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HSSC-I NATIONAL BOOK FOUNDATION FBISE

CHEMISTRY TEST - CHAPTER 6: ENERGETICS

Name:		Roll No.		Section:

Total Time: 2:30 Hour
Time Allowed: 150 Minutes

Total Marks: 150

PAPER OBJECTIVE
SECTION-A (MARKS 150)

NOTE: Section-A is compulsory. All parts of this section are to be answered on these pages and handed over to the center superintendent. Deleting / overwriting is not allowed. Do not use lead pencil.

Q.1 Fill the relevant bubble for each part. All parts carry one mark.

Questions	A	B	C	D	A B C D
1. The study of energy changes associated with physical and chemical changes is known as:	periodicity	stoichiometry	energetics	kinetics	○○●○
2. The total amount of energy in the universe:	keeps increasing	keeps decreasing	remains unchanged	becomes zero	○○●○
3. In chemical reactions, bond breaking always:	releases energy	consumes energy	produces light only	has no energy change	○●○○
4. In chemical reactions, bond making always:	releases energy	has no energy change	absorbs heat	consumes energy	●○○○
5. Energy is transferred from system to surroundings when the energy released by bond formation is:	greater than the energy consumed by bond breaking	less than the energy consumed by bond breaking	equal to zero	always constant	●○○○
6. The SI unit of heat is the:	calorie	pascal	joule	kelvin	○○●○
7. One calorie is the heat required to raise the temperature of one gram of water from:	25 °C to 35 °C	14.5 °C to 15.5 °C	0 °C to 100 °C	20 °C to 30 °C	○●○○
8. One calorie is equal to:	0.418 joules	41.8 joules	4.18 joules	418 joules	○○●○
9. A reaction that proceeds with the evolution of heat is called:	exothermic	isothermal	endothermic	adiabatic	●○○○
10. A reaction that proceeds with the absorption of heat is called:	catalytic	exothermic	spontaneous only	endothermic	○○○●
11. For an exothermic reaction, ΔH is:	negative	zero	positive	infinite	●○○○
12. For an endothermic reaction, ΔH is:	zero	equal to −1	positive	negative	○○●○
13. A balanced chemical equation that also shows the heat change is called a:	word equation	nuclear equation	thermochemical equation	ionic equation	○○●○
14. The branch of chemistry dealing with heat changes in chemical reactions is:	photochemistry	electrochemistry	thermochemistry	biochemistry	○○●○

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Questions	A	B	C	D	A B C D
15. Burning of fuels is a highly:	non-reaction	exothermic reaction	endothermic reaction	reversible reaction only	○●○○
16. Which of the following is an exothermic reaction?	Decomposition of limestone	Combustion of methane	Melting of ice	Sublimation of dry ice	○●○○
17. Which of the following is an endothermic process?	Combustion of methane	Decomposition of limestone	Condensation of steam	Freezing of water	○●○○
18. Which of the following is an exothermic process?	Melting of ice	Freezing of water	Vaporisation of water	Sublimation of dry ice	○●○○
19. If a reaction is exothermic in the forward direction, in the reverse direction it is:	endothermic with the same magnitude of ΔH	exothermic with double the ΔH	exothermic with the same ΔH	endothermic with half the ΔH	●○○○
20. For C(s) + O_2(g) → CO_2(g), ΔH° = −393.5 kJ. For the reverse reaction, ΔH° is:	+393.5 kJ	−787 kJ	+196.75 kJ	−393.5 kJ	●○○○
21. The heat evolved or absorbed in a reaction when molar quantities are as shown in the equation is called the:	heat of vaporisation	heat of fusion	heat of reaction	heat capacity	○○●○
22. Heat of reaction measured at 25 °C and 1 atm is called:	bond energy	specific heat	molar heat of fusion	standard enthalpy change	○○○●
23. Standard temperature for thermochemical data is:	0 K	298 K	273 K	373 K	○●○○
24. Because most reactions occur at constant pressure, the heat change equals the change in:	volume	mass	enthalpy	density	○○●○
25. Enthalpy change of a reaction is given by:	H(reactants) − H(products)	H(products) − H(reactants)	H(products) × H(reactants)	H(products) + H(reactants)	○●○○
26. Combustion reactions are always:	neither exothermic nor endothermic	endothermic	exothermic	non-spontaneous	○○●○
27. The standard state of a gas is a pressure of:	10 atm	1 atm	0.1 atm	760 atm	○●○○
28. The standard state of a substance in aqueous solution has a concentration of:	10 M	0.1 M	0.01 M	1 M	○○○●
29. The enthalpy change when one mole of a compound is formed from its elements in their standard states is called the standard enthalpy of:	neutralization	combustion	atomization	formation	○○○●
30. The standard enthalpy of formation of an element in its standard state is:	zero	one	positive	negative	●○○○
31. The enthalpy change when one mole of a substance is completely burnt in excess oxygen under standard conditions is called the standard enthalpy of:	formation	combustion	solution	atomization	○●○○

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Questions	A	B	C	D	A B C D
32. The standard enthalpy of combustion of hydrogen is −285.8 kJ/mol. This is also the standard enthalpy of formation of:	liquid water	ammonia	carbon dioxide	methane	●○○○
33. The standard enthalpy of combustion of graphite is:	−890.4 kJ/mol	−393.5 kJ/mol	−285.8 kJ/mol	−110.5 kJ/mol	○●○○
34. The standard enthalpy of combustion of methane is:	−890.4 kJ/mol	−285.8 kJ/mol	−1560 kJ/mol	−393.5 kJ/mol	●○○○
35. The enthalpy change when one mole of gaseous atoms is formed from the element under standard conditions is the standard enthalpy of:	neutralization	combustion	atomization	formation	○○●○
36. The enthalpy of atomization of chlorine given in the book, for $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$, is:	−121 kJ/mol	+436 kJ/mol	+218 kJ/mol	+121 kJ/mol	○○○●
37. The standard enthalpy of atomization of hydrogen, $H_2(g) \rightarrow H(g)$, per mole of H atoms is:	+436 kJ/mol	+218 kJ/mol	−218 kJ/mol	+121 kJ/mol	○●○○
38. The heat evolved when one mole of H^+ ions from an acid reacts with one mole of OH^- ions from a base to form one mole of water is the standard enthalpy of:	neutralization	combustion	formation	solution	●○○○
39. The standard enthalpy of neutralization of a strong acid with a strong base is about:	−57.4 kJ/mol	−890.4 kJ/mol	−285.8 kJ/mol	+57.4 kJ/mol	●○○○
40. The heat of neutralization of strong acids and bases is the same because the only change is the formation of water from:	Na^+ and OH^-	H^+ and Cl^-	Na^+ and Cl^- ions	H^+ and OH^- ions	○○○●
41. The enthalpy change when one mole of a substance dissolves in so much solvent that further dilution produces no heat change is the standard enthalpy of:	dilution	solution	formation	hydration	○●○○
42. Dissolving NH_4Cl in water has ΔH = +15.1 kJ/mol, so the solution:	remains unchanged	boils	becomes hot	becomes cold	○○○●
43. The first electron affinity is the energy released when one mole of gaseous atoms:	forms molecules	gains one electron each to form singly charged negative ions	gains two electrons each	loses one electron each	○●○○
44. The first electron affinity of chlorine is about:	+349 kJ/mol	−349 kJ/mol	−496 kJ/mol	+121 kJ/mol	○●○○
45. The energy released when one mole of an ionic solid is formed from its gaseous ions is called:	ionization energy	atomization energy	electron affinity	lattice energy	○○○●
46. The lattice energy of NaCl given in the book is:	−349 kJ/mol	−787 kJ/mol	+787 kJ/mol	−411 kJ/mol	○●○○
47. The energy required to break one mole of a particular bond into neutral atoms is called:	hydration energy	lattice energy	electron affinity	bond dissociation energy	○○○●
48. Enthalpy change of a reaction equals the sum of bond energies broken minus the sum of bond energies:	broken	of reactants	formed	of the catalyst	○○●○

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Questions	A	B	C	D	A B C D
49. The bond dissociation energy of H_2 is 436 kJ/mol and that of Cl_2 is 243 kJ/mol. The average bond energy of H-Cl calculated from these ($H_2 + Cl_2 \rightarrow 2HCl$, ignoring ΔH) is about:	339 kJ/mol	193 kJ/mol	431 kJ/mol	679 kJ/mol	●○○○
50. The experimental bond energy of HCl is 431.5 kJ/mol which is higher than the calculated 339 kJ/mol because theoretical values:	are exact	ignore the mass	are averages and ignore electronegativity differences	are for solids	○○●○
51. The bond energies are: H-H 436, O=O 493.6 and O-H 460 kJ/mol. The ΔH for $2H_2 + O_2 \rightarrow 2H_2O(g)$ is about:	+872 kJ	+474 kJ	−474 kJ	−1840 kJ	○○●○
52. In breaking all four C–H bonds of methane, 1662 kJ/mol is absorbed. The average C–H bond energy is:	1662 kJ/mol	831 kJ/mol	415.5 kJ/mol	6648 kJ/mol	○○●○
53. Bond energies of simple diatomic molecules such as H_2, O_2, N_2 and Cl_2 are:	always approximate	zero	relatively accurate	variable	○○●○
54. The bond energy of O-H in methanol differs from that in water because of differences in:	the atomic number of hydrogen	the mass of oxygen	the temperature scale	molecular structure and chemical environment	○○○●
55. The law of conservation of energy states that energy:	can be destroyed but not created	is always lost in reactions	can neither be created nor destroyed	can be created but not destroyed	○○●○
56. Temperature is a measure of the:	volume	average kinetic energy of molecules	density	total mass	○●○○
57. The SI unit of temperature is:	celsius	kelvin	joule	fahrenheit	○●○○
58. The amount of heat required to raise the temperature of a given amount of substance by 1 K is called:	heat capacity	enthalpy	entropy	specific heat capacity	●○○○
59. The amount of heat required to raise the temperature of one gram of a substance by 1 K is called:	specific heat capacity	latent heat	heat capacity	molar enthalpy	●○○○
60. The relationship between heat, specific heat capacity, mass and temperature change is:	$q = m / (c \Delta T)$	$q = m + c + \Delta T$	$q = c / (m \Delta T)$	$q = m\ c\ \Delta T$	○○○●
61. The specific heat capacity of copper is 0.387 J g^{-1} K^{-1}. The heat needed to raise 10 g of Cu by 10 K is:	0.387 J	3.87 J	38.7 J	387 J	○○●○
62. The molar heat capacity of Cu (atomic mass 63.54) is about:	24.6 J mol^{-1} K^{-1}	0.387 J mol^{-1} K^{-1}	63.5 J mol^{-1} K^{-1}	2.46 J mol^{-1} K^{-1}	●○○○
63. A device that measures heat flow is called a:	barometer	manometer	calorimeter	spectrometer	○○●○
64. A coffee-cup calorimeter operates at constant:	mass	pressure	temperature	volume	○●○○
65. 50 cm^3 of 1.0 M NaOH is mixed with 50 cm^3 of 1.0 M HCl. The number of moles of water formed is:	0.5	1.0	0.10	0.05	○○○●

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Questions	A	B	C	D	A B C D
66. In the neutralization calorimeter example, the temperature rises from 25 °C to 31.9 °C. The ΔT is:	25 K	6.9 K	56.9 K	31.9 K	○●○○
67. The heat evolved when 100 g of solution (c = 4.2 J g^{-1} K^{-1}) rises by 6.9 K is about:	290 J	29 kJ	0.29 kJ	2.9 kJ	○○○●
68. In the example, 2.9 kJ is released by 0.05 mole of water formation. The enthalpy of neutralization per mole is about:	−0.145 kJ/mol	−2.9 kJ/mol	−580 kJ/mol	−58 kJ/mol	○○○●
69. Hess's law was stated by G.H. Hess in the year:	1869	1776	1840	1905	○○●○
70. Hess's law states that the enthalpy change in a reaction is the same whether it takes place in:	a single step or in several steps	the presence of a catalyst only	a closed system only	an open container only	●○○○
71. Hess's law is based on the fact that enthalpy is a:	state function	constant	vector	path function	●○○○
72. For a complete energy cycle, the total enthalpy change is:	infinite	zero	negative	positive	○●○○
73. C + ½O_2 → CO (ΔH = −110.52 kJ) and CO + ½O_2 → CO_2 (ΔH = −282.98 kJ). The enthalpy change for C + O_2 → CO_2 is about:	−110.5 kJ	+393.5 kJ	−172.5 kJ	−393.5 kJ	○○○●
74. Hess's law is useful for finding enthalpies of formation of compounds which:	are easily formed	are gases only	cannot be prepared directly from their elements	are liquids only	○○●○
75. The enthalpy of formation of methane cannot be measured directly. It is found using:	colour of methane	enthalpies of combustion of CH_4, H_2 and C	boiling point of methane	density of methane	○●○○
76. The enthalpy of formation of methane obtained by Hess's law in the book is:	−890.4 kJ/mol	−393.5 kJ/mol	−74.6 kJ/mol	+74.6 kJ/mol	○○●○
77. When a thermochemical equation is multiplied by an integer, the ΔH° is:	multiplied by the same integer	divided by the integer	reversed in sign only	unchanged	●○○○
78. When a thermochemical equation is reversed, the sign of ΔH is:	changed	halved	unchanged	doubled	●○○○
79. ΔH° of a reaction from enthalpies of formation is:	Σ $\Delta H_f°$(products) + Σ $\Delta H_f°$(reactants)	Σ $\Delta H_f°$(products) × Σ $\Delta H_f°$(reactants)	Σ $\Delta H_f°$(reactants) − Σ $\Delta H_f°$(products)	Σ $\Delta H_f°$(products) − Σ $\Delta H_f°$(reactants)	○○○●
80. Elements in their standard states are not included in ΔH° calculations because their $\Delta H_f°$ is:	one	negative	positive	zero	○○○●
81. The enthalpy of combustion of ethyne (acetylene) calculated in the book is about:	−1309.5 kJ/mol	−393.5 kJ/mol	−1560 kJ/mol	−2802 kJ/mol	●○○○
82. Combustion of glucose releases about 2818 kJ/mol. The energy released per gram of glucose (molar mass 180 g/mol) is about:	15.7 kJ	2818 kJ	50.7 kJ	0.64 kJ	●○○○

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Questions	A	B	C	D	A B C D
83. The Born–Haber cycle is an application of Hess's law to:	organic compounds	covalent gases	binary ionic compounds	noble gases	○○●○
84. The Born–Haber cycle is used to calculate:	bond angle	molar volume	boiling point	lattice energy	○○○●
85. The standard enthalpy of formation of NaCl is:	+496 kJ/mol	+108 kJ/mol	−411 kJ/mol	−787 kJ/mol	○○●○
86. In the Born–Haber cycle of NaCl, the enthalpy of sublimation of sodium is:	+121 kJ/mol	−349 kJ/mol	+496 kJ/mol	+108 kJ/mol	○○○●
87. In the Born–Haber cycle of NaCl, the first ionization energy of Na is:	−349 kJ/mol	+121 kJ/mol	+496 kJ/mol	+108 kJ/mol	○○●○
88. In the Born–Haber cycle of NaCl, the formation of Cl^- releases energy equal to the:	electron affinity of chlorine	ionization energy of chlorine	lattice energy of NaCl	sublimation energy of sodium	●○○○
89. All steps in the Born–Haber cycle of NaCl except the electron affinity and lattice energy steps are:	endothermic	spontaneous only	neutral	exothermic	●○○○
90. Electron affinities of atoms are usually calculated from the Born–Haber cycle because they are:	difficult to determine directly	always zero	equal to ionization energies	very easy to measure	●○○○
91. The lattice energy of a compound increases with:	increase in ionic radius	increase in ionic charge	decrease in ionic charge	increase in temperature	○●○○
92. As the ionic radius increases, the lattice energy:	increases	remains constant	becomes infinite	decreases	○○○●
93. MgO has a higher lattice energy than NaCl because Mg^{2+} and O^{2-} have:	greater radii	greater charges	smaller charges	fewer electrons	○●○○
94. LiF has a higher lattice energy than KF because Li^+ is:	more massive than K^+	smaller than K^+	larger than K^+	less charged than K^+	○●○○
95. Charge density of an ion is defined as:	ionic mass / ionic charge	ionic charge / ionic radius	ionic charge × ionic radius	ionic radius / ionic charge	○●○○
96. When a solute dissolves, breaking the solute lattice is:	neutral	spontaneous only	endothermic	exothermic	○○●○
97. Hydration of solute particles by solvent molecules is:	endothermic	exothermic	impossible	neutral	○●○○
98. The energy released when gaseous ions are surrounded by water molecules is called the enthalpy of:	hydration	formation	combustion	solution	●○○○
99. Enthalpy of solution is related to lattice energy (reverse lattice) and hydration enthalpy by:	$\Delta H_{solution} = \Delta H_{reverse\ lattice} + \Delta H_{hydration}$	$\Delta H_{solution} = \Delta H_{hydration} / \Delta H_{lattice}$	$\Delta H_{solution} = 0$	$\Delta H_{solution} = \Delta H_{lattice} \times \Delta H_{hydration}$	●○○○

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Questions	A	B	C	D	A B C D
100. For NaCl, lattice energy is −786 kJ/mol and hydration enthalpy −783 kJ/mol. The enthalpy of solution is about:	−1569 kJ/mol	+1569 kJ/mol	−3 kJ/mol	+3 kJ/mol	○○○●
101. When ammonium nitrate dissolves in water (ΔH = +25.7 kJ/mol), the flask becomes:	cold	warm then hot	hot	unchanged	●○○○
102. When NaOH dissolves in water (ΔH = −44.5 kJ/mol), the flask becomes:	unchanged	frozen	cold	hot	○○○●
103. A solute dissolves with release of heat when the solvent binds the solute particles ___ than the crystal does.	less tightly	equally	more tightly	not at all	○○●○
104. The hydration enthalpy becomes more exothermic as the ionic charge:	increases	decreases	remains unchanged	becomes zero	●○○○
105. Smaller ions have ___ hydration enthalpy because of higher charge density.	more exothermic	zero	less exothermic	positive	●○○○
106. The hydration enthalpy of $MgSO_4$ is more exothermic than that of $CaSO_4$ because Mg^{2+} is:	larger than Ca^{2+}	less charged than Ca^{2+}	smaller than Ca^{2+}	heavier than Ca^{2+}	○○●○
107. Three factors affecting electron affinity are nuclear charge, shielding effect and:	boiling point	colour	density	atomic size	○○○●
108. The degree of disorder or randomness of a system is called:	free energy	heat capacity	enthalpy	entropy	○○○●
109. The symbol of entropy is:	H	S	E	G	○●○○
110. Which state of matter has the highest entropy?	Crystal	Gas	Solid	Liquid	○●○○
111. When ice melts, the entropy of the system:	increases	decreases	becomes zero	remains constant	●○○○
112. When water vapour condenses to liquid, the entropy of the system:	increases	decreases	remains constant	doubles	○●○○
113. A positive value of ΔS means that disorder:	increases	remains the same	is zero	decreases	●○○○
114. In general, when the temperature of a system increases, its entropy:	becomes zero	decreases	increases	stays constant	○○●○
115. The entropy change of a reaction is given by:	$\Delta S^\circ = \Sigma S^\circ(\text{products}) \times \Sigma S^\circ(\text{reactants})$	$\Delta S^\circ = \Sigma S^\circ(\text{reactants}) - \Sigma S^\circ(\text{products})$	$\Delta S^\circ = \Sigma S^\circ(\text{products}) + \Sigma S^\circ(\text{reactants})$	$\Delta S^\circ = \Sigma S^\circ(\text{products}) - \Sigma S^\circ(\text{reactants})$	○○○●
116. As the number of gaseous molecules in a reaction increases, the entropy:	increases	becomes zero	stays constant	decreases	●○○○

Questions	A	B	C	D	A B C D
117. For $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, the sign of ΔS is:	cannot be known	negative	positive	zero	○●○○
118. For $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$, the sign of ΔS is:	negative	cannot be known	positive	zero	○○●○
119. Gibbs free energy is also known as:	kinetic energy	heat energy	nuclear energy	available energy	○○○●
120. Gibbs free energy was introduced by Josiah Willard Gibbs in:	1905	1932	1876	1840	○○●○
121. The Gibbs–Helmholtz equation at constant temperature is:	ΔG = ΔS − TΔH	ΔG = TΔH − ΔS	ΔG = ΔH + TΔS	ΔG = ΔH − TΔS	○○○●
122. If ΔG° is negative, the process is:	at equilibrium	impossible	spontaneous in the forward direction	non-spontaneous	○○●○
123. If ΔG° is positive, the process is:	exothermic always	spontaneous	at equilibrium	non-spontaneous in the forward direction	○○○●
124. If ΔG° = 0, the system is:	exothermic	spontaneous	non-spontaneous	at equilibrium	○○○●
125. At the melting point of a solid, the value of ΔG° is:	zero	positive	infinite	negative	●○○○
126. A reaction has ΔH = −100 kJ and ΔS = +0.1 kJ/K. At 300 K, ΔG is:	−70 kJ	−130 kJ	+130 kJ	−100 kJ	○●○○
127. A reaction has ΔH = +50 kJ and ΔS = +0.1 kJ/K. At 300 K, ΔG is:	+80 kJ	−80 kJ	−20 kJ	+20 kJ	○○○●
128. The energy content of food is measured in:	kilopascals	kilocalories	kilograms	kelvins	○●○○
129. One kilocalorie is approximately equal to:	41.84 kJ	418.4 kJ	0.4184 kJ	4.184 kJ	○○○●
130. Enthalpy is defined as the internal energy plus the product of:	pressure and volume	density and volume	mass and velocity	temperature and entropy	●○○○
131. The condition of a system when its properties have definite values is called the:	phase rule	state of the system	path of the system	reaction rate	○●○○
132. Properties determined only by the state of the system, not by how it was reached, are called:	variables	constants	path functions	state functions	○○○●
133. The sum of all kinds of energies of the particles of a system is called its:	entropy	enthalpy	internal energy	free energy	○○●○

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134. Calories in food refer to the energy:	stored in the packet	stored in the food and released by metabolism	lost in cooking	used in packaging	○●○○
135. Cellular respiration in the body is an example of a reaction that:	makes salt	absorbs light	produces nuclear energy	releases energy from food	○○○●
136. The reaction $2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$ has $\Delta H° = -850$ kJ. It is:	exothermic	neither	a physical change	endothermic	●○○○
137. The reaction $N_2(g) + O_2(g) \rightarrow 2NO(g)$ has $\Delta H° = +180.5$ kJ. It is:	spontaneous at all temperatures	a combustion reaction	exothermic	endothermic	○○○●
138. The reaction $H_2(g) + I_2(g) \rightarrow 2HI(g)$ has $\Delta H° = +53.8$ kJ. Its reverse reaction has $\Delta H°$ equal to:	+53.8 kJ	+26.9 kJ	−107.6 kJ	−53.8 kJ	○○○●
139. The standard enthalpy of formation of liquid water is:	−890.4 kJ/mol	−285.8 kJ/mol	+285.8 kJ/mol	−393.5 kJ/mol	○●○○
140. The standard enthalpy of formation of $CO_2(g)$ from C(graphite) and $O_2(g)$ is:	+393.5 kJ/mol	−110.5 kJ/mol	−285.8 kJ/mol	−393.5 kJ/mol	○○○●
141. The enthalpy of combustion of ethanol at 1 atm and 25 °C is −1367 kJ/mol. The energy released by burning one mole of ethanol is:	683.5 kJ	2734 kJ	−1367 kJ	1367 kJ	○○○●
142. The standard state of an element or compound is its most stable physical state at:	1 atm and 298 K	10 atm and 298 K	1 atm and 273 K	0.1 atm and 373 K	●○○○
143. Bond breaking is ___ and bond formation is ___.	exothermic; endothermic	exothermic; exothermic	endothermic; exothermic	endothermic; endothermic	○○●○
144. The difference between the energy needed to break bonds and the energy released on forming bonds determines whether a reaction:	is gas or liquid	is acidic or basic	absorbs or releases energy	is fast or slow	○○●○
145. Heat capacity is expressed in the units:	kJ mol^{-1}	J K^{-1}	g K^{-1}	J g^{-1} K^{-1}	○●○○
146. Specific heat capacity is expressed in the units:	J g^{-1} K^{-1}	kJ mol^{-1}	J K^{-1}	g J^{-1}	●○○○
147. Water in a jar and water in a glass are at the same temperature. The water in the jar has more heat energy because it has:	a higher temperature	faster molecules	more water molecules	stronger bonds	○○●○
148. The energy released on burning 2 moles of methane ($\Delta H° = -890.4$ kJ/mol) is:	3561.6 kJ	890.4 kJ	1780.8 kJ	445.2 kJ	○○●○
149. $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ has $\Delta H° = -571.6$ kJ. The heat released when 1 mole of H_2 burns is:	1143.2 kJ	571.6 kJ	143 kJ	285.8 kJ	○○○●
150. The heat absorbed by a body is directly proportional to its:	temperature change	volume only	colour	density only	●○○○

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