

This set of review questions is designed to help you in preparation of the final exam, which will include the following two major topics: (1) **Energetics in Chemical Reactions** (Gibbs' free energy, entropy, standard molar free energy, phase transitions, ΔG , ΔG° and the equilibrium constant); and (2) **Electrochemistry** (oxidation and reduction, half-cells, standard reduction potential, cell potential, Nernst equation and electrolysis). These questions are intended to provide you with practice at solving quantitative problems. It is in your best interests to work through **all** these questions independently *before the exam*. Review Chapter 13 "Acid Rain" and Chapter 15 "Stratospheric Ozone Depletion".

You are assumed to have access to tables with Thermodynamic data and Standard Reduction Potentials (SRP).

PART A Energetics in Chemical Reactions

- For each of the reactions listed below:
 - $4 \text{HCl(g)} + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O(g)} + 2 \text{Cl}_2(\text{g})$
 - $\text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2 \text{HCl(g)}$
 - $\text{C(graphite)} \rightarrow \text{C(diamond)}$
 - Use tabulated Thermodynamic data to calculate ΔG° at 298 K.
 - Use tabulated Thermodynamic data to calculate ΔH° and ΔS° at 298 K and verify that $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$.
 - Will the reaction occur spontaneously at 298 K with all reactants and products in their standard states? If so, which factor, enthalpy or entropy, provides the principal driving force?
 - Where the reaction will not occur spontaneously at 298 K, determine whether it would become spontaneous at some higher or lower temperature. For each reaction, determine the range of temperatures, if any, over which it will be spontaneous.
- Over what range of temperatures is each of the following reactions spontaneous?
 - a reaction with $\Delta H^\circ = +53.4 \text{ kJ}$ and $\Delta S^\circ = +112.4 \text{ J K}^{-1}$
 - a reaction with $\Delta H^\circ = -29.4 \text{ kJ}$ and $\Delta S^\circ = -91.2 \text{ J K}^{-1}$
- You need a chemical reaction that proceeds spontaneously at low temperatures but proceeds in the reverse direction at elevated temperatures. What are the signs of ΔH° and ΔS° for such a reaction?
- Assume that ΔH° and ΔS° do not change significantly with temperature. Calculate the vapour pressure of each of the following compounds at the stated conditions.
 - $\text{CS}_2(\ell)$ at 5°C
 - $\text{CCl}_4(\ell)$ at 29°C
 - $\text{CH}_3\text{CHO}(\ell)$ at 45°C
- Given the following data for ethanol:

$$\text{CH}_3\text{CH}_2\text{OH(s)} \rightarrow \text{CH}_3\text{CH}_2\text{OH}(\ell)$$

$$\Delta H^\circ_{\text{rxn}} = 4.60 \text{ kJ mol}^{-1}$$

$$\Delta S^\circ_{\text{rxn}} = +29.7 \text{ J K}^{-1} \text{ mol}^{-1}$$
 Estimate the freezing point for $\text{CH}_3\text{CH}_2\text{OH}$.
- Use tabulated Thermodynamic data to estimate K_c for the reaction at 100°C .

$$\text{H}_2\text{O}(\ell) \rightarrow \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$$
 Note: $S^\circ(\text{OH}^-, \text{aq}) = -10.5 \text{ J K}^{-1} \text{ mol}^{-1}$
- At 298 K, the reaction $\text{CO}_2(\text{g}) \rightleftharpoons \text{CO}_2(\text{aq})$ has:

$$\Delta H^\circ = -23.0 \text{ kJ mol}^{-1}$$

$$\Delta G^\circ_f[\text{CO}_2(\text{g})] = -394.0 \text{ kJ mol}^{-1}$$

$$\Delta G^\circ_f[\text{CO}_2(\text{aq})] = -386.0 \text{ kJ mol}^{-1}$$
 - Calculate Henry's Law Constant, K_{H} , for $\text{CO}_2(\text{g})$ at 298 K.
 - Calculate ΔS° for this reaction at 298 K.
- Given the thermodynamic data:

$$\Delta G^\circ_f[\text{H}_2\text{O}(\ell)] = -237.2 \text{ kJ mol}^{-1}$$

$$\Delta H^\circ_f[\text{H}_2\text{O}(\ell)] = -285.8 \text{ kJ mol}^{-1}$$

$$\Delta S^\circ[\text{H}_2\text{O}(\ell)] = 69.9 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$\Delta S^\circ[\text{H}_2(\text{g})] = 130.6 \text{ J K}^{-1} \text{ mol}^{-1}$$

$$\Delta S^\circ[\text{O}_2(\text{g})] = 205 \text{ J K}^{-1} \text{ mol}^{-1}$$
 for the reaction:

$$\text{H}_2(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\ell)$$
 - Calculate ΔG° and the equilibrium constant, K_p , at 298 K.
 - Calculate ΔG° and the equilibrium constant, K_p , at 2000 K.
 - Calculate ΔG at 2000 K when $\text{H}_2(\text{g})$ and $\text{O}_2(\text{g})$ are at 10.0 atm pressure each. In which direction will the reaction proceed under these conditions?

9. Copper sulfate (CuSO_4) forms two distinctly different coloured hydrates which can be interconverted as follows:
 $\text{CuSO}_4 \cdot \text{H}_2\text{O}(\text{s, white}) + 4 \text{H}_2\text{O}(\text{g}) \rightarrow \text{CuSO}_4 \cdot 5\text{H}_2\text{O}(\text{s, blue})$
- | At 298 K | $\Delta H_f^\circ (\text{kJ mol}^{-1})$ | $S^\circ (\text{J K}^{-1} \text{mol}^{-1})$ |
|---|---|---|
| $\text{CuSO}_4 \cdot \text{H}_2\text{O}(\text{s, white})$ | -1084 | 150 |
| $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}(\text{s, blue})$ | -2278 | 305 |
| $\text{H}_2\text{O}(\text{g})$ | -242 | 189 |
- (a) Calculate ΔG for the reaction at 15°C if the pressure of water vapour in the atmosphere is 0.025 atm.
 (b) Is the reaction spontaneous as written under the condition given in Part (a)?
10. With access to Thermodynamic data, calculate the values of solubility product constant, K_{sp} , and the solubility of $\text{AgCl}(\text{s})$ at 25°C .
11. For the reaction: $2 \text{H}_2\text{O}(\ell) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$
 $K_{\text{w}} = 1.0 \times 10^{-14} \text{ M}^2$ at 25°C
 $\Delta H = 55.8 \text{ kJ mol}^{-1}$
- (a) What is K_{w} at 100°C ?
 (b) What is pH of boiling water?
12. The equilibrium constant for the formation of phosgene is measured at two different temperatures.
 $\text{CO}(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow \text{COCl}_2(\text{g})$
- (a) At 506°C , $K_{\text{eq}} = 1.3$; and at 530°C , $K_{\text{eq}} = 0.78$. Calculate ΔH° and ΔS° for this reaction.
 (b) Under standard-state conditions, over what temperature range is the reaction spontaneous?

MULTIPLE CHOICE QUESTIONS

1. The vapour pressure of water at 25°C is 23.76 torr. ΔG for the reaction below at 25°C is:
 $\text{H}_2\text{O}(\ell) \rightarrow \text{H}_2\text{O}(\text{g}, 23.76 \text{ torr})$
- (a) 0 (b) 8.59 kJ (c) 7.85 kJ (d) -7.85 kJ (e) 20.7 kJ
2. For the reaction: $\text{HCO}_2\text{H}(\ell) \rightarrow \text{HCO}_2\text{H}(\text{g})$ at 298 K
 $\Delta H^\circ = 46.60 \text{ kJ}$, $\Delta G^\circ = 10.3 \text{ kJ}$, and $\Delta S^\circ = 122 \text{ J K}^{-1}$
 The normal boiling point of formic acid, $\text{HCO}_2\text{H}(\ell)$, is therefore about:
- (a) 382°C (b) 109°C (c) 84°C (d) 84 K (e) 262°C
3. The value of K_{p} for the reaction:
 $\text{Hg}(\ell) \rightarrow \text{Hg}(\text{g})$
 at 100°C is 36.38 Pa. ΔG° of vaporization of $\text{Hg}(\ell)$ at 100°C is therefore:
- (a) 6.6 kJ (b) 11.1 kJ (c) 15.2 kJ (d) 24.6 kJ (e) 56.6 kJ
4. For which of the following reactions would you predict a positive value for $\Delta S_{\text{rxn}}^\circ$?
- (a) $\text{CO}_2(\text{g}) \rightarrow \text{CO}_2(\text{s})$
 (b) $\text{NH}_4^+(\text{aq}) + \text{NO}_3^-(\text{aq}) \rightarrow \text{NH}_4\text{NO}_3(\text{s})$
 (c) $\text{Zn}^{2+}(\text{aq}) + \text{H}_2(\text{g}) \rightarrow \text{Zn}(\text{s}) + 2 \text{H}^+(\text{aq})$
 (d) $\text{C}_6\text{H}_6(\ell) + 9/2 \text{O}_2(\text{g}) \rightarrow 6 \text{CO}_2(\text{g}) + 3 \text{H}_2\text{O}(\ell)$
 (e) $5 \text{CO}_2(\text{g}) + 6 \text{H}_2\text{O}(\text{g}) \rightarrow \text{C}_5\text{H}_{12}(\text{g}) + 8 \text{O}_2(\text{g})$
5. The reaction: $2 \text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow 2 \text{CO}(\text{g})$ is exothermic. The reaction is:
- (a) spontaneous at all temperatures (b) spontaneous at low temperatures
 (c) spontaneous at high temperatures (d) nonspontaneous at all temperatures
 (e) nonspontaneous at low temperatures
6. The standard free energy change for the dissociation of silver chromate is $+66.2 \text{ kJ mol}^{-1}$ at 25°C .
 $\text{Ag}_2\text{CrO}_4(\text{s}) \rightleftharpoons 2 \text{Ag}^+(\text{aq}) + \text{CrO}_4^{2-}(\text{aq})$
 The concentration of Ag^+ ions that can exist in equilibrium with $2.0 \times 10^{-3} \text{ M}$ of CrO_4^{2-} ions is:
- (a) $3.5 \times 10^{-5} \text{ M}$ (b) $6.2 \times 10^{-7} \text{ M}$
 (c) $1.2 \times 10^{-9} \text{ M}$ (d) $6.3 \times 10^{-10} \text{ M}$
 (e) $2.5 \times 10^{-12} \text{ M}$
7. At 120°C , the value of K_{p} is 0.495 atm^{-2} for the reaction:
 $2 \text{H}_2(\text{g}) + \text{CO}(\text{g}) \rightleftharpoons \text{CH}_3\text{OH}(\text{g})$
 If the ΔG° at 240°C is $+42.2 \text{ kJ mol}^{-1}$, then the standard enthalpy of the reaction is:
- (a) 0 (b) $-18.4 \text{ kJ mol}^{-1}$
 (c) $-42.2 \text{ kJ mol}^{-1}$ (d) -128 kJ mol^{-1}
 (e) cannot be determined
8. A plot of $\ln(\text{vapour pressure})$ versus $1/T$ for: $\text{C}_2\text{H}_5\text{OH}(\ell) \rightleftharpoons \text{C}_2\text{H}_5\text{OH}(\text{g})$
 gave a straight line of slope $-5.09 \times 10^3 \text{ K}$. This result indicates that:
- (a) $\Delta H_{\text{vap}} = 42.3 \text{ kJ mol}^{-1}$ (b) $\Delta G_{\text{vap}} = 42.3 \text{ kJ mol}^{-1}$
 (c) $\Delta H_{\text{vap}} = 97.4 \text{ kJ mol}^{-1}$ (d) $\Delta H_{\text{vap}} = -18.4 \text{ kJ mol}^{-1}$
 (e) $\Delta H_{\text{vap}} = -42.3 \text{ kJ mol}^{-1}$

PART B Electrochemistry

- For the cell: $\text{Pt(s)} \mid \text{Fe}^{2+}(\text{aq}), \text{Fe}^{3+}(\text{aq}) \parallel \text{Hg}^{2+}(\text{aq}) \mid \text{Hg}(\ell)$
 - Write the anode half-reaction.
 - Write the cathode half-reaction.
 - Calculate the standard cell potential for the cell.
- Calculate the potential for the half-cell: $\parallel \text{Cu}^{2+}(\text{aq}, 0.0356 \text{ M}) \mid \text{Cu(s)}$
- Calculate the potential of the cell: $\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}, 0.322 \text{ M}) \parallel \text{Cu}^{2+}(\text{aq}, 0.0665 \text{ M}) \mid \text{Cu(s)}$
- Calculate the concentration of $\text{Cl}^{-}(\text{aq})$ in the cathode compartment of the following cell which has a potential of 1.80 V at 298 K.
 $\text{Ni(s)} \mid \text{NiCl}_2(\text{aq}, 0.0200 \text{ M}) \parallel \text{CaCl}_2(\text{aq}) \mid \text{Cl}_2(\text{g}, 1.97 \text{ atm}) \mid \text{Pt(s)}$
- For the following reaction:
 $\text{AgCl(s)} + \text{Fe}^{2+}(\text{aq}) \rightarrow \text{Ag(s)} + \text{Cl}^{-}(\text{aq}) + \text{Fe}^{3+}(\text{aq})$
 - Write a "shorthand" cell diagram for the reaction at an inert platinum anode.
 - Calculate \mathcal{E}° for the reaction.
 - Calculate ΔG° for the reaction.
 - Calculate the equilibrium constant for the reaction.
- The cell potential of the cell: $\text{Cd(s)} \mid \text{Cd}^{2+}(\text{aq}, 0.0250 \text{ M}) \parallel \text{Pb}^{2+}(\text{aq}, 0.150 \text{ M}) \mid \text{Pb(s)}$ is 0.293 V at 298 K. With **NO** access to SRP Table, calculate ΔG and ΔG° for the cell reaction at 298 K.
- Calculate $[\text{Zn}^{2+}]$ in the anode compartment, given that the cell below has $\mathcal{E} = 0.015$ volts.
 $\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}, ? \text{ M}) \parallel \text{Zn}^{2+}(\text{aq}, 0.050 \text{ M}) \mid \text{Zn(s)}$
- Calculate the reduction potential of the hydrogen electrode at 25°C
 $2 \text{H}^{+}(\text{aq}) + 2 \text{e}^{-} \rightarrow \text{H}_2(\text{g})$
 when the partial pressure of $\text{H}_2(\text{g})$ is 2.5 atm and the $\text{pH} = 6.00$.
- A possible reaction for a fuel cell is: $\text{C}_2\text{H}_6(\text{g}) + 7/2 \text{O}_2(\text{g}) \rightarrow 2 \text{CO}_2(\text{g}) + 3 \text{H}_2\text{O}(\ell)$
 Calculate ΔG° and \mathcal{E}° for this potential fuel cell. (Assume standard conditions.)
- From SRP Table, choose reagents that could reduce $\text{Cu}^{2+}(\text{aq})$ to Cu(s) , but not $\text{Al}^{3+}(\text{aq})$ to Al(s) .
- The Cu^{2+} ion concentration in a solution at 25°C is determined as follows:
 A silver electrode is dipped into 0.50 M AgNO_3 , connected by a salt bridge to a second half-cell containing a Cu electrode dipping into the solution whose Cu^{2+} concentration is to be determined. As the cell works, concentration of $\text{Cu}^{2+}(\text{aq})$ is increasing. The measured initial cell potential is 0.62 V. Calculate the initial value of $[\text{Cu}^{2+}]$.
- Using SRP Table, calculate the solubility product constant, K_{sp} , of AgBr(s) .
 $\text{AgBr(s)} + \text{e}^{-} \rightarrow \text{Ag(s)} + \text{Br}^{-}(\text{aq}) \quad \mathcal{E}^{\circ} = +0.071 \text{ V}$
- Calculate the mass of Cu metal produced during the passage of 2.50 A for 50.0 minutes through aqueous copper sulfate. O_2 gas is also produced in this electrolysis. Write the equations for both half-reactions. What volume of gas is produced at STP?
- Write the cathode and the anode reactions that takes place during the electrolysis of aqueous solutions of:
 - NaBr
 - AgF
 - NaF
- You decide to electrolyse an aqueous solution of NaOH. What products will you get in each electrode? Write the overall redox reaction.

MULTIPLE CHOICE QUESTIONS

- The oxidation numbers of chlorine in $\text{Mg}(\text{ClO})_2$ and ClO_3^{-} are, respectively:
 - 1, +6
 - +2, +5
 - 1, -1
 - +1, +7
 - +1, +5
- Under standard condition, the following redox reactions are observed to occur in aqueous solution.
 $\text{A}^{+} + \text{B} \rightarrow \text{A} + \text{B}^{+}$
 $\text{A}^{+} + \text{C} \rightarrow \text{no reaction}$
 $2 \text{B}^{+} + \text{D} \rightarrow 2 \text{B} + \text{D}^{2+}$
 The decreasing order of reactivity as reducing agent is:
 - $\text{C} > \text{A} > \text{D} > \text{B}$
 - $\text{D} > \text{B} > \text{A} > \text{C}$
 - $\text{C} > \text{A} > \text{B} > \text{D}$
 - $\text{B} > \text{D} > \text{A} > \text{C}$
 - cannot be determined
- When an acidic solution of potassium dichromate is mixed with a sodium chloride solution, Cr^{3+} ions and chlorine gas are produced. The oxidizing agent is:
 - $\text{H}_2\text{O}(\ell)$
 - $\text{Cr}^{3+}(\text{aq})$
 - $\text{Cl}_2(\text{g})$
 - $\text{Cr}_2\text{O}_7^{2-}(\text{aq})$
 - $\text{Cl}^{-}(\text{aq})$

4. For the balanced redox reaction:

$$5 \text{H}_2\text{S}(\text{aq}) + 2 \text{MnO}_4^-(\text{aq}) + 6 \text{H}^+(\text{aq}) \rightarrow 2 \text{Mn}^{2+}(\text{aq}) + 5 \text{S}(\text{s}) + 8 \text{H}_2\text{O}(\ell)$$
 the number of electrons transferred is:
 (a) 3 (b) 4 (c) 6 (d) 8 (e) 10
5. For the reaction: $\text{C}_6\text{H}_{12}\text{O}_6(\text{s}) + 6 \text{O}_2(\text{g}) \rightarrow 6 \text{CO}_2(\text{g}) + 6 \text{H}_2\text{O}(\ell)$
 the number of electrons transferred is:
 (a) 24 (b) 4 (c) 6 (d) 12 (e) 8
6. The standard potential, \mathcal{E}° , for the cell: $\text{Sc}(\text{s}) \mid \text{Sc}^+(\text{aq}) \parallel \text{Cl}^-(\text{aq}) \mid \text{AgCl}(\text{s}) \mid \text{Ag}(\text{s})$ is 0.56 V.
 Given that: $\text{AgCl}(\text{s}) + \text{e}^- \rightarrow \text{Ag}(\text{s}) + \text{Cl}^-(\text{aq})$ $\mathcal{E}^\circ = +0.22 \text{ V}$
 The standard reduction potential of $\text{Sc}^+(\text{aq})$ is:
 (a) 0.34 V (b) -0.34 V (c) 0.78 V (d) -0.78 V (e) -0.56 V
7. Consider the following redox reaction:

$$2 \text{Ag}^+(\text{aq}) + \text{H}_2(\text{g}) \rightarrow 2 \text{Ag}(\text{s}) + 2 \text{H}^+(\text{aq})$$
 If $[\text{Ag}^+(\text{aq})] = 0.10 \text{ M}$, the pressure of $\text{H}_2(\text{g})$ is at 1 atm and the pH of the solution is 8.00. The cell potential at 25°C is:
 (a) -0.03 V (b) +0.39 V (c) +1.01 V (d) +1.21 V (e) +1.63 V
8. A 1.0 M solution of KI is subject to electrolysis. Which product(s) will be forming at the anode?
 (a) $\text{I}_2(\text{s})$ (b) $\text{K}(\text{s})$ (c) $\text{H}_2(\text{g})$ (d) $\text{O}_2(\text{g})$ (e) $\text{H}_2(\text{g})$ and $\text{O}_2(\text{g})$
9. How long would it take to electrodeposit (plate out) 1.0 g of Ni(s) from a 500 mL of a 0.100 M NiSO_4 solution using a constant current of 250 mA?
 (a) 5.5 min (b) 22 min (c) 55 min (d) 110 min (e) 220 min
10. An electrolysis of 250 mL of a 0.500 M $\text{Fe}(\text{NO}_3)_3$ solution is carried out for 30 min using a constant current of 25.0 A. The mass of Fe(s) deposition at the cathode is:
 (a) 1.88 g (b) 2.32 g (c) 6.98 g (d) 8.63 g (e) 13.0 g

ANSWERS

PART A Energetics in Chemical Reactions

1. (i) spontaneous up to 888 K (ii) spontaneous at all temp (iii) nonspontaneous at any temp
 2. (a) $T > 475 \text{ K}$ (b) $T < 322 \text{ K}$ 3. ΔH° and ΔS° both negative
 4. (a) 0.207 atm (b) 0.177 atm (c) 2.58 atm
 5. 155 K (or -118°C) 6. $9.34 \times 10^{-13} \text{ M}^2$
 7. (a) $3.96 \times 10^{-2} \text{ M atm}^{-1}$ (b) $-104 \text{ J K}^{-1} \text{ mol}^{-1}$
 8. (a) -237.2 kJ ; $3.79 \times 10^{41} \text{ atm}^{-3/2}$ (b) 40.6 kJ ; $8.70 \times 10^{-2} \text{ atm}^{-3/2}$ (c) -16.8 kJ ; Left to Right
 9. (a) $-17.6 \text{ kJ mol}^{-1}$ (b) yes, blue form more stable 10. $1.75 \times 10^{-10} \text{ M}^2$; $1.32 \times 10^{-5} \text{ M}$
 11. (a) $9.3 \times 10^{-13} \text{ M}^2$ (b) 6.02
 12. (a) $\Delta H = -1.1 \times 10^2 \text{ kJ mol}^{-1}$; $\Delta S = -1.4 \times 10^2 \text{ J K}^{-1} \text{ mol}^{-1}$ (b) spontaneous at $T < 791 \text{ K}$

MULTIPLE CHOICE QUESTIONS

1. a 2. b 3. d 4. d 5. a
 6. a 7. d 8. a

PART B Electrochemistry

1. (c) +0.080 V 2. +0.294 V 3. +1.079 V 4. $7.35 \times 10^{-3} \text{ M}$
 5. (a) $\text{Pt}(\text{s}) \mid \text{Fe}^{2+}(\text{aq}), \text{Fe}^{3+}(\text{aq}) \parallel \text{Cl}^-(\text{aq}) \mid \text{AgCl}(\text{s}) \mid \text{Ag}(\text{s})$
 (b) -0.549 V (c) +53.0 kJ mol⁻¹ (d) $5.17 \times 10^{-10} \text{ M}$
 6. $\Delta G = -56.5 \text{ kJ mol}^{-1}$ and $\Delta G^\circ = -52.1 \text{ kJ mol}^{-1}$ 7. $1.6 \times 10^{-2} \text{ M}$ 8. -0.37 V
 9. $-1467.332 \text{ kJ mol}^{-1}$; +1.09 V 11. $1.1 \times 10^{-6} \text{ M}$ 12. $4.81 \times 10^{-13} \text{ M}^2$
 13. Cu mass: 2.47 g; O_2 volume: 0.435 L
 14. (a) cathode: $\text{H}_2(\text{g})$; anode: $\text{Br}_2(\ell)$ (b) cathode: $\text{Ag}(\text{s})$; anode: $\text{O}_2(\text{g})$
 (c) cathode: $\text{H}_2(\text{g})$; anode: $\text{O}_2(\text{g})$
 15. cathode: $\text{H}_2(\text{g})$; anode: $\text{O}_2(\text{g})$; $2 \text{H}_2\text{O}(\ell) \rightarrow 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g})$

MULTIPLE CHOICE QUESTIONS

1. e 2. b 3. d 4. e 5. a
 6. b 7. d 8. a 9. e 10. c