oxidised releases electrons and something that is reduced accepts electrons. It depends on the readiness to accept or gain electrons that a reaction will take place.

An example is a zinc strip dropped into an aqueous solution of copper(II) sulfate. The reaction is



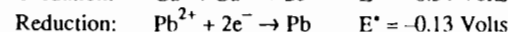
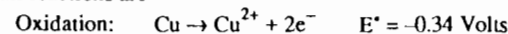
To compare the readiness of two species, Zn to be oxidised, and Cu^{2+} to be reduced, a chemist will use a table of standard reduction potentials. From the Table, Zn^{2+} can be reduced to Zn and the negative value (-0.76) indicates the reluctance of Zn^{2+} to be reduced; Zn can be oxidised to Zn^{2+} and the positive value (+0.76) indicates the readiness of Zn to be oxidised. The more positive the number (the *oxidation* potential) the more ready the substance is to be oxidised, and the more positive the reduction potential the more ready the substance is to be reduced. From the table it follows that:



Now, if we add the potentials (oxidation + reduction) we must get a positive number (the cell voltage) to indicate that the reaction occurs. Otherwise, if we get a negative number this indicates that the reaction does not occur. In this case the cell voltage is +1.12 volts and that means the reaction takes place.

Another example is the reaction of copper metal with solution of lead(II) nitrate. The possible reaction would be $\text{Cu} + \text{Pb}^{2+} \rightarrow \text{Cu}^{2+} + \text{Pb}$

The half reactions are



The total potential (cell voltage) is -0.47 volts which is negative and shows that the reaction cannot occur.

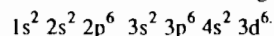
One significant difference between acid-base and redox reactions is: with acid-base reactions, an equilibrium is often set up where reactants and products are present together (although in varying amounts); while with redox reactions, the reaction often goes close to completion or goes not at all.

Essay 3 (Transition metals)

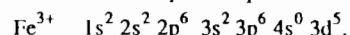
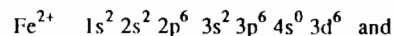
The transition metals are found in the middle section of the periodic table which is also known as the 'd block'. This is due to the fact that these metals have valence electrons in d orbitals within their atoms. Because of this, these 'd block' elements have both similar physical properties and chemical properties.

The transition metals are all hard dense metals with relatively high melting and boiling points. They all conduct electricity. They are also malleable and ductile. These physical properties can be attributed to the metallic bonding these elements exhibit - the lattice of cations surrounded by a sea of delocalised electrons. The melting and boiling points are high due to the strong electrostatic forces between the fixed cations and delocalised electrons. The non-directional nature of the bonds mean that the metals can be hammered into sheets or drawn into wires without disrupting the bonding. Thus they are malleable and ductile. The mobile charged particles, that is the delocalised electrons, can conduct heat and electricity well. Therefore they are all good conductors.

Indeed, not only do these metals have similar physical properties but also chemical properties. One of these is their ability to exist in variable oxidation states. Taking the transition metals of the fourth period of the periodic table, each element in this group has electrons in both the 4s and 3d sublevels. An example is iron with electron configuration

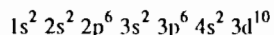


The 4s and 3d sublevels have very similar energies. Thus, these metals can lose electrons from either sublevel. Iron forms the ions



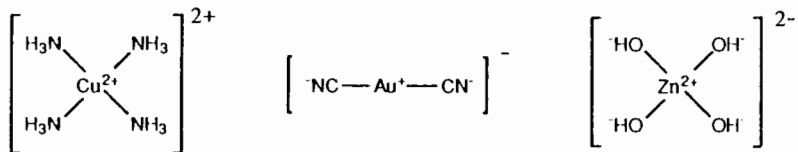
Fe^{3+} is more stable than Fe^{2+} as it has a half-filled 3d subshell. Another example of the tendency of transition metals to form various oxidation states is the existence of both the Cu^+ and Cu^{2+} ions.

Cu^{2+} salts are brightly coloured as are many other salts of transition metals. This property arises from the unfilled 3d orbitals in the metal atoms. The 3d orbitals have very similar energies. Thus an electron need only absorb a small amount of energy to 'jump' to another 3d orbital. This absorbed energy corresponds to that of visible light. Some colours from white light are absorbed and the salt appears coloured. Evidence of the fact that colours are due to incompletely filled 3d orbitals comes from zinc:



As shown, zinc has a completely filled 3d sublevel which is unaffected in the formation of the Zn^{2+} ion, and therefore its salts are not coloured but white. However, chromium which does not have a full 3d sublevel, like most transition elements forms coloured salts—in this case deep green, iron forms pale green and brown salts, and copper forms blue salts.

However, perhaps one of the most striking coloured chemicals is the bright blue $[\text{Cu}(\text{NH}_3)_4]^{2+}$. This is an example of a transition element's ability to form complex ions. Once again this is due to the incompletely filled 3d orbitals. Complex ions consist of a central metal cation surrounded by ions or molecules called ligands. Some examples of complex ions are illustrated below:



The number of ligands linked by coordinate covalent bonds to the central metal ion is called the coordination number. For the above complex ions the coordination numbers are 4, 2 and 4 respectively.

The transition metals are not as reactive as the Group I and Group II metals and thus have wide applications in society due to their electrical conductivity, high tensile strengths and

other properties. Examples of these applications are copper used in electrical wiring, iron used for building, and gold used for jewelry and as an international monetary standard. The transition metals' unique properties can be attributed to their atomic structure.

Essay 3 (Transition metals)

Transition elements are located in the d-block of the periodic table. They arise from the filling of the third, fourth and fifth principle energy levels. They are known as transition metals as they constitute a transition in the periodic table between the traditional metals of Groups I and II, and the less metallic elements of Groups III and IV. They possess similar properties to traditional metals, such as sodium, magnesium, and aluminium, but they also exhibit properties that are distinctly different from those metals.

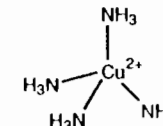
Like most metals, transition metals are soft, ductile and malleable. They also are good conductors of electricity and heat. They have a similar metallic lattice to traditional metals, and have delocalised valence electrons. This accounts for their ability to conduct. The bonding is said to be non-directional, allowing one layer to slide freely over another without disturbing the metallic bonds. This allows them to be malleable and ductile.

However, a unique property of transition metals is their ability to form a number of oxidation states. For example, iron can form Fe^{2+} or Fe^{3+} . This is due to the fact that the metals of this period can lose electrons from the 4s and 3d sub levels. This occurs because these two sub levels have similar radii and energy levels. So, when iron forms the Fe^{2+} ion it loses two 4s electrons, whereas when it forms the Fe^{3+} ion it loses two 4s electrons and one 3d electron from its original configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$.

Another property of transition metals is that their ions tend to be coloured and they form coloured compounds. This is attributed to the fact that the difference between the d orbitals is very small and only a small amount of energy is required to move an electron to a d orbital with a slightly higher energy. Therefore, when viewed in white light, the compound of the transition metal, tends to absorb certain frequencies of light resulting in it appearing a certain colour. Examples of this include: Ni^{2+} - green; Fe^{2+} - pale green; Fe^{3+} - red or orange; Cu^{2+} - blue.

Finally, transition metals are able to form complex ions. When a transition metal ion is mixed with a ligand—a negatively charged ion with a non-bonding pair of electrons (such as OH^- or CN^-), or a molecule with a non-bonding pair of electrons (such as H_2O or NH_3)—they form complex ions with the electron pairs of the ligands sharing their electrons with the transition metal ion. The number of ligands in a complex ion (or co-ordination complex) is either 2, 4 or 6, however, it is most commonly 6. Examples include $[\text{Au}(\text{CN})_2]^-$ and $[\text{Cu}(\text{NH}_3)_4]^{2+}$.

The diagram shows the tetrahedral shape of $[\text{Cu}(\text{NH}_3)_4]^{2+}$.



So, although they exhibit fairly similar properties to traditional metals, transition metals also have many unique properties that arise as a result of the filling of the d orbitals in the third, fourth and fifth principal energy levels.