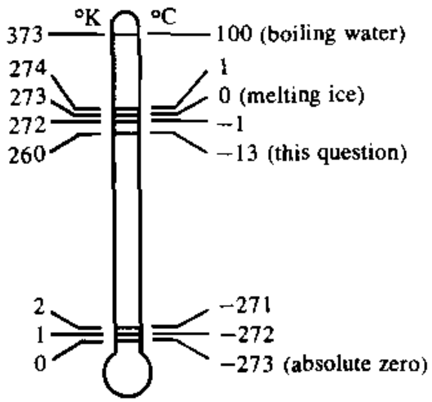
ANSWERS EXPLAINED
PART ONE

1. 1

$$T\text{ (in K)} = t\text{ (in °C)} + 273$$
$$K = °C + 273$$
$$K = (-13°C) + 273$$
$$K = 260$$

The connection between the K and C scales may be seen in this diagram of a thermometer. Notice that, while the two scales place 0 at different points, both scales use the same size steps, or intervals, between temperatures.



2. **4** Only compounds can be decomposed: either into simpler compounds or into elements. Elements cannot be decomposed into any simpler material by chemical means. Copper is the only element listed.

Mercury (II) oxide can be decomposed into mercury and oxygen.
Potassium chlorate can be decomposed into potassium chloride and oxygen.
Water can be decomposed into hydrogen and oxygen.
Copper cannot be decomposed.

3. **2** If the pressure of a gas is kept constant, then its volume changes with absolute temperature according to Charles' Law:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$$\frac{30\text{ ml}}{273\text{ K}} = \frac{V_2}{364\text{ K}}$$

$$(273)(V_2) = (30\text{ ml})(364)$$

$$V_2 = \frac{(30\text{ ml})(364)}{(273)} = \frac{(30\text{ ml})(91)(4)}{(91)(3)} = \frac{(30\text{ ml})(4)}{(3)}$$

$$V_2 = (30\text{ ml})(1.333\ldots) = 40\text{ ml}$$

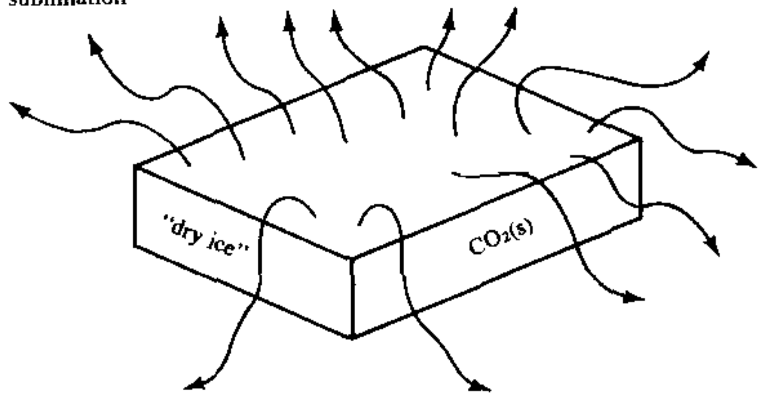
This answer is consistent also with the fact that as temperature increases the volume of a gas also increases. Choices (3) and (4) are inconsistent with this fact.

4. **1** To sublime means to change from a solid to a gaseous state directly, without first melting into a liquid.

solid state → gas state (sublimation)
(s) → (g)

Of the choices given, only CO_2 is given as being solid; the others are therefore eliminated. Indeed, CO_2(s) does sublime, as anyone who has seen the "dry ice" used by ice cream vendors can attest. It evaporates without melting.

CO_2(g) cloud
observed forming
by sublimation

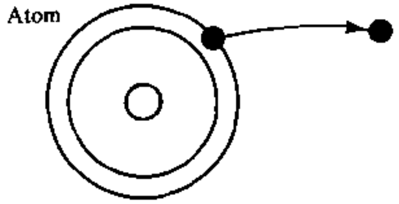


5. **2** A beta particle is actually an electron that emerges at very high speeds from the radioactive decay of atoms. The symbol β^- ("beta") is used for such an electron. Since it is an electron, it has a negative electric charge.
Reference Table H can be used to analyze this question.

Particle	Symbol	Notation Showing Mass and Charge	Charge
alpha	α	${}^4_2He$	+2
beta	β^-	${}^{0}_{-1}e$	−1
proton	p	${}^1_1H$	+1
neutron	n	${}^1_0n$	0

6. **3** When an electron is removed from an atom, that atom becomes an ion of +1 charge. The energy required to remove that electron is therefore called the *ionization energy*. Values for ionization energies (in kcal/mol) are given in Reference Table I.

Electron removed by ionization energy



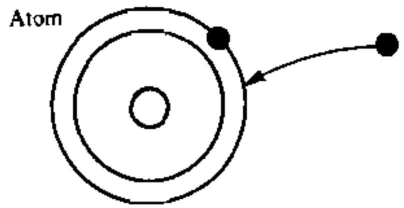
Wrong Choices Explained:

(1) *Kinetic energy* derives from the motion of atoms and molecules in a substance and is reflected in the temperature of the substance.

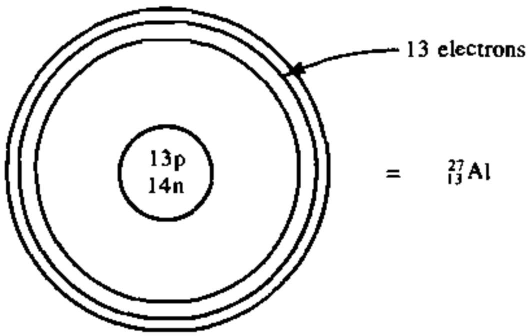
(2) *Potential energy* measures the heat or work that can be liberated in a chemical reaction or in some change of position in an electric or gravitational field.

(4) *Electron affinity* (sometimes called "electronegativity") relates to how well an atom can *take in* an additional electron to become a −1 ion. This concept is the reverse of ionization.

Extra electron added according to electron affinity value of atom



7. **1** The atomic number tells how many protons are in the atom's nucleus. Therefore, an atomic number of 13 implies the presence of 13 protons. Since in a neutral atom the number of electrons must equal the number of protons, 13 electrons must also be present.



$= {}^{27}_{13}\text{Al}$

8. **4** The ground states for the four atoms are diagramed below. Notice that only Na has *one* partially occupied orbital. Si has *two* partially filled orbitals, while Ne and Ca have their orbitals completely filled.

Orbitals

	1s	2s	2p x	2p y	2p z	3s	3p x	3p y	3p z	4s
Si (14 electrons)	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑		
Ne (10 electrons)	↑↓	↑↓	↑↓	↑↓	↑↓					
Ca (20 electrons)	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓
Na (11 electrons)	↑↓	↑↓	↑↓	↑↓	↑↓	↑				

The partially filled orbital (Na 3s)

9. **3** The chart below shows the sublevels possible in each principal energy level.

Principal Energy Level	Sublevels	Number of Sublevels
1	*s*	1
2	*s, p*	2
3	*s, p, d*	3
4	*s, p, d, f*	4

The third level has 3 possible sublevels within it.

This can also be seen if the sublevels are written out in ascending order of energy—a sequence chemistry students should memorize!

$1s^2 \quad 2s^2 \quad 2p^6 \quad 3s^2 \quad 3p^6 \quad 4s^2 \quad 3d^{10} \quad 4p^6 \quad 5s^2 \quad 4d^{10} \ldots$

($3s^2$, $3p^6$, $3d^{10}$ →) 3 sublevels (*s*, *p*, and *d*) in energy level 3

10. **3** Isotopes of an element have the same atomic number but different mass numbers.

Wrong Choices Explained:

(1) ${}^{158}_{64}X$ and ${}^{158}_{64}X$ are identical, not isotopes.

(2) ${}^{64}_{158}X$ and ${}^{158}_{64}X$ have different atomic numbers, so they can't be isotopes of the same element. (Incidentally, ${}^{64}_{158}X$ cannot exist. Can you see why?)

(4) ${}^{158}_{64}X$ and ${}^{158}_{65}X$ have different atomic numbers, so they are different elements. They have the same mass number coincidentally, but this does not qualify them to be isotopes.

11. **1** The valence electrons are those in the *outermost* (highest) principal energy level.

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ — These are the valence electrons, all in the outer third level.

12. **2** The compound's name, nitrogen(IV) oxide, tells you that nitrogen is in the +4 oxidation state. Oxygen always has a −2 oxidation state in oxides. By the criss-cross method you obtain the compound's formula:

a. Assign oxidation numbers — $\overset{+4}{N}\ \overset{-2}{O}$

b. Criss-cross oxidation numbers — $\overset{+4}{N} \times \overset{-2}{O}$

c. Possible formula — N_2O_4

d. Reduce to smallest whole number subscripts — NO_2

Wrong Choices Explained:

Choice	Charge on Nitrogen Atom	Name
(1)	+2	NO, nitrogen(II) oxide
(3)	+6	NO_3, nitrogen(VI) oxide
(4)	+8	NO_4, nitrogen(VIII) oxide

13. **4** An empirical formula has subscripts that cannot be reduced to any smaller whole numbers. K_2O meets this criterion. The other choices can be reduced.

Wrong Choices Explained:

Choice	Molecular Formula	Empirical Formula
(1)	C_2H_2	CH
(2)	C_2H_4	CH_2
(3)	Al_2Cl_6	$AlCl_3$

14. **3** This equation is best balanced by trial and error. Begin with the fact that, since there are *two* C atoms on the left, *two* CO_2 molecules should be produced. Next, balance the H atoms by placing a "2" in front of H_2O. The result:

$$C_2H_4 + 3O_2 \longrightarrow 2CO_2 + 2H_2O$$

15. **4** Molecules are polar if they meet two conditions:

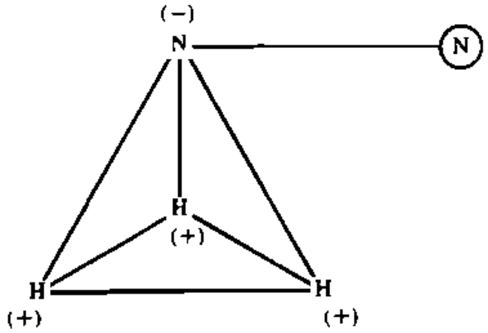
a. the bonds *within* the molecule are polar, *and*

b. the molecule has an asymmetric shape.

These conditions create an asymmetric charge distribution. NH_3, or ammonia, meets these criteria because

a. the N—H bond is polar, since N is more electronegative than H, *and*

b. the molecule is not symmetric, being a triangular-based pyramid with the N atom above the three H atoms.



Wrong Choices Explained:

(1) Being a gas is not directly relevant to whether a molecule is polar or not. Many gases are; many are not.

(2) N—H bonds are polar.

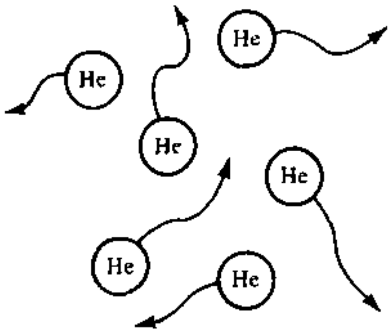
(3) N is a nonmetal; H, a metalloid. In any case, identifying the type of element is irrelevant to identifying polar molecules.

16. **1** The degree of ionic character in a bond can be determined from the difference between the electronegativity values of the elements in the bond. These are given in Reference Table I. The greater the difference in these electronegativity values, the more ionic is the bond formed.

Choice	Element	Electronegativities		Difference	Ionic Character
(1)	oxygen	O = 3.5	S = 2.5	1.0	greatest
(2)	chlorine	Cl = 3.0	S = 2.5	0.5	
(3)	bromine	Br = 2.8	S = 2.5	0.3	least
(4)	phosphorus	P = 2.1	S = 2.5	0.4	

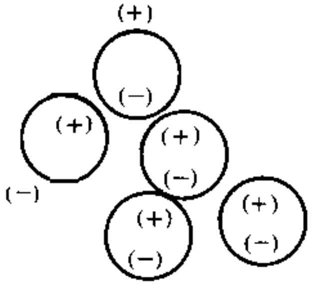
17. **3** Helium atoms have full outer electron shells. They do not normally give or receive electrons from other atoms or from each other. Therefore, helium is inert and remains as single (monoatomic) atoms in the gas state. Since these atoms are not attracted to one another, there exists no bond to make them coalesce to form a liquid or solid.

single He(g) atoms with no mutual attraction forces



However, the electrons around each atom do revolve, thereby creating *weak, temporary charge fluctuations* which can serve as the "glue" to form a liquid state under low temperatures. These relatively weak attractive forces are known as van der Waals forces.

He(l)—atoms held together by weak, partial, fluctuating van der Waals charges, (+) and (−)



Wrong Choices Explained:

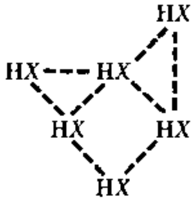
(1) Hydrogen bonding results from highly polar bonds that contain H as one atom. Here we are dealing with the pure element helium, not hydrogen.

(2) Covalent bonds are formed by the sharing of electrons between two atoms, as in the diatomic molecules H_2, Cl_2, and O_2. As discussed above, helium atoms are inert and monoatomic, and do not do this.

(4) Ionic bonding begins when electrons are transferred to make (+) and (−) ions. This does not occur with helium.

18. **2** Hydrogen bonds are formed between molecules that contain some hydrogen. If the bond between the hydrogen atom and its partner is very polar, then the attractive forces between neighboring molecules will be strong.

δ(+) δ(−) Partial charges
H—*X*
Polarized bond
within molecule



Hydrogen bonds
between
molecules

The HX molecule above will have the greatest polarization *within* itself and *between* its neighbors when element *X* has an electronegativity high enough to pull electrons closer to itself than to H.

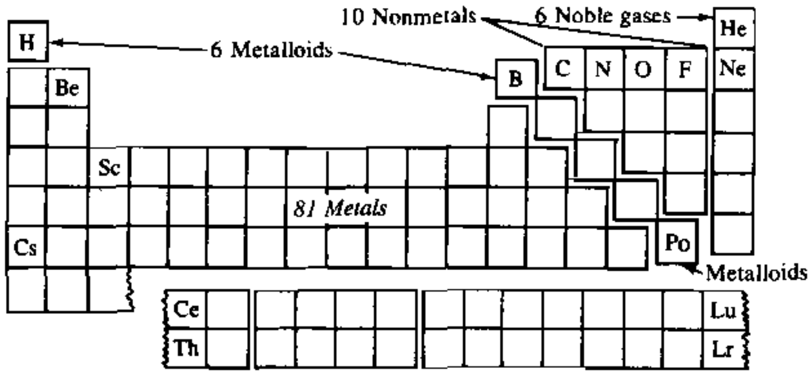
δ(+) H :Xδ(−)

Atoms with high electronegativities also have small radii. Note, for example, that fluorine (F) has the highest electronegativity (4.0, from Reference Table I) and a small radius (0.72 Å, from the Periodic Table). In fact, HF does have extremely strong hydrogen bonds between molecules.

19. **3** Elements within the same family (or group) behave in similar ways because they have the same electron configuration in their outermost energy levels. Mg and Ca are both metals in Group IIA on the Periodic Table. Notice that they have similar outer s^2 orbitals.

The other pairs of elements are not from the same groups and so will not be similar in chemical properties.

20. **2**



21. **3** The halogen family is Group VIIA on the Periodic Table.

22. **1** Reference Table I gives you the electronegativities for these elements:

C	2.5	(lowest)
F	4.0	
N	3.0	
O	3.5	

23. **2** Nonmetals are poor conductors of electricity because they do not easily release their outer electrons as metals do. In other words, nonmetals require a lot of energy to release even a single electron—they have high ionization energies.

This is confirmed by Reference Table I. Note that the nonmetals have ionization energies from 200 to 400 kcal/mol. Compare this to the metals, which have ionization energies primarily in the 90 to 200 kcal/mol range.

24. **4** The Periodic Table gives electron structures. Looking down each group, you can detect the similarity in outer electrons among group members.

Choice	Group	Typical Outer Electron Pattern	Total Outermost Electrons
(1)	IA	s^1	1
(2)	O	s^2	2
(3)	VIA	s^2p^4	6
(4)	IVA	s^2p^2	4

25. **4** To say that an element has a crystalline structure is to imply that it's a solid. Under normal conditions, such as STP, iodine is a solid, bromine is a liquid, and fluorine and chlorine are gases.

26. **2**

a. Fe_2O_3

Fe mass: 2 × 56 (at. wt.) = 112 amu
O mass: 3 × 16 = 48 amu
formula mass = 160 amu

b.

$$\%\text{ oxygen, by mass} = \frac{\text{oxygen mass}}{\text{formula mass}} \times 100$$

$$\%\text{ oxygen} = \frac{48\text{ amu}}{160\text{ amu}} \times 100$$

$$= \frac{(3)(16)}{(10)(16)} \times 100$$

$$= \frac{3}{10} \times 100 = 30\%$$

27. **4** *Obtain the formula mass of LiF.*

Li = 7 amu
F = 19 amu
LiF = 26 amu

Therefore, 1 mole of LiF will have a mass of 26 grams.

Now solve for the number of moles actually given.

$$\text{number of moles} = \frac{\text{mass given}}{\text{mass of 1 mole}} = \frac{39\text{ g}}{26\text{ g/mol}} = 1.5\text{ moles LiF}$$

28. **2**

a. Calculate the number of moles in the solution taken before dilution.

12-molar solution = 12 moles of solute/each liter of solution

$$\text{number of moles taken before dilution} = 0.50\text{ liter} \times \frac{12\text{ moles solute}}{1\text{ liter}} = 6\text{ moles}$$

b. Calculate new molarity after dilution.

$$\text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution made}} = \frac{6\text{ moles}}{1\text{ liter}} = 6\text{ molar}$$

29. **4** This is a "volume-volume" gas problem. One result of Avogadro's Law is that the volumes of gases in a gas reaction (at constant pressure) are in the same ratio as the coefficients in the balanced equation.

a. Balanced equation: $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

b. Mole ratios:	4	5	4	6	← Identical
c. Gas volume ratios:	4	5	4	6	←
d. Gas volumes given in problem:		X	40. liters		

e. Set up ratios:

$$\frac{5}{X} = \frac{4}{40.\text{ liters}}$$

f. Cross-multiply:

$4(X) = 200$ liters
$X = 50.$ liters of $O_2(g)$ required

30. **1** Avogadro's Law states that 1 mole of *any* gas at STP will occupy 22.4 liters. This is the basis for our solution.

a. Calculate the moles of He in the sample.

atomic weight of He = 4 amu
so 1 mole of He = 4 grams

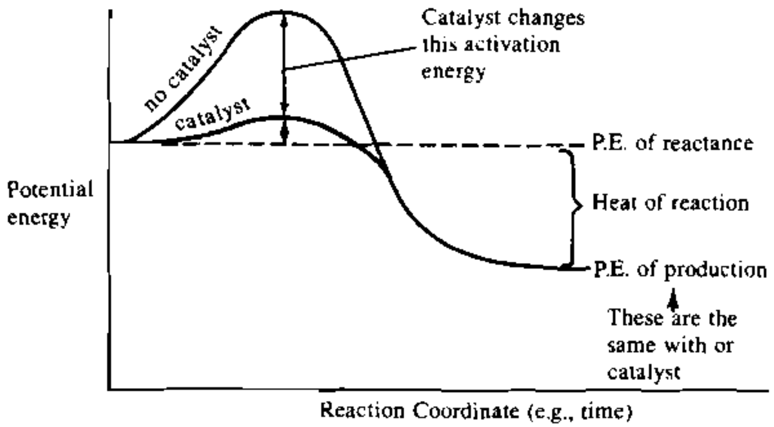
$$\text{number of moles} = \frac{\text{mass given}}{\text{mass of 1 mole}} = \frac{2.00\text{ g}}{4.00\text{ g/mol}} = 0.500\text{ mole He}$$

b. Use Avogadro's Law.

volume of gas = number of moles (at STP) × 22.4 liters/mol
volume = 0.500 mole He × 22.4 liters/mole
volume = 11.2 liters

31. **1** A catalyst speeds up a reaction's rate by making it easier for the reactants to come together. The energy needed to start the reactants is called the *activation energy*, which the catalyst changes. Without the catalyst the reaction would proceed at a slower rate. However, the potential energy of the reactants and products, and the net heat of the reaction, are the same with or without the catalyst.

The standard diagram of energy changes during a chemical reaction shows these concepts.



32. **4** Reference Table G shows 1 mole of each of these materials burning (oxidizing) to produce $CO_2(g)$ and $H_2O(\ell)$. The ΔH value tells how many kilocalories are released in this process. The negative sign (−) in front of the ΔH values means these reactions are all exothermic (they release heat). Clearly, $C_6H_{12}O_6(s)$ releases the largest amount of heat, 669.9 kcal/mol.

33. **4** Energy is being absorbed in this reaction. *By definition*, that makes it endothermic.

Wrong Choices Explained:

(1) Exothermic reactions release energy as a product. That's not the case in this reaction, which requires energy as a reactant.

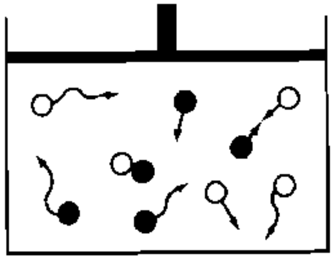
(2), (3) These are not even feasible choices since they have the wrong word matched to the process described.

Remember: EXO = "out" ± energy released as a product
ENDO = "in" ∓ energy absorbed as a reactant

34. **3** The rate of a reaction depends on several factors, including how often molecules or atoms collide with one another. By increasing the pressure on a gas system, the molecules or atoms are "squeezed" closer, so they collide more frequently. This makes their rate of reaction increase.

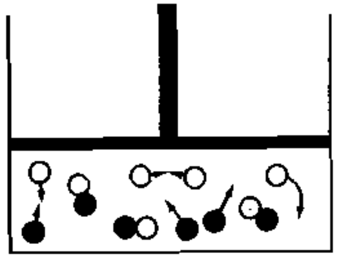
This is analogous to reactions between solutions of compounds, which go at a faster rate if the concentration of solutions is increased. Increasing the pressure on a gas is like increasing its concentration, in moles/liter.

760 torr



Lower pressure,
less frequent collisions,
lower reaction rate

900 torr



Higher pressure,
more frequent collisions,
higher reaction rate

35. **4** In general, an equilibrium constant will change *only* with an increase in temperature. The concentration of reactants or products and the pressure may be adjusted, causing a shift in the equilibrium amounts present, but the equilibrium constant will not have changed. Only temperature changes the constant.

In this specific example, we have a very insoluble material, AgCl(s), which, nevertheless, does dissolve to a tiny extent as ions.

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

In this case the equilibrium constant is modified and becomes a solubility product constant.

$K_{eq} = K_{sp} = [AG^+][Cl^-]$ without a denominator
[ ] brackets indicate chemical concentration in moles/liter.

An increase in temperature will permit more moles of AgCl(s) to dissolve as ions, so the K_{sp} will change.

Wrong Choices Explained:

(1), (2) Changing the concentration of either Ag^+ or Cl^- will cause an opposite shift in the concentration of the other ion. This is called the *common ion effect*. However, the K_{sp} equilibrium constant will remain the same throughout the shift. This can be illustrated by the following equations:

$$K_{sp} = [Ag^+][Cl^-] \quad \text{or} \quad [\mathbf{Ag^+}][Cl^-] \quad \text{or} \quad [Ag^+][\mathbf{Cl^-}]$$

The concentration of each may
shift, but not
their product, K_{sp}.

(3) Pressure causes shifts in gaseous systems but, like concentration, cannot change the K_{eq} of those systems. In this solid/ion equilibrium the pressure has no effect at all—either on K_{eq} or on equilibrium concentrations.

36. 4 Acids ionize to release an H^+ ion. According to the Brönsted Theory, acids donate protons—which is the same as saying they donate H^+ ions.

$$HX \rightleftharpoons H^+ + X^-$$

The degree to which an acid ionizes is reflected in the K_{eq} for the reaction, which for acids is designated as K_a.

$$K_{eq} = K_a = \frac{[H^+][X^-]}{[HX]}$$

An acid that ionizes to a large extent will be "strong," because it releases a lot of H^+ ions, resulting in a large K_a value.

Reference Table J gives the K_a values for certain acids.

Acid	K_a	Relative Strength
HNO_2	5.1×10^{-4}	
H_2S	1.0×10^{-7}	weakest (smallest K_a)
CH_3COOH	1.8×10^{-5}	
H_3PO_4	7.1×10^{-3}	strongest (largest K_a)

37. 2

$$K_a = \frac{[\text{conc. of product ions}]}{[\text{conc. of remaining acid}]} = \frac{[H^+][NO_2^-]}{[HNO_2]}$$

38. 3 The pH scale is defined so that pure water (at 25°C) has a pH of 7.00. The reason for this comes from the amount of ionization in pure water:

H_2O	$\rightleftharpoons$	$H^+ + OH^-$
0.0000001 mole/liter will ionize.		0.0000001 mole/liter of each of these is produced.

Therefore, the concentration of H^+ in pure water is 0.0000001, or 1×10^{-7}, mole/liter.

$$pH = -\log(H^+)$$
$$pH = -\log(1 \times 10^{-7}) = -\log(10^{-7}) = -(-7) = 7$$

39. 4 Mg is a fairly active metal and readily reacts with most acids to release hydrogen.

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

The other metals, Au (gold), Ag (silver), and Cu (copper), are relatively unreactive with acids.

Reference Table L confirms this. The reduction potential for Mg^{2+} is *lower* than that for H^+, while the reduction potentials for the ions of the other atoms are *higher* than that of H^+.

Half-reaction	Standard Electrode/Reduction Potential, in volts
$Au^{3+} + 3e^- \rightarrow Au(s)$	$E° = 1.50$
$Ag^+ + e^- \rightarrow Ag(s)$	$E° = 0.80$
$Cu^{2+} + 2e^- \rightarrow Cu(s)$	$E° = 0.34$
$2H^+ + 2e^- \rightarrow H_2(g)$	$E° = 0.00$
$Mg^{2+} + 2e^- \rightarrow Mg(s)$	$E° = -2.37$

If Au, Ag, or Cu were to react with H^+, each would form its corresponding ion. But these ions would recapture the electron more readily than the H^+ ion could, so it could not be reduced to H_2. In other words, Au, Ag, and Cu will not give electrons to H^+, and so no reaction can occur between them and the acid. However, Mg^{2+}, when formed in the reaction, will not compete with H^+ for the electron. In other words, Mg(s) will react with acids.

40. 3 Conjugate pairs are any two species (molecules or ions) that are converted into one another by a mere gain or loss of H^+.

$$HCl \rightleftharpoons H^+ + Cl^- \quad \text{Conjugates}$$

In the full equation there are two such conjugate pairs.

$$HCl + H_2O \rightleftharpoons H_3O^+ + Cl^- \quad \text{Conjugate acid-base pairs}$$

Wrong Choices Explained:

(1) HCl donates an H^+ during the forward reaction; H_3O^+ donates an H^+ during the reverse reaction. Both are serving as Brönsted acids, but they are not a conjugate acid-base pair.

(2) HCl is giving its H^+ to H_2O in the forward reaction. We say the HCl is acting as the acid and the H_2O is acting as the base. This is an acid-base pair, but not a *conjugate* acid-base pair, because they are not formed from one another by an H^+ exchange.

(4) H_2O accepts an H^+ during the forward reaction; Cl^- accepts an H^+ during the reverse reaction. Both are serving as Brönsted bases, but they are not a conjugate acid-base pair.

In summary:

HCl	+	H_2O	$\rightarrow$	H_3O^+	+	Cl^-	(forward reaction)
Brönsted acid		Brönsted base					

HCl	+	H_2O	$\leftarrow$	H_3O^+	+	Cl^-	(reverse reaction)
				Brönsted acid		Brönsted base	

41. 2 Write the underlying half-reactions of this balanced ion equation.

$$Sn^{2+} \rightarrow Sn^{4+} + 2e^- \quad \text{(oxidation)}$$
$$2Fe^{3+} + 2e^- \rightarrow 2Fe^{2+} \quad \text{(reduction)}$$

Note that for every Sn^{2+} ion there are 2 electrons lost during oxidation. So, for 1 mole of Sn^{2+}, 2 moles of electrons will be lost.

42. 3 The molecule must add up to be neutral (0 charge). The oxidation numbers of H and O are standard: +1 and −2, respectively. We solve for S:

$$H_2SO_4$$
$$2(H) + S + 4(O) = 0$$
$$2(+1) + S + 4(-2) = 0$$
$$2 + S - 8 = 0$$
$$S - 6 = 0$$
$$S = 6$$

43. 1 Write the underlying half-reactions of this balanced equation. Begin by assigning the oxidation number to each atom of the equation. This helps you spot the *changes* in number for some of the elements involved.

$$\overset{+1}{4\,H}\ \overset{-1}{Cl} + \overset{+4}{Mn}\ \overset{-2}{O_2} \longrightarrow \overset{+2}{Mn}\ \overset{-1}{Cl_2} + \overset{+1}{2\,H_2}\ \overset{-2}{O} + \overset{0}{Cl_2}$$

Oxidation: 2 Cl^{-1} of the 4 Cl^{-1} appear to have changed.
$2\,Cl^{-1} \longrightarrow Cl_2^0 + 2e^-$

Reduction: the single Mn^{4+} ion appears to have changed.
$Mn^{+4} + 2e^- \rightarrow Mn^{+2}$

We see that manganese has been reduced since
a. its oxidation number was reduced from +4 to +2, and
b. it *gained* electrons, a change that is the heart of reduction.

Wrong Choices Explained:

(2), (3) These are not even feasible choices since they match the words "reduced" and "oxidized" with the wrong oxidation number changes.
"Reduced" means the oxidation numbers decreased.
"Oxidized" means the oxidation numbers increased.
These choices have it backward.

(4) This choice is correct in stating that a change from +2 to +4 would be oxidation; however, it is not true of the manganese in the given equation. As seen above, it was the chlorine that was oxidized, for its oxidation number went from −1 to 0.

44. 3 Write the underlying half-reactions.

$$\text{Oxidation. } \overset{0}{2Na} \rightarrow 2Na^+ + 2e^-$$
$$\text{Reduction. } 2H^+ + 2e^- \rightarrow \overset{0}{H_2}$$

The Na is oxidized; the H^+ (in the H_2O) is reduced.

45. 1 Note oxidation number changes and write the underlying half-reactions.

Oxidation: $3\,\overset{0}{Cu} \longrightarrow 3\,\overset{+2}{Cu} + 6e^-$
Reduction: Two of the nitrogens in the 8 HNO_3 are changed. You must determine for yourself that N within the HNO_3 (within the NO_3^- ion) has an oxidation number of +5.

$$2\,\overset{+5}{N} + 6e^- \rightarrow 2\,\overset{+2}{N} \text{ (within NO)}$$
$$\text{(within } HNO_3\text{)}$$

The N^{+5} ion was reduced; however, the agent that allowed it to obtain electrons was the Cu^0 that lost electrons. The Cu^0 is called the reducing agent. Note that it was itself oxidized.

Reducing agents (here Cu^0) are oxidized.
Oxidizing agents (here N^{+5}) are reduced.

46. 1 The reduction half-reaction is

$$Cu^{+2} + e^- \rightarrow Cu^{+1}$$

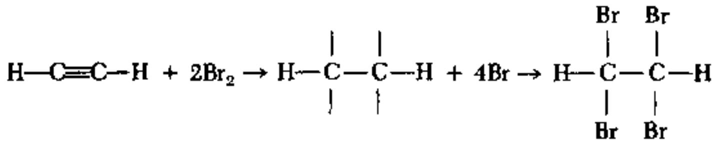
One electron is needed to reduce one ion of Cu^{+2}. Therefore, 1 mole of electrons is needed to reduce 1 mole of Cu^{+2} ions.

47. 2 Isomers of a compound have the same empirical formula, but different structural formulas.

Choice	Structural Formula	Empirical Formula
Original compound	CH_3COOCH_3	$C_2H_6O_2 = CH_3O$
(1)	CH_3OCH_3	C_2H_6O
(2)	CH_3CH_2COOH	$C_3H_6O_2 = CH_3O$
(3)	CH_3COCH_3	C_3H_6O
(4)	$CH_3CH_2CH_2OH$	C_3H_8O

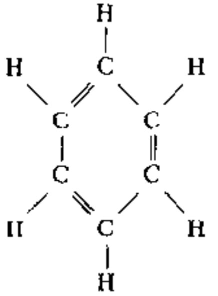
The empirical formulas for the original compound and choice (2) are the same.

48. 4 This is an addition reaction. The triple bond is opened, allowing places for 4 atoms to join the molecule.



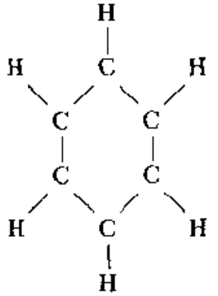
The answer can also be determined by the fact that "$2Br_2$" is given in the equation. These 4 Br atoms must appear somewhere in the product(s) of the reaction. Only choice (4) shows 4 Br atoms, thereby balancing the equation.

49. 1 These symbols represent the benzene ring. The first symbol is a shorthand way of writing the Kekulé structure:



C_6H_6, Kekulé structure

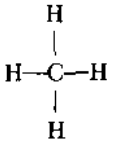
After this structure was proposed, it was discovered that the electrons in the 3 double bonds are not actually localized (fixed) on any carbon atoms. They delocalize, or spread, over the whole ring. This gave rise to the second symbol for benzene:



C_6H_6, delocalized electrons

Wrong Choices Explained:

(2) Methane is CH_4 and contains only single bonds

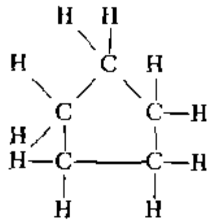


Ch_4

(3) Acetylene is the commercial name for ethyne, C_2H_2, which is straight chained and contains a triple bond.

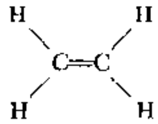
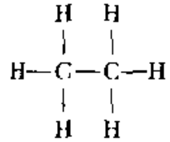
H—C≡C—H C_2H_2

(4) Cyclopropane is a ring, but contains only single bonds between its *five* carbon atoms.



C_5H_{10}

50. 2 Saturated hydrocarbons contain only single bonds between carbon atoms. The suffixes of the compound names tell you what bonds are present in the compound.

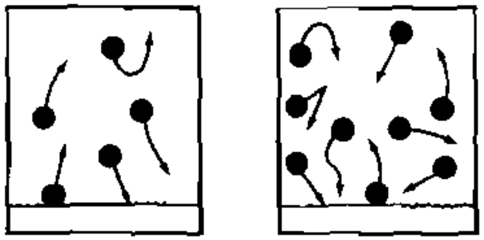
Choice	Name	Structure	
(1)	ethene	"-ene" indicates the presence of a double bond	
(2)	ethane	"-ane" indicates only single bonds	
(3)	ethylene	another name for ethene	
(4)	ethyne	"-yne" indicates the presence of a triple bond	H—C≡C—H

51. 3 C_nH_{2n-2} is the general formula for the alkyne series, in which C_2H_2 is the first member. Choice (1), C_nH_n, looks temptingly correct, but it's not applicable for the rest of the series: C_3H_4, C_4H_6, C_5H_8.

52. 3 The molecules will be squeezed into a smaller volume, but their numbers will not change. Molecules can't disappear.

53. 2 With an increase in temperature more water molecules will evaporate as they receive the energy necessary to break free of the bonds holding them in the liquid state. In the gas state these molecules have more random motions than they did when in the liquid state. This increase in randomness is what an entropy increase is defined as.

20°C



25°C More molecules in gas state lead to higher entropy.

H_2O

54. 1 The product of the concentrations of Ba^{2+} ions and SO_4^{2-} ions must remain constant.

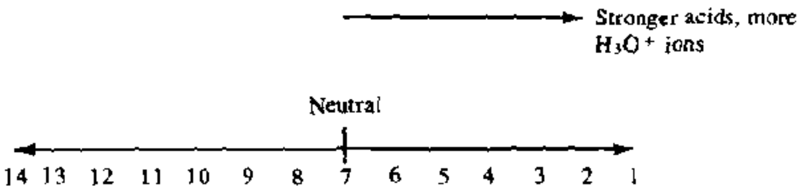
$$K_{sp} = \text{a constant} = [Ba^{2+}][SO_4^{2-}]$$

If $[Ba^{2+}]$ is increased, then $[SO_4^{2-}]$ must decrease in the same proportion so the product of both remains constant.

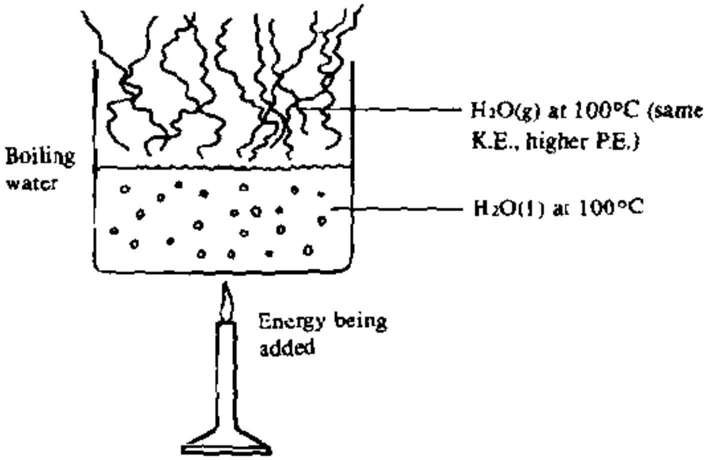
$$K_{sp} = [\mathrm{Ba}^{2+}][SO_4^{2-}]$$

See also question 35.

55. 1 More H_3O^+ ions mean a more acidic solution. The pH scale is designed so that stronger acids have *smaller* pH numbers.



56. 2 This question describes the boiling of water at 100°C. Heat energy is needed to boil the water. This energy allows molecules to break free of the bonds holding them together as a liquid and become "freer" as a gas. This energy does not make the steam hotter than the boiling water; *both are at 100°C*. Since the heat energy did not cause the kinetic energy of the water molecules to change, it must have gone into increasing the potential energy of the molecules.

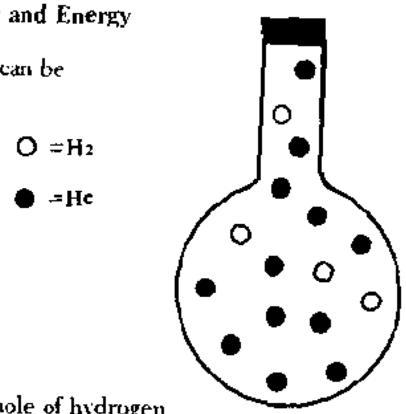


PART TWO

Group 1—Matter and Energy

57. 1 The partial pressure of a gas can be calculated from its mole fraction:

1 mole of hydrogen
+ 3 moles of helium
4 moles of gas in flask



a. Mole fraction of hydrogen $= \dfrac{\text{1 mole of hydrogen}}{\text{4 moles of gas, total}} = 0.25$

b. Partial pressure of hydrogen = mole fraction × total pressure

$$p_{H_2} = 0.25 \times 400 \text{ torr}$$
$$p_{H_2} = 100 \text{ torr}$$

58. 3 One mole is defined as 6.02×10^{23} molecules, just as one dozen eggs is defined as 12 eggs. The nitrogen here has only 1.5×10^{23} molecules, which is a fraction of a mole.

$$\text{amount of nitrogen} = \frac{1.5 \times 10^{23} \text{ molecules}}{6.02 \times 10^{23} \text{ molecules/mole}} = 0.25 \text{ mole}$$

Any material having this same number of moles must have that same number of molecules, namely, 1.5×10^{23}. Choice (3) is the answer because it also contains 0.25 mole. That it mentions O_2 is irrelevant, just as the H_2, He, and Ne named in the other choices are irrelevant. Only the moles matter.

59. 4 "Fusion" means melting: (s) → (ℓ). If you didn't know this, you might recall that "heat of fusion" is the name applied to the heat needed to melt materials, like ice. Or you could get the answer by eliminating the other choices, which are clearly wrong.

Wrong Choices Explained:

(1) Condensation means (g) → (ℓ).

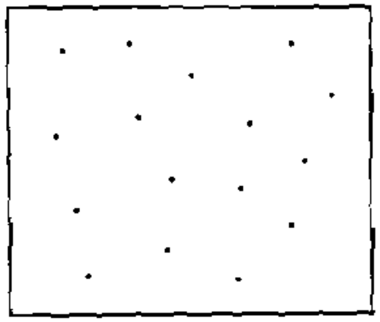
(2), (3) Vaporization and evaporation both mean (ℓ)→(g).

60. 1 The gas laws, such as Charles' Law, Boyle's Law, and $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$, are technically for "ideal" gases. The behaviors of most real gases can be adequately (but not perfectly) calculated by these equations. Here are the differences between ideal and real gases.

Ideal Gases	Real Gases
a. Molecules are mere points, having no volume contribution to the total volume of the gas.	a. Molecules themselves occupy a small fraction of the volume.
b. Molecules of the gas have no mutual attractions between themselves.	b. Attractive forces exist between molecules in varying degrees.

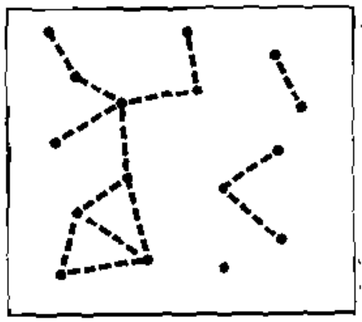
If the temperature of a real gas is high, then the attractive forces between molecules will have a proportionally diminished effect because the molecules have a higher velocity, while the attractive forces have remained constant. If the pressure of a real gas is low, then the molecules are further apart, so their attractive forces are lessened and their own volumes are even less of the total volume. These two conditions, high temperatures and low pressures, cause a real gas to behave most like an ideal gas.

Ideal Gas



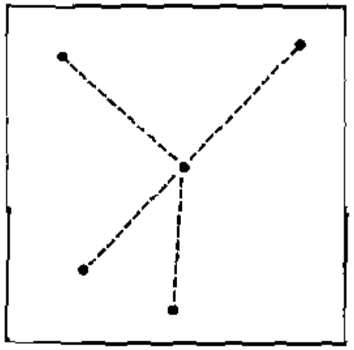
Molecules are points, with no attractive forces.

Real Gas



Molecules have measureable volumes; attractions exist between neighbors.

Real Gas at High *T* and low *P*



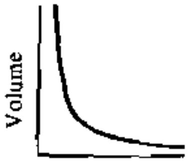
Fewer molecules mean less proportion of the total volume.
Molecules further apart mean less attractive forces between them.
Faster motion means attractive forces are diminished in their effects.

Result: Behavior closer to an ideal gas.

Fewer molecules mean less proportion of the total volume.
Molecules further apart mean less attractive forces between them.
Faster motion means attractive forces are diminished in their effects.

Result: behavior closer to an ideal gas.

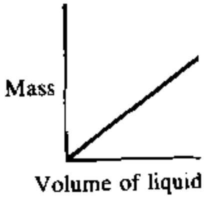
61. 2 The volume and pressure of an ideal gas are inversely proportional. Curve *B* shows such a relationship.



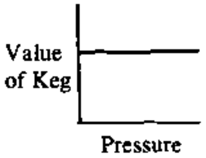
Wrong Choices Explained:

(1) *Curve A describes a direct relationship.* As one variable increases, the other variable also increases, though not by the same proportion. This particular curve might represent a square-root relation.

(3) Curve *C* describes a direct proportionality. As one variable increases, the other increases by the same proportion. The relationship between volume and mass of a liquid might be an example described by this type of curve.



(4) Curve *D* describes a constant. The value of the *y* variable does not change if the *x* variable is changed. This is true, for example, of the value of an equilibrium constant if pressure is varied.



Group 2—Atomic Structure

62. **3** The three fundamental particles of chemistry are the *electron, proton,* and *neutron*. Their exact masses are given below. For most purposes their masses are rounded off to 0, 1, and 1 amu, respectively, for e, p, and n. For most purposes the proton and neutron are regarded as having the same mass, 1 amu. The *deuteron* is not a fundamental particle, but the nucleus of deuterium (an isotope of hydrogen), and consists of a proton together with a neutron. Its mass is 2 amu.

Particle	Exact Mass, in amu	Exact Mass, in grams
electron	0.00055	9.11×10^{-28}
proton	1.00728	1.670×10^{-24}
neutron	1.00867	1.672×10^{-24}
deuterium (n + p)	2.01595	3.342×10^{-24}

63. **2** A quick approach is to note that the element with atomic number 10 will have 10 electrons, while Na (atomic number 11) will have 11 electrons. If Na should give away one electron, the Na^+ ion that results will have only 10 electrons remaining. So, the electron configuration of Na^+ must be identical to that of the element with atomic number 10.

element with atomic number 10	$10e^-$	same number of electrons
Na^+ ion = Na atom − $1e^-$ = $11e^-$ − $1e^-$ =	$10e^-$	

A longer approach would be to write out the electron configurations for the choices:

Choice	Number of Electrons	Electron Configuration
element with atomic number 10	10	$1s^22s^22p^6$ (same)
(1) Na	11	$1s^22s^22p^63s^1$
(2) Na^+	10 (one less than Na)	$1s^22s^22p^6$ (same)
(3) Cl	17	$1s^22s^22p^63s^23p^5$
(4) Cl^-	18 (one more than Cl)	$1s^22s^22p^63s^23p^6$

64. **4** Electron dot symbols show the outer, valence electrons only. These outer electrons can be only in the *s* or *p* orbitals, because *d* and *f* orbitals are always within an outer (valence) shell of *s* or *p* electrons. In other words, dot pictures can be translated into orbital notation. Using the periodic table we make a visual inspection to see if any of the orbital notations in the question match a known metalloid.

Choice	Dot Symbol	Equivalent Orbital Notation	Any Metalloid with Same Notation?
(1)*	X ·	s^1	No, this designates the active metals of Group IA. (*See note)
(2)	X :	s^2	No, this designates He and Group IIA metals.
(3)	:X: (8 dots)	s^2p^6	No, this designates Group O inert gases from Ne down.
(4)	·X· (5 dots)	s^2p^3	Yes, the element As (atomic number 33), which is a metalloid in Group VA, ends with $4s^24p^3$.

* *Note:* This may have been a misleading question. Choice (1) could be the dot symbol of H with its one s^1 electron. And on the Periodic Table, hydrogen is surrounded by a heavy-lined box, which means that it, too, can be considered a metalloid. Choice (4), however, was the official answer.

65. **2** Fluorine (atomic number 9) must have 9 electrons whether it is in the ground state or in an excited state. Choices (3) and (4) are automatically eliminated because they show 10 electrons.

Choice (1) is the ground state of **F**: $1s^22s^22p^5$
Choice (2) is an excited state of **F**: $1s^22s^12p^6$

In the excited state, one electron (from the 2*s* level) was promoted upward (to the 2*p* level).

$1s^22s^{2\,1}2p^{5\,6}$ One electron promoted

66. 1 According to Reference Table F, the half-life of radioactive ^{32}P is 14.3 days. This means that after each 14.3 days only half of the previous amount of ^{32}P remains.

Time Passed in days	Amount of ^{32}P Left, in grams
at start	32
14.3	16
28.6	8
42.9	4
57.2	2
71.5	1

This can also be solved by formula.

$$n = \text{number of half-lives passed} = \frac{\text{time}}{\text{half-life}} = \frac{71.5 \text{ days}}{14.3 \text{ days}} = 5$$

$$\text{amount remaining} = (\text{original sample})(\tfrac{1}{2})^n = (32 \text{ g})(\tfrac{1}{2})^5 = (32 \text{ g})(\tfrac{1}{32}) = 1 \text{ g}$$

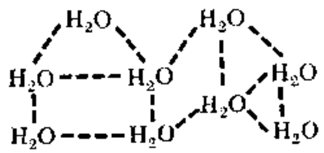
Group 3—**Bonding**

67. 1 The presence and strength of ionic bonds in a compound can be determined from the electronegativity differences between its atoms. A difference of 1.7 or greater indicates ionic bonds.

Choice	Compound	Electronegativity Differences Obtained from Reference Table I	Type of Bond
(1)	NaBr	2.8 − 0.9 = 1.9	ionic
(2)	HBr	2.8 − 2.1 = 0.7	polar covalent
(3)	$C_6H_{12}O_6$	Organic compounds are generally covalent.	
(4)	CO_2	3.5 − 2.5 = 1.0	polar covalent

68. 2 To have a nonpolar covalent bond, the electronegativity difference must be 0. This occurs only when identical atoms share electrons, as is true of diatomic molecules like F_2, Cl_2, H_2, and O_2. The electronegativity difference between atoms in the other choices will be greater than 0 but less than 1.7. Their bonds will be polar covalent.

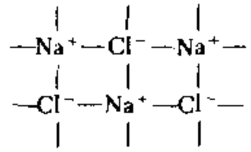
69. 4 When the temperature is low enough (0°C), water molecules will form a crystalline arrangement we call solid ice. The bonds that hold neighboring molecules of water together are only moderate in strength, being primarily hydrogen bonds between molecules. This is the essence of a *molecular solid:* molecules bound to each other by moderate attractive forces. The molecules are close to one another but they have not lost their identity as separate individual molecules.



Weak van der Waals and polar bonds

Wrong Choices Explained:

(1) *Ionic solids* are composed of an enormous lattice of + and − ions. All ions are attached equally to their neighbors in one giant molecule or lattice. NaCl is an example. The ionic bonds in ionic solids are very strong.

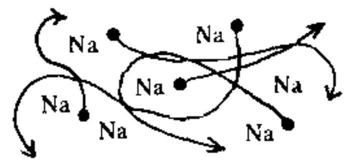


Strong ionic bonds in a giant lattice of ions

(2) In *network solids*, neighboring molecules are joined by actual covalent bonds, by actual sharing of electrons. This makes them somewhat stronger than molecular solids, which bond not by covalent bonds, but by weaker polarized attraction. Quartz, made from SiO_2 molecules, is an example of a network solid.

(3) Most solid metals such as Cu, Al, Mg, Na, and Li, are *metallic solids.* The atoms of these metal elements are held to each other because they completely exchange their outer electrons. These freely transferred electrons form a "sea of charge" that holds neighboring atoms together. This also explains why metals can conduct electricity well.

● Outer electron



Outer electrons are delocalized in a "sea of electrons" bonding the metal atoms

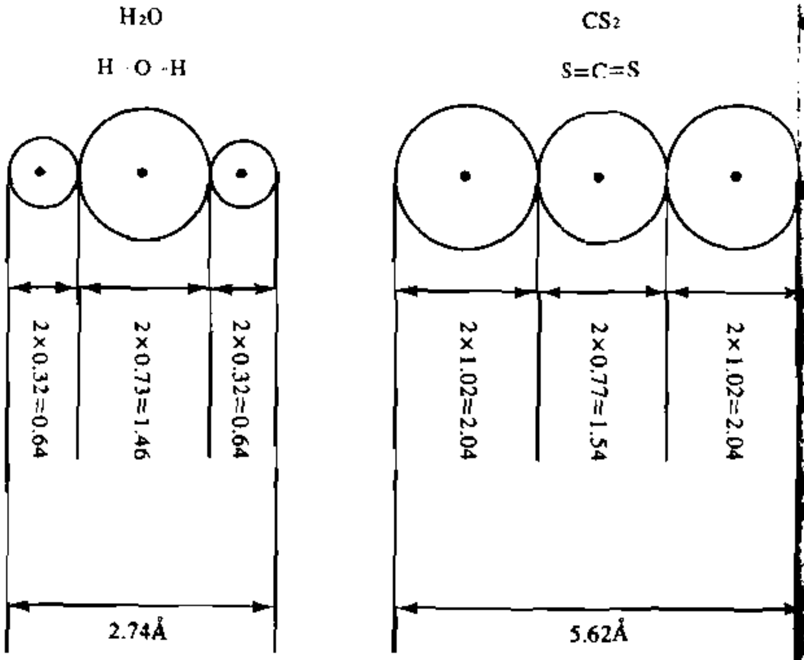
70. 3 H, N, and Br represent atoms. These elements do not remain single at STP conditions. They join with themselves to form the diatomic (2-atom) molecules H_2, N_2, and Br_2. The element krypton, however, being inert, does not bond even with itself. Kr atoms are also molecules.

The formula for a molecule of gas identifies the structure of the gas at specified conditions of temperature and pressure, here at STP. A krypton molecule is simply Kr; hydrogen, nitrogen, and bromine molecules are H_2, N_2, and Br_2, respectively.

71. 4 A H—O bond is slightly polar covalent because the electronegativity difference between O and H is 3.5 − 2.1 = 1.4. A C—S bond is nonpolar covalent because the two elements have the same electronegativities, 2.5 − 2.5 = 0. Water molecules will therefore be polar, while carbon disulfide molecules will be nonpolar. And water molecules will have attractive forces that hold them together as a solid or liquid, while CS_2 will have no such intermolecule attractions. That's why water does not vaporize to the extent that CS_2 does. H_2O has a low vapor pressure; CS_2 has a high vapor pressure.

Wrong Choices Explained:

(1) H_2O is actually a smaller molecule than CS_2, but size is not the issu here. Molecule size can be estimated using the covalent atomic radius given the Periodic Table. Remember that the atom diameters below are *twice* th radius given in the table.



(2) H_2O is actually smaller in molecular mass than CS_2. This suggests th it should vaporize more easily, but it doesn't—because of the attractive bond holding H_2O molecules to each other.

$$H_2O = 2 + 16 = 18 \text{ amu}$$
$$CS_2 = 12 + 32 + 32 = 76 \text{ amu}$$

(3) Neither H_2O nor CS_2 has ionic bonds within or between its molecules

Group 4—**Periodic Table**

72. 2

Period	*Diatomic Elements*
1	H_2
2	N_2, O_2, and F_2
3	Cl_2
4	Br_2

73. 1 The metals of Group IIA are also known as alkaline earth metals. Mg is in that group, together with Be, Ca, Sr, and Ba.

Wrong Choices Explained:

(2) Zn is in Group IIB. All the B-group elements are known as transition elements.

(3) Li is in Group IA. These are known as alkali metals.

74. 4 To become an S^{2-} ion, the S^0 atom must gain two electrons.

$$S^0 + 2e^- \rightarrow S^{2-}$$

These two electrons go into the two openings of the outer $3p$ sublevel of sulfur.

$$S^0 = 1s^2 2s^2 2p^6 3s^2 3p^4$$
$$S^{2-} = 1s^2 2s^2 2p^6 3s^2 3p^6$$

These extra two electrons will increase the repulsive forces between all the outer electrons so that the $3s$ and $3p$ sublevels will be slightly swelled, or expanded, in size.

In summary, the S^{2-} ion has a larger radius because it has more electrons than the S^0 atom.

75. 3 Transition metals, which are located in the B groups on the Periodic Table, usually give colored ions in solution. For example, Cu^{2+} is blue, Fe^{3+} is rust red, Ni^{2+} is green, Mn^{7+} is dark purple, and Cr^{3+} is dark green.

76. 2 Fluorine is element number 9; oxygen is element number 8. The F nucleus has 9 protons, giving it a +9 charge. The O nucleus has 8 protons, giving it a +8 charge. The outermost electrons of both elements are in the second principal energy level. However, the stronger +9 charge in F's nucleus attracts these electrons more than does O's nucleus. This extra nuclear charge explains two facts:

a. The F atom has a slightly smaller radius than the O atom.
b. The F nucleus keeps a stronger hold on its outer electrons. Its ionization energy is higher than oxygen's (see Reference Table I).

Group 5—**Mathematics of Chemistry**

77. 1 The easiest way to handle this question is simply to find the molecule with mass 16:

$$CH_4 = 12 + 1 + 1 + 1 + 1 = 16$$

The others are much heavier than 16, and, in fact, none of them reduces to an empirical formula of CH_4.

78. 4 One molecule of $C_6H_{12}O_6$ clearly contains 12 atoms of hydrogen. Therefore, 1 mole of this material has in it 12 moles of hydrogen.

Here's an analogy. If a box contains 6 hamburger rolls, how many dozen rolls are held by a dozen boxes? Answer: 6 dozen rolls. Of course, altogether there would be (6 × 12) = 72 rolls. The question didn't ask how many rolls there are, however, but how many *dozens* of rolls.

Similarly, this chemistry question doesn't ask how many molecules of hydrogen there are, but how many *moles* of hydrogen there are. There are 12 moles of H atoms; this does mean there are $12(6.0 \times 10^{23})$ atoms of hydrogen, but that's not asked for by the question. Choice (2) is wrong for this particular question.

79. 2 One mole of any gas at STP occupies 22.4 liters. If we know the mass of 22.4 liters of the gas, then we'll have its mole mass, which, in turn, is also its molecular mass.

a. Mole mass of this gas = $\frac{2.0\text{ g}}{\text{liter}} \times \frac{22.4\text{ liters}}{\text{mole}} = 44.8\ \frac{\text{g}}{\text{mole}}$

b. Molecular mass must be 44.8 amu, which rounds to 45 amu.

80. 1

heat absorbed = (mass of water) × (temperature change, in °C)

$$q = (m) \times (\Delta t)$$
$$q = (5\text{ g}) \times (15°\text{C} - 10°\text{C})$$
$$q = (5\text{ g}) \times (5°\text{C})$$
$$q = 25\text{ cal}$$

81. 4 Obtain the atomic mass of each element in *1* molecule of C_2H_6.

carbon: 2 × 12 = 24 amu
hydrogen: 6 × 1 = 6 amu

$$\text{carbon : hydrogen ratio} = \frac{\text{carbon mass}}{\text{hydrogen mass}} = \frac{24\text{ amu}}{6\text{ amu}} = \frac{4}{1} = 4:1$$

Group 6—Kinetics and Equilibrium

82. 4 This reaction shows the formation of 1 mole of the KCl directly from its elements.

$$K(s) + \tfrac{1}{2}Cl_2(g) \rightarrow KCl(s)$$

Reference Table E presents the energy information for such synthesis, or formation, reactions:

potassium chloride, KCl(s) $\Delta H_f^\circ = -104.2$ Kcal/mol
$\Delta G_f^\circ = -97.6$ Kcal/mol

The ΔH value tells you how much heat is released in this reaction (the minus sign means the reaction is exothermic); the ΔG value tells you that the reaction occurs spontaneously (the minus sign means the reaction was spontaneous). Reactions are spontaneous only if ΔG is negative. Values for ΔH or ΔS alone are not the determinants for spontaneousness.

83. 3 All of these compounds are highly insoluble in water. However, a tiny amount will dissolve. The extent to which they will dissolve is given by the "solubility product equilibrium constant," K_{sp}, shown in Reference Table K. The smallest K_{sp} value indicates the material that dissolves least, which would make the "most dilute" solution. That is clearly ZnS.

$$K_{sp}\text{ of ZnS} = 1.6 \times 10^{-23} = (Zn^{2+})(S^{2-})$$

84. 2 Memorize the equation: $\Delta G = \Delta H - T\Delta S$. Here is one interpretation of this equation.

ΔG	=	ΔH	−	$T\Delta S$
Energy that is "free" to drive a chemical reaction spontaneously.		Heat energy released by the reaction.		Energy lost to the increased randomness of the molecules.

85. 3 A K_{sp} expression contains only the concentration terms of the products. These concentration terms, $[X]$, are raised to an exponent equal to the coefficient in the balanced equation: $[x]^{\text{coefficient}}$.

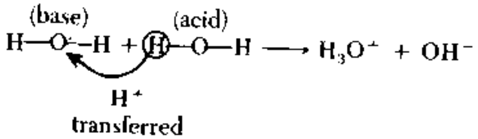
$$Mg(OH)_2 \longrightarrow Mg^{2+} + 2\,OH^-$$

$K_{sp} = [Mg^{2+}][OH^-]^2$ Coefficient becomes exponent.

86. 1 Reference Table K gives the K_{sp} for $BaSO_4$ as 1.1×10^{-10}.

Group 7—Acids and Bases

87. 3 Remember that an H^+ ion is the same thing as a proton. Here we see that one water molecule donates a proton (H^+), while the other water molecule accepts it.



Since water here acts as both a proton acceptor and donor, it is said to be acting both as a base and as an acid.

88. 2 Only a basic solution will cause litmus to turn from red to blue. We must find which one of these gases, when dissolved in water, forms a basic solution. By standard laboratory experiments and by direct memorization of required material, chemistry students will recognize that NH_3, ammonia, is the required gas.

$$NH_3 + H_2O \rightarrow NH_4OH \text{ (a base)}$$

The other gases form acid solutions.

Reference Tables J and K can also be used to determine what is formed when these gases are dissolved in, and thus react with, water. Reference Table J shows the extent to which they form acid solutions.

Reaction	K_a
$H_2S \rightarrow H^+ + HS$	1.0×10^{-7}
$H_2O + SO_2 \rightarrow H^+ + HSO_3^-$	1.7×10^{-2}
$H_2O + CO_2 \rightarrow H^+ + HCO_3^-$	4.4×10^{-7}
$NH_3 \rightarrow H^+ + NH_2^-$	very small, $< 10^{-36}$

From this we see that H_2S, SO_2, and CO_2 will form H^+ ions (acid) to a measurable extent. The NH_3 reaction is negligible, since its K_a is not even measurable. So, NH_3 can't make an acid solution. However, NH_3 can make a basic solution, as indicated in Reference Table K.

$$NH_3(aq) + H_2O \rightarrow NH_4^+ + OH^-$$

89. 4 The best conductor is the acid that ionizes most, releasing most H^+ ions. This can be determined from Reference Table J. We need to find the acid with the largest K_a value. This is HNO_3, whose K_a is given as "very large."

Wrong Choices Explained:

(1) C_2H_5OH is ethyl alcohol. It does not ionize and so cannot conduct electricity.

(2) CH_3COOH is acetic acid. It does ionize to form H^+ ions, but only to a small extent. Its K_a value is given as 1.8×10^{-5} in Reference Table J.

(3) $C_6H_{12}O_6$ is glucose. Like most organic compounds, it doesn't ionize and doesn't conduct electricity.

90. 4 This product, $[H_3O^+][OH^-]$, is always 1×10^{-14}. Its symbol is K_w.

Wrong Choice Explained:

(3) This would be the $[H_3O^+]$ in this solution. Remember that $[H_3O^+]$ and $[H^+]$ are equal, interchangeable terms.

$$K_w = (H^+)(OH^-)$$
$$1.0 \times 10^{-14} = (H^+)(1 \times 10^{-6})$$
$$(H^+) = \frac{(1.0 \times 10^{-14})}{(1 \times 10^{-6})} = 1 \times 10^{-8}$$

91. 1 Amphiprotic ions can either give out or take in a proton (which is a H^+ ion). Such ions can be located using Reference Table J because they appear on two lines and two sides in the table: first as an acid, second as a base. This is true of HSO_4^-.

$$H_2SO_4 \leftarrow H^+ + \underset{\text{(base)}}{HSO_4^-}$$

(reverse arrow from Reference Table J)

$$\underset{\text{(acid)}}{HSO_4^-} \rightarrow H^+ + SO_4^{2-}$$

Wrong Choices Explained:

(2) NH_4^+ can give out a H^+ ion, but cannot take one in.

$NH_4^+ \rightarrow H^+ + NH_3$ (in Reference Table J)

$NH_4^+ + H^+ \not\rightarrow$ does not occur

(3), (4) NO_3^- and Cl^- can take in an H^+ ion, but, since they obviously contain no H atoms, they can't give out an H^+ ion.

Group 8—**Redox and Electrochemistry**

92. 4 Once a chemical cell has reached equilibrium, there is no further chemical reaction occurring. No electricity can be made, and so the voltage is 0.

93. 2 Write the underlying half-reactions and obtain their potentials from Reference Table L.

oxidation: $2Al^0 \longrightarrow 2Al^{3+} + 6e^-$ $E° = +1.66$ (*see note)

reduction: $3Cu^{2+} + 6e^- \rightarrow 3Cu^0$ $E° = +0.34$

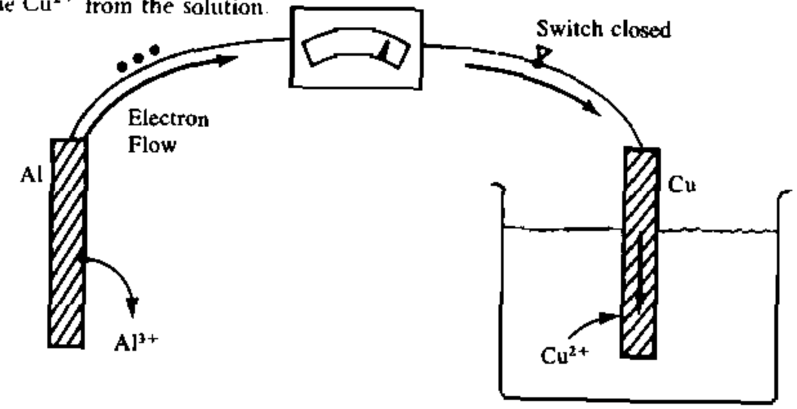
total $E° = +2.00$ volts

*Note:

a. This equation is the reverse of that printed in Reference Table L, which shows reductions, not oxidations. Therefore, $E°$ here has the sign opposite to that indicated in the Reference Table, +1.66 volts instead of −1.66 volt.

b. This equation is actually double the one that is printed in Reference Table L. The Cu^{2+} equation is triple the one given in Reference Table L. However, the $E°$ values are *not changed*. The same voltage, $E°$, is produced no matter how many times the equation and moles of reactants are multiplied.

94. 2 The half-reactions from the previous problem reveal that electrons are being given out by the Al^0 and being used up by the Cu^{2+} ions. They travel through the wire from the Al electrode over to the Cu electrode, then meet the Cu^{2+} from the solution.



95. 3 There are two ways to approach this problem.

a. Mg can replace a metal ion from solution only if the metal is less active than Mg. Ag (silver) is less active than Mg, so Mg can replace Ag^+ ions from solution.

$$Mg + 2Ag^+ \rightarrow Mg^{2+} + 2Ag$$

b. From an electric potential approach, if the Mg is to react it must be oxidized.

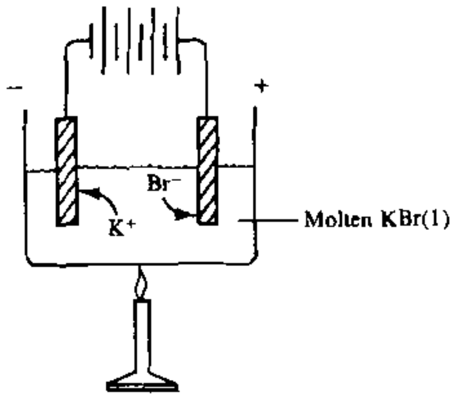
$$Mg^0 \rightarrow Mg^{2+} + 2e^-$$

The oxidizing agent must have a reduction potential greater than that of the Mg^{2+} ion just formed, or else the Mg^{2+} ion will recapture the electrons given off. From Reference Table L,

$Ag^+ + e^- \rightarrow Ag$	$E° = +0.80$ volt	Stronger than Mg^{2+}
$Mg^{2+} + 2e^- \rightarrow Mg$	$E° = -2.37$ volt	
$Sr^{2+} + 2e^- \rightarrow Sr$	$E° = -2.89$ volt	
$Ba^{2+} + 2e^- \rightarrow Ba$	$E° = -2.90$ volt	Weaker than Mg^{2+}
$Li^+ + e^- \rightarrow Li$	$E° = -3.00$ volt	

Only Ag^+ has a reduction potential that can capture the electrons from Mg and prevent Mg^{2+} from retaking them.

96. 1 When KBr is fused (melted), the K^+ and Br^- ions are free to migrate toward electrodes.

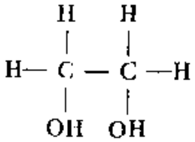


At the negative electrode, electrons are picked up by the K^+ ions, which is *reduction*.
$K^+ + e^- \rightarrow K^0$

At the positive electrode, electrons are released by the Br^- ions, which is *oxidation*.
$2Br^- \rightarrow Br_2 + 2e^-$

Group 9—Organic Chemistry

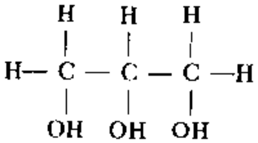
97. 4 A dihydroxy alcohol must contain *two* OH groups. Only $C_2H_4(OH)_2$ fits.



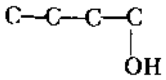
Wrong Choices Explained:

(1), (3) These contain OH^- groups, but they are not alcohols. They are inorganic bases.

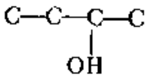
(2) $C_3H_5(OH)_3$ is a trihydroxy alcohol called glycerol.



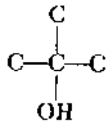
98. 3 A primary alcohol has the OH group attached to a carbon atom which is linked to *one other* C atom. This will be a C atom at the end of a straight chain.



Primary Alcohol
OH on a C attached to 1 other C atom

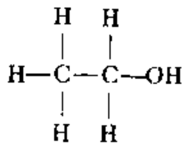


Secondary Alcohol
OH on a C attached to 2 other C atoms



Tertiary Alcohol
OH on a C attached to 3 other C atoms

Choice (3):



OH on a C attached to just 1 other C atom

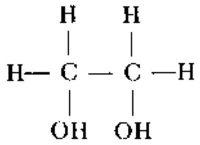
Wrong Choices Explained:

(1) This is an organic acid.
(2) This is an aldehyde.
(4) This is a secondary alcohol.

99. 1 $C_3H_5(OH)_3$ is glycerol. Its IUC name is 1,2,3-trihydroxypropane. See also explanation for choice (2) of question 97.

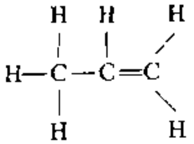
Wrong Choices Explained:

(2) Ethylene glycol is the common name for 1,2-dihydroxyethane.

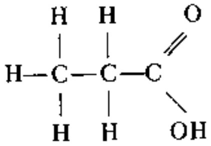


$C_2H_4(OH)_2$

(3) Propene is C_3H_6.

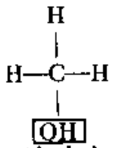


(4) Propanoic acid is C_2H_5COOH.



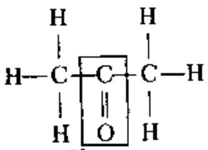
100. 2 Ketones have an O atom attached to some inner C atom.

Choice (1) CH_3OH



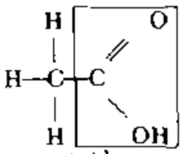
Alcohol

Choice (2) CH_3COCH_3



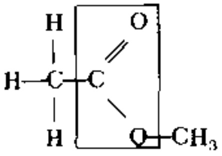
Ketone

Choice (3) CH_3COOH



Acid

Choice (4) CH_3COOCH_3



Ester

101. 3 Since these compounds are all alkanes, the electrical interactions and attractions between molecules are about equal. Therefore, the principal determinant of boiling point will be molecular mass. Materials with smaller molecular masses boil "easier," at lower temperatures, because they need less energy to be ejected from the liquid as a gas. The actual boiling point of each, taken from a reference book, is shown in the table below for comparison.

Choice	Formula	Mass in amu	Boiling Point, in °C
(1) butane	C_4H_{10}	58	−0.6
(2) ethane	C_2H_6	30	−88
(3) methane	CH_4	16	−162
(4) propane	C_3H_8	44	−42

Group 10—Applications of Chemical Principles

102. 2 Carbon can replace a metal ion only if carbon is "more active" than the element it is replacing. See question 95. Mg, Na, and Li are active metals, and so their oxides, MgO, Na_2O, and Li_2O, respectively, are very stable. They cannot easily be reduced to the pure metal, certainly not by carbon.

$$MgO + C \rightarrow \text{no reaction}$$
$$Na_2O + C \rightarrow \text{no reaction}$$
$$Li_2O + C \rightarrow \text{no reaction}$$

However, Zn is not an active metal. Carbon can replace it in its oxide.

$$2ZnO + C \rightarrow Zn + CO_2$$

This is the standard method for producing many such metals from their ores, usually oxides:

Zn from ZnO Pb from PbO
Fe from Fe_2O_3 Sn from SnO_2
Cu from CuO

103. 2 *Cracking* consists of taking large molecules and breaking them into smaller molecules. Crude oil contains a high amount of very large molecules with 20 to 36 carbon atoms and higher. These are heavy wax, asphalt, and tars. Cracking changes these into molecules with smaller numbers of carbon atoms.

gasoline—molecules with C_5 to C_{12} formulas
kerosene—molecules with C_{12} to C_{18} formulas

Wrong Choices Explained:

(1) *Oxidation* can mean either combining with oxygen or losing electrons.

Standard oxidation: $2Mg + O_2 \rightarrow 2MgO$
redox oxidation: $Mg + Cl_2 \rightarrow MgCl_2$
because $Mg^0 \rightarrow Mg^{2+} + 2e^-$

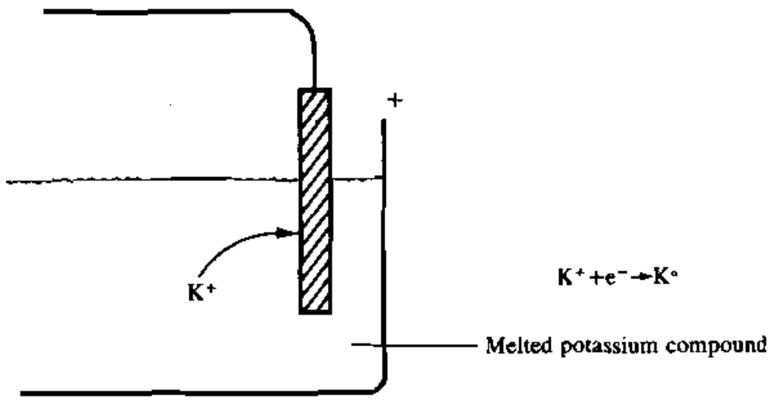
(3) The *Haber process* is the commercial method of making ammonia from nitrogen and hydrogen.
(4) The *contact process* is the commercial method of making sulfuric acid from sulfur or sulfide ores.

104. 1 In a chemical cell the positive electrode is where electrons are being *used* by the reactions; the negative electrode is where electrons are being *released* by the reaction. The underlying half-reactions for a nickel-cadmium battery are as follows:

$$Ni^{3+} + 3e^- \rightarrow Ni^0 \text{ reduction, at + electrode)}$$
$$Cd^0 \rightarrow Cd^{2+} + 2e^- \text{ (oxidation, at − electrode)}$$

The material that makes up the positive electrode must contain the Ni^{3+} ion, namely, $Ni(OH)_3$.

105. 4 Iron, tin, and lead are obtained commercially by reducing their oxides with coke (carbon). See question 102. Potassium, being an active metal, cannot be obtained from its salts or oxides by reaction with carbon. Electricity must be used. The potassium compound is first purified, then melted under high temperature. Electricity is passed through the liquid, and the K^+ ions migrate toward the positive terminal, where they are reduced to K^0.



106. 3 The contact process has three fundamental steps:
a. "Roast" the ore or sulfur to make sulfur dioxide.

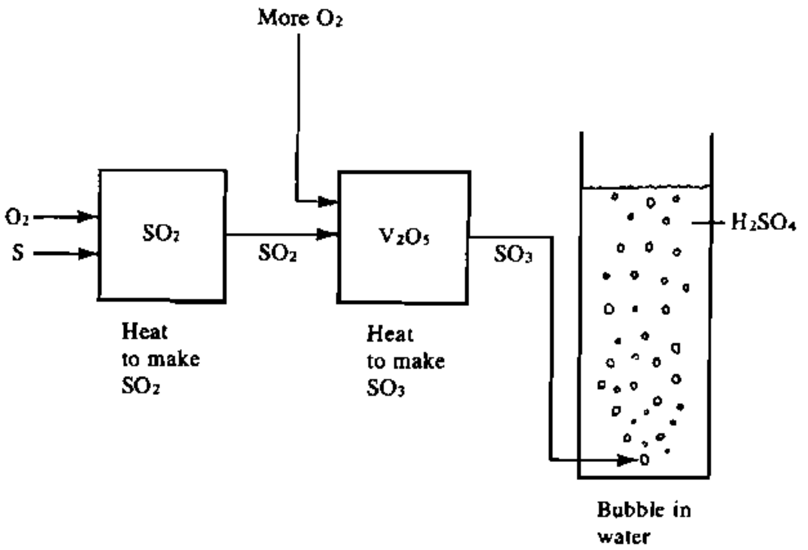
$$S + O_2 \rightarrow SO_2 \qquad 2FeS + 3O_2 \rightarrow 2FeO + 2SO_2$$

b. Oxidize sulfur dioxide to sulfur trioxide with a catalyst.

$$2SO_2 + O_2 \xrightarrow{V_2O_5} 2SO_3$$

c. Bubble the sulfur trioxide in water to make sulfuric acid.

$$SO_3 + H_2O \rightarrow H_2SO_4$$



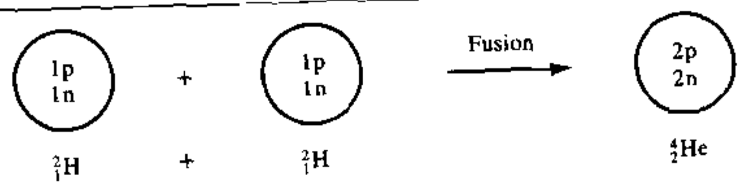
Wrong Choices Explained:
(1) The catalyst, vanadium pentoxide (V_2O_5), is used, not to make SO_3 directly from pure S, but from SO_2.
(2), (4) Ozone, O_3, is not part of the contact process. Choice (4) is also not balanced; it should be eliminated on that ground alone.

Group 11—Nuclear Chemistry

107. 2 ^{238}U disintegrates slowly into ^{206}Pb and other products. Reference Table L gives a half-life of 4.51×10^9 years, about 4½ billion years. As time passes, the ratio of lead to uranium slowly increases. This ratio can be used to date geologic formations.

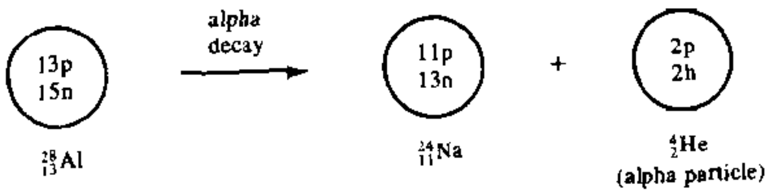
Wrong Choices Explained:
Radioactive compounds are used for these medical purposes. However, the material used must have a short half-life so that detection and treatment is not a prolonged health risk.
(1) For diagnosing thyroid disorders—radioactive iodine is used.
(3) For detecting brain tumors—radioactive potassium is used to follow the blood flow.
(4) For treating cancer—powerful radioactive sources are used to produce radiation that kills cancer cells.

108. 4 In this equation, lighter atoms are forming heavier atoms. This is the essence of fusion.

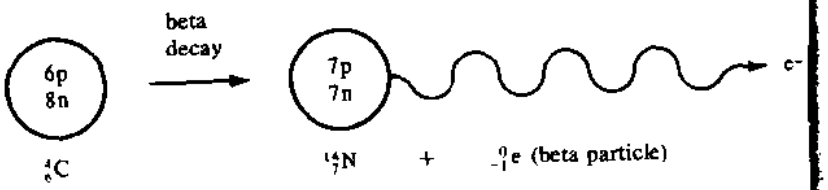


Wrong Choices Explained:

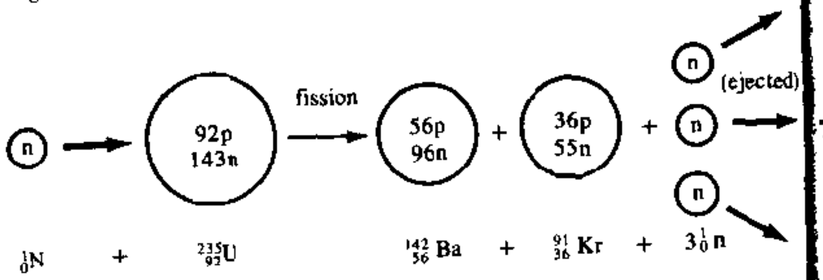
(1) Alpha decay is the process in which a larger atom disintegrates into a smaller atom and a helium nucleus (the alpha particle)



(2) Beta decay is the process in which a neutron of an atom disintegrates to produce a proton and an electron. The proton remains behind, resulting in a new element with an atomic number greater by 1 more than that of the starting element. The electron shoots away and is called a beta particle.



(3) Fission is the breaking of large nuclei into smaller ones, with the release of great energy. Fission is "triggered" by bombardment, usually by a neutron



109. 3 Heavy water, like ordinary water, has the formula H_2O. The difference is that the hydrogen atoms in heavy water are the isotope $^{2}_{1}H$ or $^{3}_{1}H$ instead of ordinary $^{1}_{1}H$.

Heavy water is used to remove the heat energy of fission reactors from the "core" to a heat exchanger. (Eventually, this heat is used to boil regular water into steam to drive electric generators.) Heavy water is also used as a moderator. Moderators cause neutrons to slow down so they become more effective as triggers to start more fission reactions. See question 108.

Wrong Choices Explained:

(1), (2) Cd and B can stop and capture neutrons, so they are used in control rods to reduce or shut off a nuclear reaction.

(4) Graphite, C, is like heavy water—it moderates or slows down neutrons to make them more effective triggers of fission. Unlike heavy water, C is not a coolant.

110. 4 Most fission reactors started out using uranium as their fuel. Recently plutonium, specifically $^{239}_{94}Pu$, has become favored as a fuel because it can be "bred" (produced) cheaply from ongoing uranium nuclear reactors.

Plutonium is a controversial fuel. If it is produced too abundantly, there is a danger that some nations and/or terrorists could use it for weapons rather than electric energy. It also has a half-life of 24,000 years, which means that, should too much of it be produced artificially, nations will have problems in storing it before it's used as a fuel. U-235 is produced in just enough quantity to be used immediately as a fuel. Pu-239 may become too abundant . . . a very real danger.

Wrong Choices Explained:

(1), (2), (3) These can be eliminated as fission fuels because they are lightweight atoms. Fission requires large atoms such as U-235 and Pu-239 as fuel.

111. 1 In all such nuclear transformations two key rules must be obeyed:

a. The total mass numbers of reactants and products must be equal.
b. The total electric charge numbers must also be equal on both sides of the equation.

$$27 + 4 = 30 + 1$$

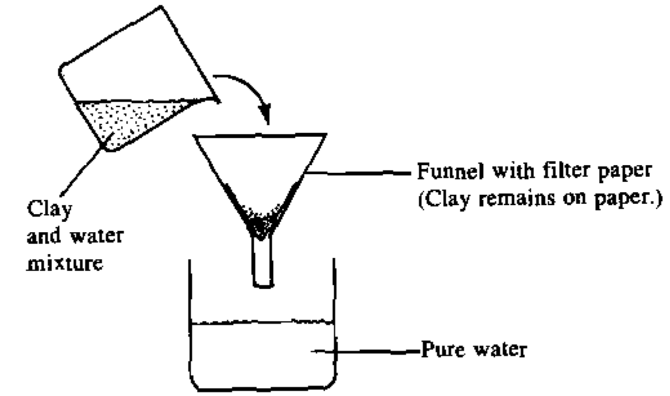
$$^{27}_{13}Al + ^{4}_{2}He \longrightarrow ^{30}_{15}P + ^{1}_{0}X \quad \begin{matrix}\text{Mass numbers}\\ \text{Electric charges}\end{matrix}$$

$$13 + 2 = 15 + 0$$

The unknown particle is $^{1}_{0}X$, which is a neutron.

Group 12—**Laboratory Activities**

112. **4** Filtering separates insoluble substances in a water mixture.

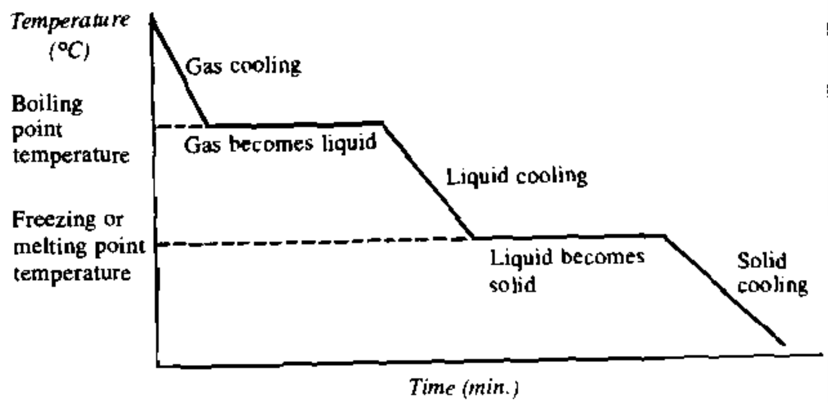


Wrong Choices Explained:

(1) Precipitates are formed by chemical reactions. Sometimes precipitations are made to settle faster by centrifuging. This involves spinning the solution at high speed in a centrifuge. The precipitate settles to the bottom of the tube used.

(2), (3) Boiling is a quick way to separate the water and dissolved particles from one another in a solution. Technically, this is called distillation, or "evaporation to dryness."

113. **2** As time goes by, the temperature of this material will decrease, but not in the systematic, linear manner of choice (3). There will be stretches of time where the temperature will not decrease *even though heat is being removed*. These will occur during phase changes.



114. **3**

$$\% \text{ error} = \frac{\text{observed value} - \text{accepted value}}{\text{accepted value}} \times 100$$

$$\% \text{ error} = \frac{55.2°C - 50.1°C}{50.1°C} \times 100 = \frac{(5.1°C)}{(50.1°C)} \times 100$$

$$\% \text{ error} = 10.2\%$$

115. **4**

$$(\text{volume of acid})(\text{strength of acid}) = (\text{volume of base})(\text{strength of base})$$
$$(20.0 \text{ ml})(x) = (40.0 \text{ ml})(0.20 \text{ M})$$
$$x = \frac{(40.0 \text{ ml})(0.20 \text{ M})}{20.0 \text{ ml}}$$
$$x = 0.04 \text{ M HCl}$$

116. **1** In addition, the sum cannot be more precise than the decimal place in the least precise of the numbers.

0.027 g		(precise to thousandths place)
+ 0.0023 g		(precise to ten-thousandths place)
0.0293 g		
0.029 g		(now precise to thousandths place also)

The digit "3" in 0.0293 g was dropped, and the number rounded off to 0.029 g.

Illustrative examples of rules for significant digits in answers.

Addition and subtraction—no digits allowed beyond the decimal place of the least precise of the numbers.

Multiplication and division—same number of significant digits as are in number with least number of such digits.

0.027 g + 0.0023 g 0.0293 g	0.027 g − 0.0023 g 0.0247 g	0.027 cm × 0.002 3 cm 0.0000621 cm²	0.027 g / 0.0023 ml = 11.73913043 g/ml
0.029 g (to thousandths place only)	**0.025 g** (rounded up to thousandths place only)	**0.000062 cm²** (two significant digits only)	**12 g/ml** (two significant digits only)