1. (D) Alkali metals react the most vigorously with water.

2. (B) HI and H₂SO₄ are strong acids, H₂SO₃ is weak.

3. (B) CuS is insoluble, while NH₄Cl is.

4. (A) Metallic character increases as one moves towards the bottom left of the periodic table.

5. (B) In this procedure, the molar mass will be determined by the density of the vapour as calculated by the ideal gas law. From the ideal gas law, we know that \( M = \frac{mRT}{PV} \). Clearly we need the volume, pressure, and temperature. But the mass in this equation is only the mass of the gas that we have, i.e. the mass that is remaining in vapor form after the rest has evaporated. This was already found with the pre-weighed beaker (simply by subtracting the mass). The mass of the liquid at the beginning is irrelevant.

6. (C) Copper and nitric acid react to form NO₂ gas, while HCl reacts with calcium carbonate to form CO₂ gas.

7. (A) Aluminum forms the oxide Al₂O₃ on its outer surface.

8. (A) NaOH absorbs water from the atmosphere and can form hydrates such as NaOH · H₂O, which can change its molar mass (as a solid) or its concentration (in solution). Having a low molar mass or ionizing in water does not have any impact on titration accuracy. NaOH is strong and reacts quite rapidly with other acids.

9. (C) Let \( x \) be the number of chlorines bonded to one atom of Rhenium. With our given data, we have that \( \frac{186.21}{186.21+35.45x} = 0.636 \implies x = 3.00 \). Hence the formula is ReCl₃.

10. (A) Since AgCl and Cl⁻ exist in a 1:1 stoichiometry, the number of moles Cl⁻ originally in the solution is equal to the number of moles of AgCl precipitated. There are \( \frac{1}{(107.87+35.45)} \) g mol⁻¹ = 0.00698 mol AgCl, so there are 0.00698 · 35.45 = 0.247 grams of chlorine.

11. (D) The equation can be balanced as

   \[
   5\text{Sn}^{2+} + 2\text{MnO}_4^- + 16\text{H}^+ \longrightarrow 5\text{Sn}^{4+} + 2\text{Mn}^{2+} + 8\text{H}_2\text{O(l)}
   \]

   So the ratio is 5/2.
12. (D) There is \( \frac{6.4}{12 + 4 \cdot 1} = 0.4 \) mol methane, so there are 400 mmol methane.

13. (B) The total amount of hydroxide used is \((23.7 \text{ mL})(1.47 \text{ M}) = 34.84 \text{ mmol.}\) But hydroxide reacts with sulfuric acid in a 2 : 1 ratio (since sulfuric acid is diprotic) so there is only \( \frac{34.84}{2} = 17.42 \text{ mmol sulfuric acid.}\) The concentration is thus \( \frac{17.42 \text{ mmol}}{1.5 \text{ mL}} = 11.6 \text{ M.}\)

14. (D) There are \( \frac{2 \text{ g}}{118.71 + 16 \cdot 2} = 0.0133 \text{ mol SnO}_2, \) so \( 2 \cdot 0.0133 = 0.0266 \text{ mol of H}_2 \) is required. At STP each mol of gas takes up 22.4 L, so \( 0.0266 \cdot 22.4 = 0.595 \) liters of hydrogen gas are required.

15. (C) We want to relate the concepts conceptually, since the mathematical variation is somewhat tedious and messy (though it can be done with the Clausius-Clapeyron equation). We proceed sequentially. For set A, high vapour pressures are indicative of weak intermolecular forces, yet a high \( \Delta H \) implies strong intermolecular forces, so A is wrong. Likewise, in set B, a high \( \Delta H \) implies strong intermolecular forces, but a low boiling point implies weak intermolecular forces, so B is also wrong. For set C, both options imply strong intermolecular forces. Finally, for set D, boiling point and rate of evaporation have no correlation (vapour pressure is about equilibrium, not kinetics).

16. (B) The volume of a gas is related to the number of moles, not the molar mass. All the other properties change based on molar mass.

17. (A) By the ideal gas law, we have that

\[
n = \frac{PV}{RT} = \frac{\left(\frac{600}{760} \text{ atm}\right)(0.1 \text{ L})}{(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(287.15 \text{ K})} = 0.00335 \text{ mol}
\]

So the molar mass of the gas is \( \frac{0.239 \text{ g}}{0.00335 \text{ mol}} = 71.4 \frac{\text{ g}}{\text{ mol}}. \) Of the choices given, Cl\(_2\) has the closest molar mass.

18. (D) Only temperature affects vapor pressure (as per Clausius-Clapyeron equation).

19. (C) Any point on the liquid-solid or liquid-gas phase boundaries, or directly in the liquid phase, can have liquid present.

20. (C) Diethylamine can form strong intermolecular attractions due to hydrogen bonding, increasing its solubility greatly.
21. (D) The high melting point means that it is not molecular; the non-conductivity means that it is not ionic (since ionic compounds conduct when dissolved) nor metallic.

22. (C) Endothermic reactions have a positive $\Delta H$; the formation of ozone has the highest $\Delta H$.

23. (B) $q = C \cdot \Delta T \implies \Delta T = \frac{q}{C}$. Thus the highest temperature change occurs with the lowest heat capacity, which gold has.

24. (C) We are given that the reaction

$$2 \text{Li} + 0.5 \text{O}_2 \rightarrow \text{Li}_2\text{O}$$

has $\Delta H = -598.8 \text{kJ}$, so $\frac{598.8}{2} = 299.4 \text{kJ}$ is released per mole of lithium. So we need $\frac{150}{299.4} = 0.501$ mol lithium to release 150 kJ, which has mass $0.501 \cdot 6.94 = 3.48$ grams.

25. (D) Gases have much higher standard entropy as compared to solids/liquids. NaCl has higher entropy than Na because there is more mass as well as more arrangements in the crystal structure of NaCl.

26. (B) The Gibbs Free Energy is $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ = -60 + 0.2T$. For the reaction to be spontaneous, we need $\Delta G^\circ < 0 \implies 0.2T < 60 \implies T < 300$.

27. (A) NO$_2$ and O$_2$ are related in a 2 : 1 ratio, so the rate of O$_2$ formation is $0.0036 \cdot \frac{1}{2} = 0.0018$.

28. (D) As can be seen from a reaction diagram, the activation energy of the reverse reaction is equal to the activation energy of the forward reaction minus the enthalpy change of the forward reaction. In this case, the activation energy would be $120 - (-85) = 205 \text{kJ/mol}$.

29. (C) By comparing the first and second trials we see that the rate doubles as the concentration of ethyl iodide doubles, so the reaction is first order in ethyl iodide. By comparing the second and third trials, we know that the rate doubles when the concentration of hydroiodic acid is doubled, so the reaction is also first order in hydroiodic acid.

30. (A) Catalysts are not used up during a reaction, so they appear on both sides of an equation.
31. (B) The integrated rate law for first order reactions is $[A] = [A]_0 e^{-kt}$, so $\ln[A] = \ln[A]_0 - kt$, which is linear in time.

32. (B) As the pressure increases, the volume decreases, so by Le Chatelier’s principle, the equilibrium shifts to favor the side with less gas, which would reduce [HI]. Since the enthalpy change of the reaction is positive, by the Vant Hoff equation, $K$ would increase, which means more HI would be formed as well.

33. (D) The equilibrium expression is obtained by taking the ratio of the (concentrations of the gaseous/aqueous components of the) products and reactants, each raised to their stoichiometric coefficients. $I_2$ here is a solid and is not taken into account.

34. (B) We know that $[H^+] = [C_6H_5COO^-] = 0.045 - [C_6H_5COOH]$ from stoichiometry, so by definition $K_a = \frac{[H^+][C_6H_5COO^-]}{[C_6H_5COOH]} = \frac{1.7 \cdot 10^{-3}}{0.045 - 1.7 \cdot 10^{-3}} \approx 6.7 \cdot 10^{-5}$. (Note: $6.4 \cdot 10^{-5}$ would have been obtained if one used $[C_6H_5COOH]=0.45 \text{ M}$, which is what the authors most likely intended (though it is wrong).)

35. (D) Since the equilibrium constant is less than 1, this means that the reaction is not favoured. This means that forming $C_6H_5O^-$ is unfavourable, so it is a stronger base than CN$^-$. The other two relevant compounds are acids.

36. (A) The reaction is the sum of the following two reactions:

$$\text{OCl}^- + H^+ \rightleftharpoons \text{HOCl}$$

$$\text{H}_2\text{O} \rightleftharpoons H^+ + \text{OH}^-$$

The first reaction has $K = \frac{1}{K_a} = 3.3 \cdot 10^7$, while the second is the autoprotolysis of water and has $K = 10^{-14}$. The equilibrium constant of the sum of the reactions is $3.3 \cdot 10^7 \cdot 10^{-14} = 3.3 \cdot 10^{-7}$.

37. (A) Acidifying the solution will lead to the formation of HF which is a weak acid and does not disassociate, causing F$^-$ to be removed from the system and increasing the solubility of PbF$_2$. On the other hand, acidifying the solution of any other halogen will not result in the formation of the acid, since those acids are strong and disassociate, which means that the solubility of PbX$_2$ ($X = Cl, Br, I$) does not change appreciably.

38. (D) We can find the oxidation numbers of nitrogen in each reaction; only in reaction D does it increase (from +3 to +5). For reaction A, it decreases from 0 to −3, and the oxidation number does not change in reactions B and C (respectively remaining at +4 and +5).
39. (A) Water’s reduction potential is much higher than that of potassium, so it is preferentially reduced at the cathode. The water is reduced to hydroxide anions and hydrogen gas, and the hydroxide turns the cathode pink due to the phenolphthalein indicator. On the cathode, iodine is formed, which has a yellow color.

40. (A) Oxidation occurs at the anode. Since aluminum has a lower reduction potential, it is easier for aluminum to undergo oxidation, so the reaction is $\text{Al} \rightarrow \text{Al}^{3+} + 3e^-$

41. (B) $E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}} = -1.18 - (-1.66) = 0.48 \text{ V}$.

42. (A) By the Nernst equation, $E = E^o - \frac{RT}{nF} \ln \left( \frac{[\text{Cu}^{2+}]}{[\text{Ni}^{2+}]} \right) = E^o + \frac{RT}{nF} \ln \left( \frac{[\text{Cu}^{2+}]}{[\text{Ni}^{2+}]} \right)$, so increasing the concentration of copper ions increases the voltage. Changing the electrode size does not change the voltage.

43. (B) The conversion is from $P$ to $P^{3-}$. $P$ has 3 unpaired electrons (and 15 total), the addition of 3 more electrons completes the electron pairs in phosphorous and brings it up to 18 total electrons.

44. (C) $\text{K}^+$, $\text{Cl}^-$, $\text{S}^{2-}$ all have 18 electrons.

45. (D) The number of neutrons is the mass number minus the number of protons. Both atoms have the same number of neutrons (30) in set D.

46. (D) By periodic trends, those elements closer to the bottom left of the periodic table have more metallic character. Only in set D does the atoms move closer to the bottom left.

47. (B) The electron configuration of iron is $[\text{Ar}]4s^23d^6$. Since there are 5 d-orbitals, one of the d-orbitals must share a pair of electrons; the rest have unpaired electrons.

48. (B) The energy gap is given by

$$\Delta E = -Z_{el}^2 \hbar c R_\infty \left( \frac{1}{(n+1)^2} - \frac{1}{n^2} \right) = Z_{el}^2 \hbar c R_\infty \left( \frac{2n+1}{n^2(n+1)^2} \right)$$

which clearly decreases as $n$ increases.

49. (A) Elements close together have similar electronegativities; oxygen and fluorine are the closest together.
50. (C) Phosphorous has 5 valence electrons and can form the stable compound \( \text{PH}_3 \) which has similar structure to ammonia. \( \text{CH}_3 \) and \( \text{SH}_3 \) would be reactive free radicals; \( \text{ClH}_3 \) could theoretically exist but would be highly unstable (Compare to the stable \( \text{ClF}_3 \) which has a similar structure, but the fluorines can stabilize the positive charge on the Cl).

51. (A) \( \text{NO}_3^- \) and \( \text{CO}_3^{2-} \) have the same number of valence electrons and the same molecular structure.

52. (D) All the compounds contain both covalent and ionic bonds.

53. (B) Each of the compounds has steric number 4 and will have similar bond angles according to VSEPR theory, but electron pairs are more repulsive and hence \( \text{H}_2\text{F}^+ \) will have the smallest bond angle, followed by \( \text{H}_3\text{O}^+ \) and \( \text{NH}_4^+ \). (Note: While this explanation is most likely the intended one, this is somewhat wrong; see ”Bent’s Rules” for a full explanation).

54. (C) Drawing the structure of the nitrite ion, we see that it has two resonance structures, one where the bond is a single bond, and the other in where the bond is a double bond. Thus the actual bond length would be somewhere between a single and double bond.

55. (D) Carboxyl groups are those with carboxylic acids; only propanoic acid is a carboxylic acid.

56. (A) Alkanes are saturated hydrocarbons; alkenes, alkynes, and aromatics each have degrees of unsaturation.

57. (C) There are four: butan-1-ol, butan-2-ol, 2-methylpropan-1-ol, and 2-methylpropan-2-ol. (Note: if counting optical isomers, there are five compounds, since butan-2-ol can exist as a pair of enantiomers).

58. (A) ortho compounds are 1,2 disubstituted; meta compounds are 1,3 disubstituted, and para compounds are 1,4 disubstituted.

59. (B) cis-trans isomerism occurs in double bonds where there are non-equivalent substituents, as in this case, or in optical isomers.

60. (C) Drawing out a few units of the polymer we see that it is made of distinct \( \text{CCl}_2-\text{CH}_2 \) units, so we can conclude that compound C is used for its polymerization (specifically addition polymerization).